PEARSON

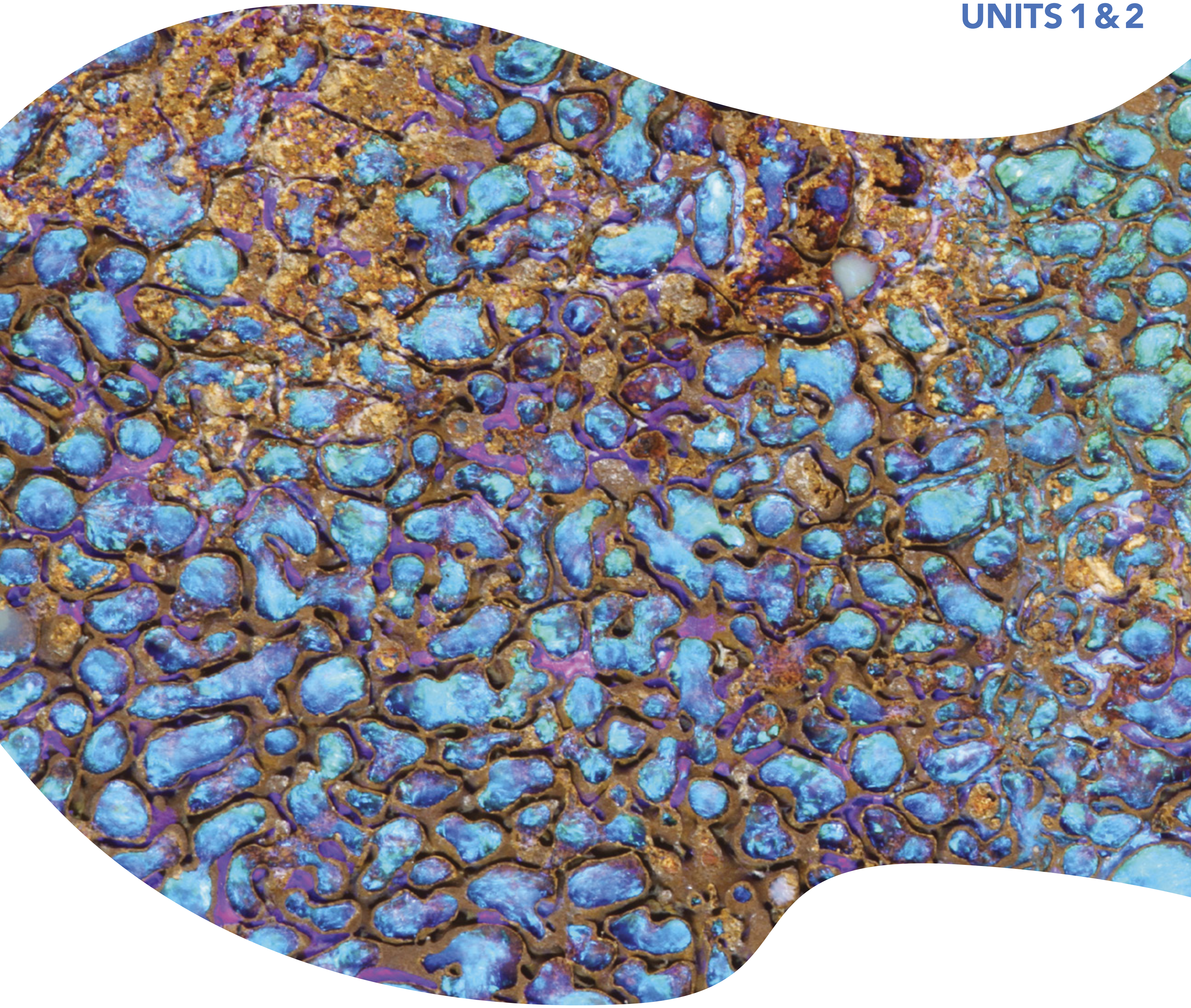
CHEMISTRY

QUEENSLAND

STUDENT BOOK

11

UNITS 1 & 2



QCE 2019 SYLLABUS

PEARSON

CHEMISTRY

QUEENSLAND

STUDENT BOOK

11

UNITS 1 & 2

Pearson Australia
(a division of Pearson Australia Group Pty Ltd)
707 Collins Street, Melbourne, Victoria 3008
PO Box 23360, Melbourne, Victoria 8012
www.pearson.com.au

First published 2019 by Pearson Australia
2022 2021 2020 2019
10 9 8 7 6 5 4 3 2 1

Lead Publisher: Alicia Brown
Project Manager: Michelle Thomas
Production Editor: Laura Pietrobon, Virginia O'Brien
Development Editor: Antonietta Anello
Editor: Marcia Bascombe
Designer: Anne Donald
Rights & Permissions Editor: Lisa Woodland
Senior Publishing Services Analyst: Rob Curulli
Proofreader: Camha Pham
Indexer: Bruce Gillespie
Illustrator: DiacriTech
Printed in China

A catalogue record for this book is available from the National Library of Australia

ISBN: 978 1 4886 1955 7

Pearson Australia Group Pty Ltd ABN 40 004 245 943

This syllabus forms part of a new senior assessment and tertiary entrance system in Queensland. Along with other senior syllabuses, it is still being refined in preparation for implementation in schools from 2019.

For the most current syllabus versions and curriculum information please refer to the QCAA website https://www.qcaa.qld.edu.au/

Disclaimer
The selection of internet addresses (URLs) provided for this book was valid at the time of publication and was chosen as being appropriate for use as a secondary education research tool. However, due to the dynamic nature of the internet, some addresses may have changed, may have ceased to exist since publication, or may inadvertently link to sites with content that could be considered offensive or inappropriate. While the authors and publisher regret any inconvenience this may cause readers, no responsibility for any such changes or unforeseeable errors can be accepted by either the authors or the publisher. Some of the images used in *Pearson Chemistry 11 Queensland Student Book* might have associations with deceased Indigenous Australians. Please be aware that these images might cause sadness or distress in Aboriginal or Torres Strait Islander communities.

Practical activities
All practical activities, including the illustrations, are provided as a guide only and the accuracy of such information cannot be guaranteed. Teachers must assess the appropriateness of an activity and take into account the experience of their students and facilities available. Additionally, all practical activities should be trialled before they are attempted with students and a risk assessment must be completed. All care should be taken: appropriate protective clothing should be worn, the correct equipment used and appropriate preparation and clean up procedures followed whenever conducting any practical activity. Although all practical activities have been written with safety in mind, Pearson Australia and the authors do not accept any responsibility for the information contained in or relating to the practical activities, and are not liable for any loss and/or injury arising from or sustained as a result of conducting any of the practical activities described in this book.

The Chemistry Education Association (CEA) was formed in 1977 by a group of teachers from secondary and tertiary institutions. It aims to promote the teaching of chemistry, particularly in secondary schools.

The CEA has established a tradition of providing up-to-date text, electronic material and support resources for both students and teachers, and professional development opportunities for teachers.

PEARSON
CHEMISTRY 11
QUEENSLAND

Writing and development team

We are grateful to the following people for their time and expertise in contributing to the Pearson Chemistry Unit 1 and 2 project.

Nick Holmes
Head of Science
Queensland Chemistry Series Consultant
Lead author

Elizabeth O'Connell
Head of Science
Queensland Chemistry Subject Lead
Author

Kahlia Ashworth
Senior Science Teacher
Reviewer

Elissa Huddart
VCAA Exam Assessor and Trial Exam Writer (Chemistry)
Queensland Skills & Assessment Book Lead Author

Erin Bruns
VCAA and International Baccalaureate Educator, Medical Laboratory Technician
Author

Dr Chris Commons
Lecturer and Teacher
Past Chief Examiner VCAA Chemistry
Author and Reviewer

Lanna Derry
Author, Head of Science, Past President of Chemistry Education Association
Reviewer

Penny Commons
Tutor and Transition Fellow at the University of Melbourne
Queensland Skills & Assessment Book Lead Author

Claire Hoffman
Lecturer and Tutor, Griffith University
Author

Bob Hogendoorn
VCAA Exam writer and assessor (Chemistry)
Author

Juanita Jacobs
QCAA District Panel Chair, Former Acting SEO, Member of the Expert Writing Team for new syllabus, Chief Assessor
Author

Erica McLean
Head of Science
Reviewer

Professor Richard John
Program Lead Griffith Science Education Alliance, Queensland STEM Education Network
Author

Marguerite van der Klashorst
Teacher, Qualified Pharmacist
Author

Dr Allan Knight
Head of Science, Lecturer, Murdoch University
Author

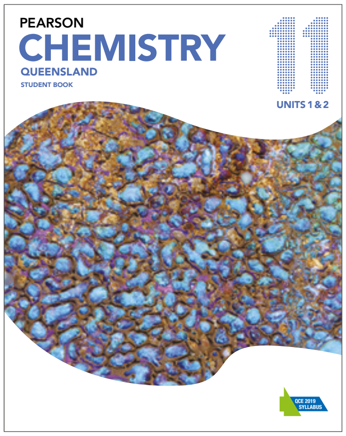
Dr Wendy Mercer
Teacher, QCAA District Panel (Chemistry)
Author

Djaisi Robinson
Head of Science,
Author

Dr Trish Weekes
Science Literacy Consultant
Teacher

Raphael Johns
Chemistry Laboratory Technician

Penny Lee
School Laboratory Technician
Safety Consultant



Professor David F Treagust PhD FRSC FRSB FAERA
John Curtin Distinguished Professor
STEM Education Research Group
School of Education, Curtin University
Consultant and Reviewer

Pearson wishes to acknowledge and thank the following individuals for their contribution to this series:

CONTRIBUTING AUTHORS

Dr Warrick Clarke
Author, Curriculum writer

Pat O'Shea
Author, Head of Science

Maria Porter
Author, Teacher

Geoff Quinton
Author, Head of Science, President of ASTA

Bob Ross
Author, Teacher

Patrick Sanders
Author, Head of Science

Robert Sanders
Author Education Consultant

Chemistry
Education
Association

Pearson wishes to thank the following Queensland Chemistry Teachers for their careful work in answer checking this series:
Michelle Brown
Patricia Corbin
Brenden McDonnell
Erica McLean
Jamie Selby-Pham
Sophie Selby-Pham
Sally Phun
Susan Scheiwe
Mark Spanevello

Contents

ONLINE CHAPTER

Chapter 1 Chemistry Skills and Assessment toolkit

Go to your eBook to access this chapter.
Page numbering begins at 'e1' for this eBook chapter.
Unit 1 starts page 1, followed by Chapter 2.

Unit 1 Chemical fundamentals—structure, properties and reactions

Unit 2: Molecular interactions and reactions

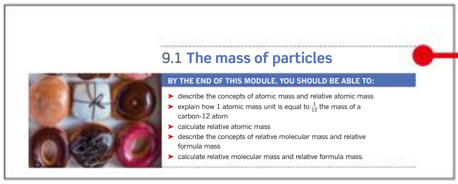
How to use this book

PEARSON CHEMISTRY 11 UNITS 1 & 2 QUEENSLAND

Pearson Chemistry 11 Queensland has been written to the new QCE Chemistry Syllabus. The book is an easy-to-use resource that covers Units 1 & 2 as well as comprehensively addresses the Skills and Assessment.
Explore how to use this book below.

Design

Featuring best-practice literacy and instructional design, this series supports all learners with careful scaffolding of concepts and defined learning objectives.
A simple to navigate, predictable design enables ease of use. The high-quality, relevant photos and illustrations assist student understanding of concepts.



9.1 The mass of particles

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the concepts of atomic mass and relative atomic mass
- explain how 1 atomic mass unit is equal to $\frac{1}{12}$ the mass of a carbon-12 atom
- calculate relative atomic mass
- describe the concepts of relative molecular mass and relative formula mass
- calculate relative molecular mass and relative formula mass.

Chapter opener

The Syllabus subject matter addressed in each chapter is clearly listed, along with any Science as a Human Endeavour features and Mandatory Practicals.

CHAPTER 07 Ionic compounds

Syllabus subject matter

Topic 1 • Properties and structure of atoms

INTRODUCTION TO BONDING

Topic 2 • Properties and structure of materials

BONDING AND PROPERTIES

SCIENCE AS A HUMAN ENDEAVOUR

Module opener

Module openers outline the key concepts and skills developed and link to syllabus subject matter listed in the Chapter opener.

Science as a Human Endeavour

The SHE features provides an opportunity to appreciate the development of science and its use and influence on society.
The SHE features provide a segue into the development of claims and research questions for the Research Investigation. Questions are included to help students formulate ideas and delve more deeply into the concepts.

SCIENCE AS A HUMAN ENDEAVOUR

Nanomaterials

Nanomaterials have unique electrical, catalytic, magnetic, mechanical, thermal and imaging characteristics. This makes them attractive for use in medical, pharmaceutical, electronic and engineering sectors.

Gold nanoparticles in cancer treatment

Gold nanoparticles are the subject of substantial research with a wide range of applications (Figure 6.5.6). One area of development is in using gold nanoparticles as a targeted chemotherapy treatment method.

FIGURE 6.5.6 Applications of gold nanoparticles include cancer treatment research and detecting biological toxins.

Gold nanoparticles can be attached to molecules of a tumour-killing agent known as tumour necrosis factor (TNF). The nanoparticles hide the molecule from the body's immune system.

The nanoparticles carrying TNF tend to accumulate in cancer tumours, allowing TNF to destroy tumours. The nanoparticles do not appear to accumulate in other regions of the body, which means healthy cells are not affected.

Silver nanoparticles kill bacteria

Silver ions have long been known to kill bacteria. The ions can rapidly penetrate bacterial membranes and interact with proteins in the bacteria, destroying the cell structure of the bacteria and preventing them from reproducing.

Technology has enabled silver nanoparticles to be included in many different types of wound dressings. When the dressing (Figure 6.5.7) comes into contact with moisture from the wound, silver nanoparticles are slowly but continuously released from the wound pad. They then enter the wound and kill bacteria.

FIGURE 6.5.7 A wound dressing with silver nanoparticles to kill bacteria

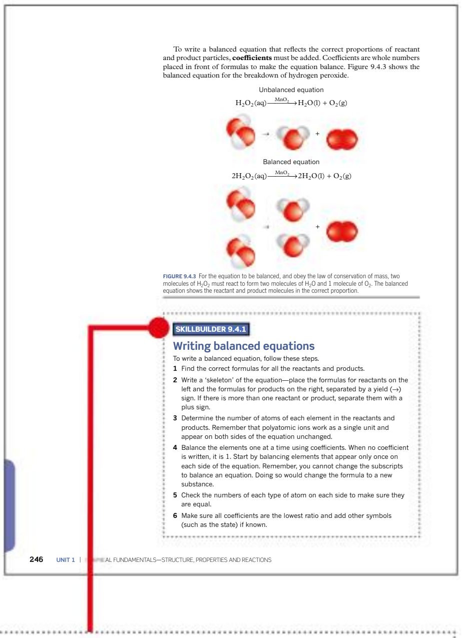
In similar antibacterial applications, Samsung has created and marketed a material called Silver Nano, which adds silver nanoparticles to the surfaces of household appliances. Silver nanoparticles have been embedded in the surfaces of plastic storage bins, as well as in fabrics used by astronauts, babies and outdoor enthusiasts.

Copper nanoparticles go into space

Solder is a filler metal used to join two or more metals. Solders are essential to plumbing and metal constructions, including in satellites and spacecraft. For most of history, solders have contained a high amount of lead. Concerns about the toxicity of lead have driven the development of lead-free solder.

The complex electronics in satellites, such as the solar-powered satellite in Figure 6.5.8 on page 162, must be reliable and efficient over a very long time. Space scientists have developed a nanotechnology copper-based solder that offers far superior performance over the materials currently in use. It is expected that the new solder material will produce up to 10 times the electrical and thermal conductivity of current solders, with a wide range of space and defence applications.

CHAPTER 6 | METALS 161



To write a balanced equation that reflects the correct proportions of reactant and product particles, **coefficients** must be added. Coefficients are whole numbers placed in front of formulas to make the equation balance. Figure 9.4.3 shows the balanced equation for the breakdown of hydrogen peroxide.

Unbalanced equation

$$H_2O_2(aq) \xrightarrow{MnO_2} H_2O(l) + O_2(g)$$

Balanced equation

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

FIGURE 9.4.3 For the equation to be balanced, and obey the law of conservation of mass, two molecules of H_2O_2 must react to form two molecules of H_2O and 1 molecule of O_2. The balanced equation shows the reactant and product molecules in the correct proportion.

SKILLBUILDER 9.4.1

Writing balanced equations

To write a balanced equation, follow these steps.

1 Find the correct formulas for all the reactants and products.
2 Write a 'skeleton' of the equation—place the formulas for reactants on the left and the formulas for products on the right, separated by a yield (→) sign. If there is more than one reactant or product, separate them with a plus sign.
3 Determine the number of atoms of each element in the reactants and products. Remember that polyatomic ions work as a single unit and appear on both sides of the equation unchanged.
4 Balance the elements one at a time using coefficients. When no coefficient is written, it is 1. Start by balancing elements that appear only once on each side of the equation. Remember, you cannot change the subscripts to balance an equation. Doing so would change the formula to a new substance.
5 Check the numbers of each type of atom on each side to make sure they are equal.
6 Make sure all coefficients are the lowest ratio and add other symbols (such as the state) if known.

246 UNIT 1 | [illegible]AL FUNDAMENTALS—STRUCTURE, PROPERTIES AND REACTIONS

Heat (*Q*) is the energy that flows from one object to another because of a difference in temperature.

Highlight box

Highlight features focus students' attention on important information such as key definitions, formulas and salient points.

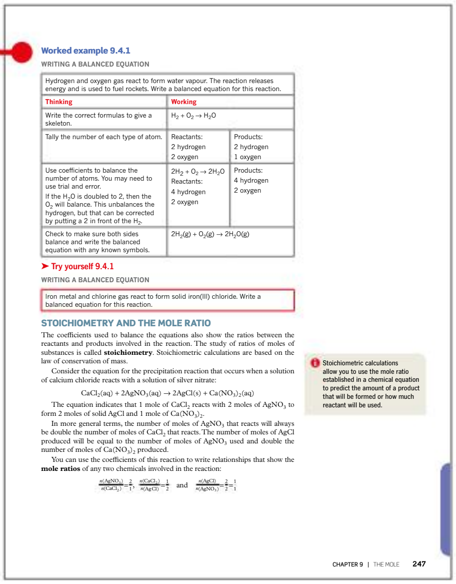
Skillbuilder

A Skillbuilder outlines a method or technique. Each is instructive and self-contained. Skillbuilders step students through the skills to support science application required when analysing or utilising knowledge.

Worked Examples

Worked Examples use sequential steps of thinking and working. This research-based approach greatly enhances student understanding and application of formulas to subject matter. Each Worked Example is followed by a Try Yourself task where students apply their learning to a mirrored problem to practise the skill.

Fully worked solutions to all Try Yourself problems are available on *Pearson Chemistry 11 Queensland* Teacher Support.



Worked example 9.4.1

WRITING A BALANCED EQUATION

Hydrogen and oxygen gas react to form water vapour. The reaction releases energy and is used to fuel rockets. Write a balanced equation for this reaction.		
Thinking	**Working**	
Write the correct formulas to give a skeleton.	$H_2 + O_2 \rightarrow H_2O$	
Tally the number of each type of atom.	Reactants: 2 hydrogen 2 oxygen	Products: 2 hydrogen 1 oxygen
Use coefficients to balance the number of atoms. You may need to use trial and error. If the H_2O is doubled to 2, then the O_2 will balance. This unbalances the hydrogen, but that can be corrected by putting a 2 in front of the H_2.	$2H_2 + O_2 \rightarrow 2H_2O$ Reactants: 4 hydrogen 2 oxygen	Products: 4 hydrogen 2 oxygen
Check to make sure both sides balance and write the balanced equation with any known symbols.	$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$	

➤ **Try yourself 9.4.1**

WRITING A BALANCED EQUATION

Iron metal and chlorine gas react to form solid iron(III) chloride. Write a balanced equation for this reaction.

STOICHIOMETRY AND THE MOLE RATIO

The coefficients used to balance the equations also show the ratios between the reactants and products involved in the reaction. The study of ratios of moles of substances is called **stoichiometry**. Stoichiometric calculations are based on the law of conservation of mass.

Consider the equation for the precipitation reaction that occurs when a solution of calcium chloride reacts with a solution of silver nitrate:

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

The equation indicates that 1 mole of $CaCl_2$ reacts with 2 moles of $AgNO_3$ to form 2 moles of solid AgCl and 1 mole of $Ca(NO_3)_2$.

In more general terms, the number of moles of $AgNO_3$ that reacts will always be double the number of moles of $CaCl_2$ that reacts. The number of moles of AgCl produced will be equal to the number of moles of $AgNO_3$ used and double the number of moles of $Ca(NO_3)_2$ produced.

You can use the coefficients of this reaction to write relationships that show the **mole ratios** of any two chemicals involved in the reaction:

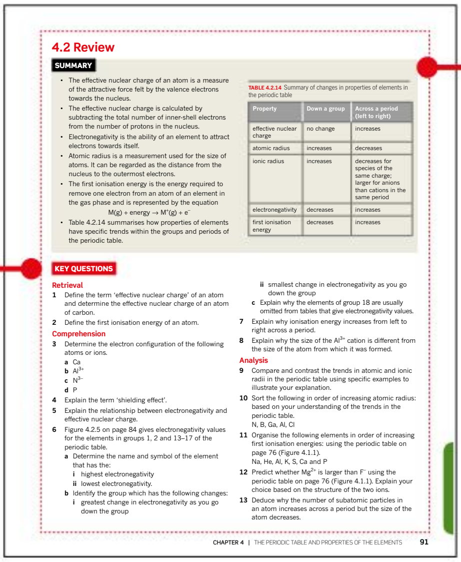
$$\frac{n(AgNO_3)}{n(CaCl_2)}=\frac{2}{1}, \frac{n(CaCl_2)}{n(AgCl)}=\frac{1}{2} \text{ and } \frac{n(AgCl)}{n(AgNO_3)}=\frac{2}{2}=\frac{1}{1}$$

Stoichiometric calculations allow you to use the mole ratio established in a chemical equation to predict the amount of a product that will be formed or how much reactant will be used.

CHAPTER 9 | THE MOLE 247

Module review

Each module finishes with key questions to test students' understanding and ability to recall the key concepts of the module. Questions are carefully categorised under the relevant cognitive level—Retrieval, Comprehension or Analysis—and are developed to assess the syllabus requirements.



4.2 Review

SUMMARY

- The effective nuclear charge of an atom is a measure of the attractive force felt by the valence electrons towards the nucleus.
- The effective nuclear charge is calculated by subtracting the total number of inner-shell electrons from the number of protons in the nucleus.
- Electronegativity is the ability of an element to attract electrons towards itself.
- Atomic radius is a measurement used for the size of atoms. It can be regarded as the distance from the nucleus to the outermost electrons.
- The first ionisation energy is the energy required to remove one electron from an atom of an element in the gas phase and is represented by the equation
  $M(g) + energy \rightarrow M^+(g) + e^-$
- Table 4.2.14 summarises how properties of elements have specific trends within the groups and periods of the periodic table.

TABLE 4.2.14 Summary of changes in properties of elements in the periodic table

Property	Down a group	Across a period (left to right)
effective nuclear charge	no change	increases
atomic radius	increases	decreases
ionic radius	increases	decreases for species of the same charge; larger for anions than cations in the same period
electronegativity	decreases	increases
first ionisation energy	decreases	increases

KEY QUESTIONS

Retrieval

1 Define the term 'effective nuclear charge' of an atom and determine the effective nuclear charge of an atom of carbon.

2 Define the first ionisation energy of an atom.

Comprehension

3 Determine the electron configuration of the following atoms or ions.
  a Ca
  b Al^{3+}
  c N^{3-}
  d P

4 Explain the term 'shielding effect'.

5 Explain the relationship between electronegativity and effective nuclear charge.

6 Figure 4.2.5 on page 84 gives electronegativity values for the elements in groups 1, 2 and 13–17 of the periodic table.
  a Determine the name and symbol of the element that has the:
    i highest electronegativity
    ii lowest electronegativity.
  b Identify the group which has the following changes:
    i greatest change in electronegativity as you go down the group
    ii smallest change in electronegativity as you go down the group
  c Explain why the elements of group 18 are usually omitted from tables that give electronegativity values.

7 Explain why ionisation energy increases from left to right across a period.

8 Explain why the size of the Al^{3+} cation is different from the size of the atom from which it was formed.

Analysis

9 Compare and contrast the trends in atomic and ionic radii in the periodic table using specific examples to illustrate your explanation.

10 Sort the following in order of increasing atomic radius: based on your understanding of the trends in the periodic table.
N, B, Ga, Al, Cl

11 Organise the following elements in order of increasing first ionisation energies: using the periodic table on page 76 (Figure 4.1.1).
Na, He, Al, K, S, Ca and P

12 Predict whether Mg^{2+} is larger than F^- using the periodic table on page 76 (Figure 4.1.1). Explain your choice based on the structure of the two ions.

13 Deduce why the number of subatomic particles in an atom increases across a period but the size of the atom decreases.

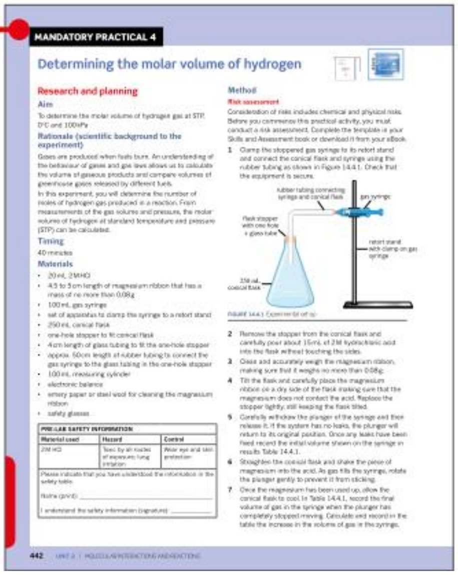
CHAPTER 4 | THE PERIODIC TABLE AND PROPERTIES OF THE ELEMENTS 91

Module summary

Each module concludes with a summary to help students consolidate the key points and concepts.

Mandatory practicals

All Mandatory practicals are included in the Student Book and have been comprehensively developed to ensure they fully address the syllabus requirements. Each practical has been trialled and tested to ensure it can be safely performed and yields effective results, and includes a depth of questions and applications that enable students to develop and demonstrate required manipulative skills.



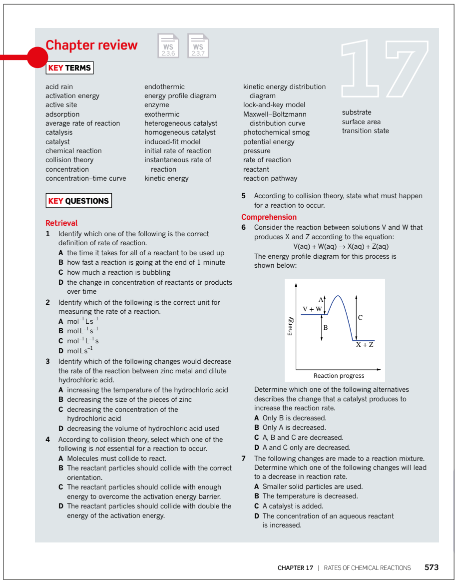
MANDATORY PRACTICAL 4

Determining the molar volume of hydrogen

How to use this book

Chapter review

Each chapter finishes with a list of key terms covered in the chapter and a set of questions to test students' abilities to apply the knowledge gained from the chapter.



Chapter review WS 2.3.6 WS 2.3.7 17

KEY TERMS

acid rain
activation energy
active site
adsorption
average rate of reaction
catalysis
catalyst
chemical reaction
collision theory
concentration
concentration–time curve
endothermic
energy profile diagram
enzyme
exothermic
heterogeneous catalyst
homogeneous catalyst
induced-fit model
initial rate of reaction
instantaneous rate of reaction
kinetic energy
kinetic energy distribution diagram
lock-and-key model
Maxwell–Boltzmann distribution curve
photochemical smog
potential energy
pressure
rate of reaction
reactant
reaction pathway
substrate
surface area
transition state

KEY QUESTIONS

Retrieval

1 Identify which one of the following is the correct definition of rate of reaction.
A the time it takes for all of a reactant to be used up
B how fast a reaction is going at the end of 1 minute
C how much a reaction is bubbling
D the change in concentration of reactants or products over time

2 Identify which of the following is the correct unit for measuring the rate of a reaction.
A $mol^{-1} L s^{-1}$
B $mol L^{-1} s^{-1}$
C $mol^{-1} L^{-1} s$
D $mol L s^{-1}$

3 Identify which of the following changes would decrease the rate of the reaction between zinc metal and dilute hydrochloric acid.
A increasing the temperature of the hydrochloric acid
B decreasing the size of the pieces of zinc
C decreasing the concentration of the hydrochloric acid
D decreasing the volume of hydrochloric acid used

4 According to collision theory, select which one of the following is *not* essential for a reaction to occur.
A Molecules must collide to react.
B The reactant particles should collide with the correct orientation.
C The reactant particles should collide with enough energy to overcome the activation energy barrier.
D The reactant particles should collide with double the energy of the activation energy.

5 According to collision theory, state what must happen for a reaction to occur.

Comprehension

6 Consider the reaction between solutions V and W that produces X and Z according to the equation:

$$V(aq) + W(aq) \rightarrow X(aq) + Z(aq)$$

The energy profile diagram for this process is shown below:

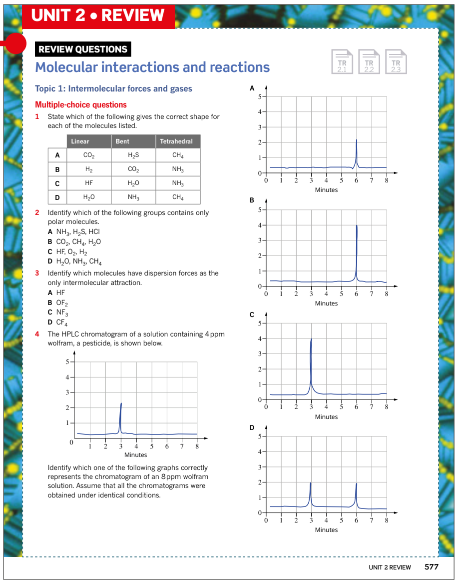
Determine which one of the following alternatives describes the change that a catalyst produces to increase the reaction rate.
A Only B is decreased.
B Only A is decreased.
C A, B and C are decreased.
D A and C only are decreased.

7 The following changes are made to a reaction mixture. Determine which one of the following changes will lead to a decrease in reaction rate.
A Smaller solid particles are used.
B The temperature is decreased.
C A catalyst is added.
D The concentration of an aqueous reactant is increased.

CHAPTER 17 | RATES OF CHEMICAL REACTIONS 573

Unit review

Each Unit concludes with a comprehensive set of exam-style questions, including multiple choice and short answer, that assist students to draw together their knowledge and understanding of the whole Unit.



UNIT 2 • REVIEW

REVIEW QUESTIONS

Molecular interactions and reactions

TR 2.1 TR 2.2 TR 2.3

Topic 1: Intermolecular forces and gases

Multiple-choice questions

1 State which of the following gives the correct shape for each of the molecules listed.

	Linear	Bent	Tetrahedral
A	CO_2	H_2S	CH_4
B	H_2	CO_2	NH_3
C	HF	H_2O	NH_3
D	H_2O	NH_3	CH_4

2 Identify which of the following groups contains only polar molecules.
A NH_3, H_2S, HCl
B CO_2, CH_4, H_2O
C HF, O_2, H_2
D H_2O, NH_3, CH_4

3 Identify which molecules have dispersion forces as the only intermolecular attraction.
A HF
B OF_2
C NF_3
D CF_4

4 The HPLC chromatogram of a solution containing 4 ppm wolfram, a pesticide, is shown below.

Identify which one of the following graphs correctly represents the chromatogram of an 8 ppm wolfram solution. Assume that all the chromatograms were obtained under identical conditions.

UNIT 2 REVIEW 577

Icons

Go To icons make important links to relevant content within the student books in the course. The Go To icons indicate where to engage with Chapter 1 in your eBook.

GO TO ➤

Every Mandatory practical is supported by a complementary **SPARKlab** alternative practical.

The ***Pearson Chemistry 11 Skills and Assessment Book*** icons indicate the best time to engage with an activity for practice, application and revision. The type of activity is indicated as follows:

Worksheet (WS) WS 3.1

Practical Activity (PA) PA 3.2

Mandatory Practical (MP) MP 3.2

Sample Assessment Task (SAT). SAT 3.3

The **Reader+** icon indicates when to engage with an asset via your reader+ eBook. Assets may include videos and interactive activities.

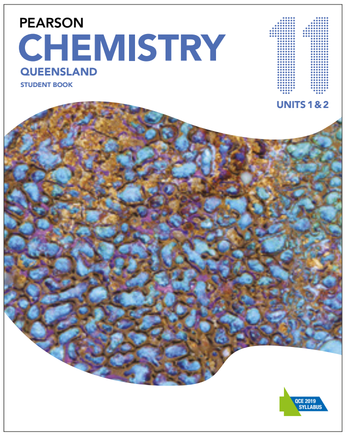
Glossary

Key terms are shown in **bold** throughout the Student Book and are listed at the end of each chapter. A comprehensive glossary at the end of the book defines all the key terms. The glossary aligns with the syllabus context and includes the QCAA defined terminology.

Answers

The Teacher Reader+ eBook provides comprehensive answers and fully worked solutions for all module reviews, Try yourself, Science as a Human Endeavour, chapter reviews and Unit reviews.

Pearson Chemistry 11 Queensland



Student Book

Pearson Chemistry 11 Units 1 & 2 Queensland has been developed by experienced Queensland teachers to address all the requirements of the new QCE Chemistry 2019 Syllabus. The series features the very latest developments and applications of chemistry, literacy and instructional design to ensure the content and concepts are fully accessible to all students.

Skills and Assessment Book

The *Pearson Chemistry 11 Skills and Assessment Book* gives students the edge in preparing for all forms of assessment. Specifically prepared to provide opportunities to consolidate, develop and apply subject matter and science inquiry skills, this resource features a toolkit, key knowledge summaries, worksheets, practical activities and guidance, assessment practice and exam-style Topic Review sets.

Reader+ the next generation eBook

Pearson Reader+ lets you use your Student Book online or offline, anywhere and anytime, on any device. Pearson Reader+ retains the look and integrity of the printed book.

Teacher Support

Pearson Chemistry 11 Units 1 & 2 Queensland Teacher Support provides:

- complete answers, fully worked solutions or suggested answers to all *Student Book and Skills and Assessment Book* tasks
- mandatory practical expected results, common mistakes, suggested answers and full safety notes and risk assessments
- teaching, learning and assessment programs

Pearson

Access your digital resources at **pearsonplaces.com.au**
Browse and buy at **pearson.com.au**

Understanding Assessment Instructions

The Queensland Certificate of Education (QCE) 2019 Chemistry Syllabus uses an inquiry approach, guiding students in a systematic way to a better understanding of the world.

You will notice as you read through the Syllabus for Units 1 to 4, that many terms are underlined. Two examples are provided:

a the General Senior Syllabus objectives
b a sample of Topic 1 subject matter.

The General Senior syllabus objectives showing the underlining of terms

Syllabus objective		Unit 1	Unit 2	Unit 3	Unit 4
1.	describe and explain scientific concepts, theories, models and systems and their limitations	●	●	●	●
2.	apply understanding of scientific concepts, theories, models and systems within their limitations	●	●	●	●
3.	analyse evidence	●	●	●	●
4.	interpret evidence	●	●	●	●
5.	investigate phenomena	●	●	●	●
6.	evaluate processes, claims and conclusions	●	●	●	●
7.	communicate understandings, findings, arguments and conclusions.	●	●	●	●

The QCAA Syllabus states that students 'are required to use a range of cognitive processes to demonstrate they meet the syllabus'. Many of the underlined words in the Syllabus are action verbs. These verbs are often placed at the start of dot points, to identify the level of thinking (cognitive process) you are expected to demonstrate. Note the action verbs from the two Syllabus extracts.

Remember, the cognitive verbs used in the Syllabus subject matter dot points, indicate the highest level of thinking and subject engagement to be covered. You will not be assessed at a higher cognitive level.

A subject matter description from Chemistry Unit 1, Topic 1 showing the underlining of terms

Subject matter

Periodic table and trends

- recall that elements are represented by symbols and recognise that the structure of the periodic table is based on the atomic number and the properties of the elements
- describe and explain that elements of the periodic table show trends across periods and down groups, including atomic radii, valencies, ionic radii, 1st ionisation energy and electronegativities as exemplified by groups 1, 2, 13–18 and period 3
- explain how successive ionisation energy data is related to the electron configuration of an atom
- compare and explain the metallic and non-metallic behaviours of elements, including group trends and the reactivity for the alkali metals (Li–Cs) and the halogens (F–I)
- recognise that oxides change from basic through amphoteric to acidic across a period
- analyse, evaluate and interpret data to explain and justify conclusions for periodic trends, patterns and relationships.

UNDERSTANDING COGNITIVE PROCESSES AND VERBS

It is important to understand that the verbs driving the syllabus objectives and topic subject matter are not randomly chosen. By gaining a better understanding of cognitive verbs, you will be able to respond more satisfactory to questions and instructions in assessment tasks.

Cognitive verbs are signals to the learner, of the type of thinking to be demonstrated. For example:

- the verb *evaluate* indicates that an assessment or judgement must be made
- the verb *describe* requires that an account or outline be provided.

There is a difference between the thinking needed by each of these verbs. To *evaluate* is of a higher level of thinking than to *describe*. Generally, the higher the thinking level required in a task, the more challenging it is.

Cognitive verbs can be arranged or classified into different levels of thinking (also known as cognitive processes) ranging from remembering to complex thinking. The QCAA Syllabus uses an arrangement (taxonomy) of cognitive processes devised by educational researchers Robert Marzano and John Kendall.

In this arrangement, four levels of cognitive process are identified: retrieval, comprehension, analysis and knowledge utilisation. An outline of these levels is provided in Chart A. A large number of different cognitive verbs are used in the Syllabus. These verbs can be aligned with different levels of thinking as shown in Chart B.

Pearson Chemistry 11 Queensland Student Book provides a comprehensive number of questions and instructions. The review instruction sets are arranged by cognitive levels using the Marzano and Kendall taxonomy and provide students with the opportunity to demonstrate knowledge and application of the subject matter at the following levels:

Module review – Retrieval, Comprehension and Analysis

Chapter review – Retrieval, Comprehension, Analysis and Knowledge utilisation

Unit review – Retrieval, Comprehension, Analysis and Knowledge utilisation.

CHART A Cognitive processes, as arranged by Marzano and Kendall

Cognitive processes – levels of thinking (Marzano and Kendall taxonomy)

Retrieval	Comprehension	Analysis	Knowledge utilisation
Level 1 – basic level of thinking • Involves remembering, recalling, recognising and executing information.	**Level 2** – higher level of thinking than Retrieval • Involves understanding and identifying key information.	**Level 3** – more complex thinking than comprehension • Involves examination of information and the identification and separation into its separate parts.	**Level 4** – most complex thinking level • Involves applying information to investigate, experiment, problem solve and make decisions.

Increasing complexity of thinking

Each level of thinking builds upon lower levels. For example, you must be able to retrieve information and comprehend it before you can analyse it.

CHART B Cognitive processes, associated verbs with sample questions and instructions

Cognitive processes, associated verbs and sample questions

Retrieval		Comprehension		Analysis		Knowledge utilisation	
Processes: • recognising • recalling • symbolising		**Processes:** • recognising • recalling • symbolising		**Processes:** • recognising • recalling • symbolising		**Processes:** • recognising • recalling • symbolising	
Cognitive VERBS		**Cognitive VERBS**		**Cognitive VERBS**		**Cognitive VERBS**	
define demonstrate describe identify indicate label list name	paraphrase recall recognise (e.g. features) select show state use	calculate (e.g. numerical answer; mathematical processes) clarify comprehend (meaning) construct (e.g. a diagram) demonstrate describe determine develop discuss	draw (visual depiction) explain illustrate implement (e.g. plan, proposal) recognise (e.g. features) represent select show summarise symbolise (e.g. through diagram, illustration, model) understand use	analyse apply assess calculate (e.g. numerical answer; mathematical processes) categorise classify compare conclude consider contrast critique deduce derive determine diagnose differentiate	discriminate distinguish edit evaluate extrapolate explore identify errors/problems infer interpret e.g. meaning judge organise/sequence/structure predict reflect (on) sort scrutinise	adapt appraise appreciate argue assess comment (make a judgement) conduct (e.g. investigations) conclude construct (e.g. an argument) convince create design (e.g. a methodology, an artefact, a proposal) decide determine develop (e.g. a strategy, product or process) devise discuss/explore draw conclusions evaluate	experiment/test (e.g. hypotheses) (e.g. information, ideas, components) into a whole, in order to create new meaning generate/test (e.g. hypotheses) hypothesise/propose (e.g. arguments, concept) investigate/examine (e.g. an argument, statement or conclusion) judge justify/prove (e.g. an argument, statement or conclusion) make decisions manipulate (e.g. language texts; skills; technologies) modify persuade predict (e.g. a result) propose prove research realise/resolve (e.g. artistic works) solve (e.g. problems) synthesise (e.g. information, ideas, components) test

ℹ Note that some cognitive verbs appear in more than one cognitive level.

ℹ Note that a question may not necessarily include a cognitive verb.

CHART B continued

Cognitive processes, associated verbs and sample questions			
Sample instruction and question	**Sample instruction and question**	**Sample instruction and question**	**Sample instruction and question**
Identify the electron configuration of Cr and Cu. What is the electron configuration of Cr and Cu?	**Explain** how successive ionisation energy data is related to the electron configuration of an atom. How is the electron configuration of an atom related to successive ionisation energy data?	**Compare and explain** the metallic and non-metallic behaviour of elements for Li–Cs and F–I. What are the similarities and differences between the metallic and non-metallic behaviours of elements?	**Make predictions** based on supplied solubility curves. What prediction can you make from the solubility curve with which you are supplied?

UNDERSTANDING THE INSTRUCTION/QUESTION

The cognitive verb alone is not enough of a guide to understanding what the question or instruction requires as a response. As seen in the previous chart, questions may not include a cognitive verb. In addition, a cognitive verb may apply to more than one level of thinking.

Consider these examples:

Examples	Use of the cognitive verb 'explain'	Scope and context of the instruction or question
	What is the question asking?	**How much do I write?**
Example 1a: Explain diffusion. What is diffusion?	The key aspect/s of the question: **a** definition response—retrieval **b** paraphrase response—comprehension	A question that focuses on: **a** one key term—diffusion
Example 1b: Explain convection. What is convection?		A question that focuses on one concept: **a** convection
Example 2a: Explain the **difference** between simple and facilitated diffusion. What is **different** between simple and facilitated diffusion?	The key aspect/s of the question: **a** compare and contrast—analysis	A question that focuses on: **a** two terms: simple diffusion and facilitated diffusion **b** differences between the two terms
Example 2b: Explain the **difference** between convection and conduction. What is **different** between convection and conduction?		A question that focuses on: **a** two terms: convection and conduction **b** differences between the two terms
Example 3a: Explain how you would conduct an experiment using chromatography to **analyse** proteins. What experimental procedure would you use to **analyse** proteins using chromatography?	The key aspect/s of the question: **a** create, design, experiment **b** knowledge utilisation	A question that focuses on: **a** two terms: proteins and chromatography **b** their analysis **c** design of an experiment
Example 3b: Explain how you would conduct an experiment using convection to **analyse** the specific heat capacity of a metal. What experimental procedure would you use to analyse the specific heat capacity of a metal using convection?		A question that focuses on: **a** two terms: convection and specific heat capacity **b** their analysis **c** design of an experiment

Strategies for understanding the question or instruction

As you can see, analysing a question or instruction can be quite challenging. This particularly applies when more complex cognitive processes are required.

The following steps provide a framework for understanding and analysing questions and instructions.

Strategies for understanding the question or instruction	**Example** Element X forms a compound with oxygen in which the ratio of the number of atoms of X to the number of atoms of oxygen is 2:5. The element makes up 59.7% of the mass of the compound. Calculate the molar mass of X and use it to identify this element.
1 Underline the cognitive verb/s and identify a plausible thinking level for the verb.	'Calculate the molar mass of X and use it to identify this element.' **Calculate** involves assessment, deductive thinking, so has **analysis** thinking level qualities. **Identify** is found on the list for both **retrieval** and **analysis**.
2 Determine the scope, context of the question and its thinking level.	The scope of the question includes calculations using given data and matching against the periodic table. The context is the understanding and analysing of the data. This targets the analysis cognitive level.
3 Consider some cognitive actions from Chart B are required to complete the question.	The question is complex. It asks for **analysis** of the data provided then deduction based on that analysed data.
4 Make sure you know the meaning of every word in the question or instruction.	**Calculate means** to 'determine or find (e.g. a number, answer) by using mathematical processes; obtain a numerical answer showing the relevant stages in the working; ascertain/determine from given facts, figures or information.' **Identify** means to 'distinguish, locate, recognise and name; establish or indicate who or what someone or something is; provide an answer from a number of possibilities; recognise and state a distinguishing factor or feature'. [QCAA definitions]
5 Rephrase the question or instruction in your own words, elaborating on all details required.	Rephrasing may be in the form of a question or instruction. For example: 'I can look up the molar mass of oxygen and I know oxygen makes up 40.3% of this compound and the unknown element is 59.7% of the compound. Use this to work out the molar mass of the unknown element.'

ASSESSMENT TASKS AND COGNITIVE PROCESSES

Pearson Chemistry 11 Units 1 & 2 Queensland Student Book provides a solid foundation for undertaking all assessment tasks in the Syllabus. Comprehensive sets of key instructions and tasks, arranged by cognitive thinking levels, and mirroring the instruction types of the examination, are provided at the end of each module and chapter.

Module reviews have tasks under Retrieval, Comprehension and Analysis.

Chapter review tasks cover these three levels as well as Knowledge utilisation.

In addition, **Unit reviews** provide the opportunity to consolidate and test you on a broader area of subject matter.

Mandatory practicals provide support in this skill area through practice in this cognitive level.

This approach will support you in developing the skills and level of application required to complete the assessment tasks.

You are required to complete the following assessment tasks:

a Data test
b Student experiment
c Research investigation
d Examination

The following charts provide an indication of the cognitive processes you should expect to encounter in each type of assessment task. Note that the student experiment and research investigation assessment tasks are designed for thinking at the highest cognitive levels. While retrieval, comprehension and analysis are required to complete the tasks, they are the underlying thinking levels necessary to complete the tasks. Hence the differences in the sizes of the ticks, with the largest tick indicating the focal cognitive thinking level.

Data test (IA1)

Retrieval	Comprehension	Analysis	Knowledge utilisation	
✔	✔	✔	✔	The task requires you to demonstrate thinking that is complex and at the high levels of analysis and knowledge utilisation. Retrieval and comprehension underlie the thinking so data can be analysed in the test.

Student experiment (IA2)

Retrieval	Comprehension	Analysis	Knowledge utilisation	
✔	✔	✔	✔	The task requires you to demonstrate thinking that is complex and at a high level. Retrieval, comprehension and analysis underlie the experimenting and problem-solving required for this task.

Research investigation (IA3)

Retrieval	Comprehension	Analysis	Knowledge utilisation	
✔	✔	✔	✔	The task requires you to demonstrate thinking that is complex and at a high level. Retrieval, comprehension and analysis underlie the investigation and decision-making required for this task.

The examination (EA) will include two papers. Each paper consists of a number of different types of items, including short and combination responses.

	Retrieval	Comprehension	Analysis	Knowledge utilisation	
Short response: • multiple choice • single-word • sentences • calculating using algorithms.	✔	✔	✔	✔	Short responses generally draw on factual subject matter in the retrieval and comprehension cognitive processes areas but may require analysis where calculations and data interpretation are involved.
Combination response: • short items requiring single-word, sentence or short paragraph responses • calculating using algorithms • interpreting graphs, tables or diagrams • responding to unseen data and/or stimulus.	✔	✔	✔	✔	The calculations and responses to unseen data move the cognitive processes required to the highest levels of thinking.

DEFINITIONS OF COGNITIVE VERBS

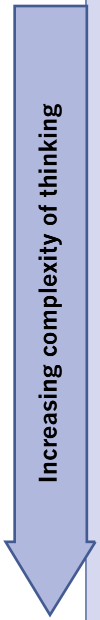
The list that follows provides definitions for cognitive verbs. Where available, the definitions are taken from the QCAA Syllabus. Those verbs whose definitions are not in the QCAA Syllabus appear in grey text. Refer to the list to clarify exactly what is required when any of these verbs appear in a question or instruction. Verbs are organised according to cognitive levels of thinking.

Increasing complexity of thinking

Level of thinking	Cognitive verb	Definition of cognitive verb
Retrieval: **processes of recognising, recalling, symbolising**	define	give the meaning of a word, phrase, concept or physical quantity; state meaning and identify or describe qualities
	demonstrate	prove or make clear by argument, reasoning or evidence, illustrating with practical example; show by example; give a practical exhibition
	describe	give an account (written or spoken) of a situation, event, pattern or process, or of the characteristics or features of something
	identify	distinguish; locate, recognise and name; establish or indicate who or what someone or something is; provide an answer from a number of possibilities; recognise and state a distinguishing factor or feature
	indicate	suggest, show or recommend a course of action
	label	Identify by applying a name to an object or person
	list	write the names of connected items, usually one below the other
	name	specify or give a label to an object or person
	paraphrase	use different words to convey the same meaning
	recall	remember; present remembered ideas, facts or experiences; bring something back into thought, attention or into one’s mind
	recognise	identify or recall particular features of information from knowledge; identify that an item, characteristic or quality exists; perceive as existing or true; be aware of or acknowledge
	select	choose in preference to another or others; pick out
	show	provide the relevant reasoning to support a response
	state	express something definitely and clearly
	use	operate or put into effect; apply knowledge or rules to put theory into practice

Increasing complexity of thinking

Level of thinking	Cognitive verb	Definition of cognitive verb
Comprehension: **processes of integrating, symbolising**	calculate (e.g. numerical answer, mathematical processes)	work out using mathematical processes and determine by reasoning
	clarify	make clear or intelligible; explain: make a statement or situation less confused or more comprehensible
	comprehend (meaning)	understand the meaning or nature of; grasp mentally
	construct	create or put together (e.g. an argument) by arranging ideas or items; display information in a diagrammatic or logical form; make; build
	demonstrate	prove or make clear by argument, reasoning or evidence, illustrating with practical example; show by example; give a practical exhibition
	describe	give an account (written or spoken) of a situation, event, pattern or process, or of the characteristics or features of something
	determine	establish, conclude or ascertain after consideration, observation or calculation; decide or come to a resolution
	develop	elaborate, expand or enlarge in detail; add detail and fullness to; cause to become more complex or intricate
	discuss	examine by argument; sift the considerations for and against; debate; talk or write about a topic, including a range of arguments, factors or hypotheses; consider, taking into account different issues and ideas, points for and/or against, and supporting opinions or conclusions with evidence
	draw (visual depiction)	produce a picture, diagram or other visual representation
	explain	make an idea or situation plain or clear by describing it in more detail or revealing relevant facts; give an account; provide additional information
	illustrate	provide pictures, provide an example for a point being made
	implement	put something into effect, e.g. a plan or proposal
	recognise	identify or recall particular features from knowledge; identify that an item, characteristic or quality exists; perceive as existing or true; be aware of or acknowledge
	represent	scientific representations are a verbal, physical or mathematical demonstration of understanding of a science concept or concepts; a concept can be represented in a range of ways and using multiple models (ACARA 2015c)
	select	choose in preference to another or others; pick out
	show	provide the relevant reasoning to support a response
	summarise	give a brief statement of a general theme or major point/s; present ideas and information in fewer words and in sequence
	symbolise	represent or identify by a symbol or symbols
	understand	perceive what is meant by something; grasp; be familiar with (e.g. an idea); construct meaning from messages, including oral, written and graphic communication
	use	operate or put into effect; apply knowledge or rules to put theory into practice



Level of thinking	Cognitive verb	Definition of cognitive verb
Analysis: **processes of matching, classifying, analysing errors, generalising, specifying**	analyse	dissect to ascertain and examine constituent parts and/or their relationships; break down or examine in order to identify the essential elements, features, components or structure; determine the logic and reasonableness of information; examine or consider something in order to explain and interpret it, for the purpose of finding meaning or relationships and identifying patterns, similarities and differences
	apply	use knowledge and understanding in response to a given situation or circumstance; carry out or use a procedure in a given or particular situation
	assess	measure, determine, evaluate, estimate or make a judgement about the value, quality, outcomes, results, size, significance, nature or extent of something
	calculate	work out using mathematical processes and determine by reasoning
	categorise	place in or assign to a particular class or group; arrange or order by classes or categories; classify, sort out, sort, separate
	classify	arrange, distribute or order in classes or categories according to shared qualities or characteristics
	compare	compare: display recognition of similarities and differences and recognise the significance of these similarities and differences
	conclude	judgement based on evidence (ACARA 2015c)
	consider	think deliberately and carefully about something, typically before making a decision; taking something into account when making a judgement; view attentively or scrutinise; reflect on
	contrast	contrast: display recognition of differences by deliberate juxtaposition of contrary elements; show how things are different or opposite; give an account of the differences between two or more items or situations, referring to both or all of them throughout
	critique	review (e.g. a theory, practice, performance) in a detailed, analytical and critical way
	deduce	reach a conclusion that is necessarily true, provided a given set of assumptions is true; arrive at, reach or draw a logical conclusion from reasoning and the information given
	derive	arrive at a reasoning: manipulate a mathematical relationship to give a new equation or relationship; in mathematics, obtain the derivative of a function
	determine	establish, conclude or ascertain after consideration, observation, investigation or calculation; decide or come to a resolution
	diagnose	identify the nature of a problem or illness
	differentiate	identify the difference/s in or between two or more things; distinguish, discriminate; recognise or ascertain what makes something distinct from similar things; in mathematics, obtain the derivative of a function
	discriminate	note, observe or recognise a difference; make or constitute a distinction in or between; differentiate; note or distinguish as different
	distinguish	recognise as distinct or different; note points of difference between; discriminate; discern; make clear a difference/s between two or more concepts or items
	edit	correct written material by careful checking
	evaluate	make an appraisal by weighing up or assessing strengths, implications and limitations; make judgements about ideas, works, solutions or methods in relation to selected criteria; examine and determine the merit, value or significance of something, based on criteria
	extrapolate	infer or estimate by extending or projecting known information; conjecture; infer from what is known; extend the application of something (e.g. a method or conclusion) to an unknown situation by assuming that existing trends will continue or similar methods will be applicable
	explore	look into both closely and broadly; scrutinise; inquire into or discuss something in detail
	identify (errors, problems)	recognise and establish things such as groupings of similar items, mistakes or issues

Increasing complexity of thinking

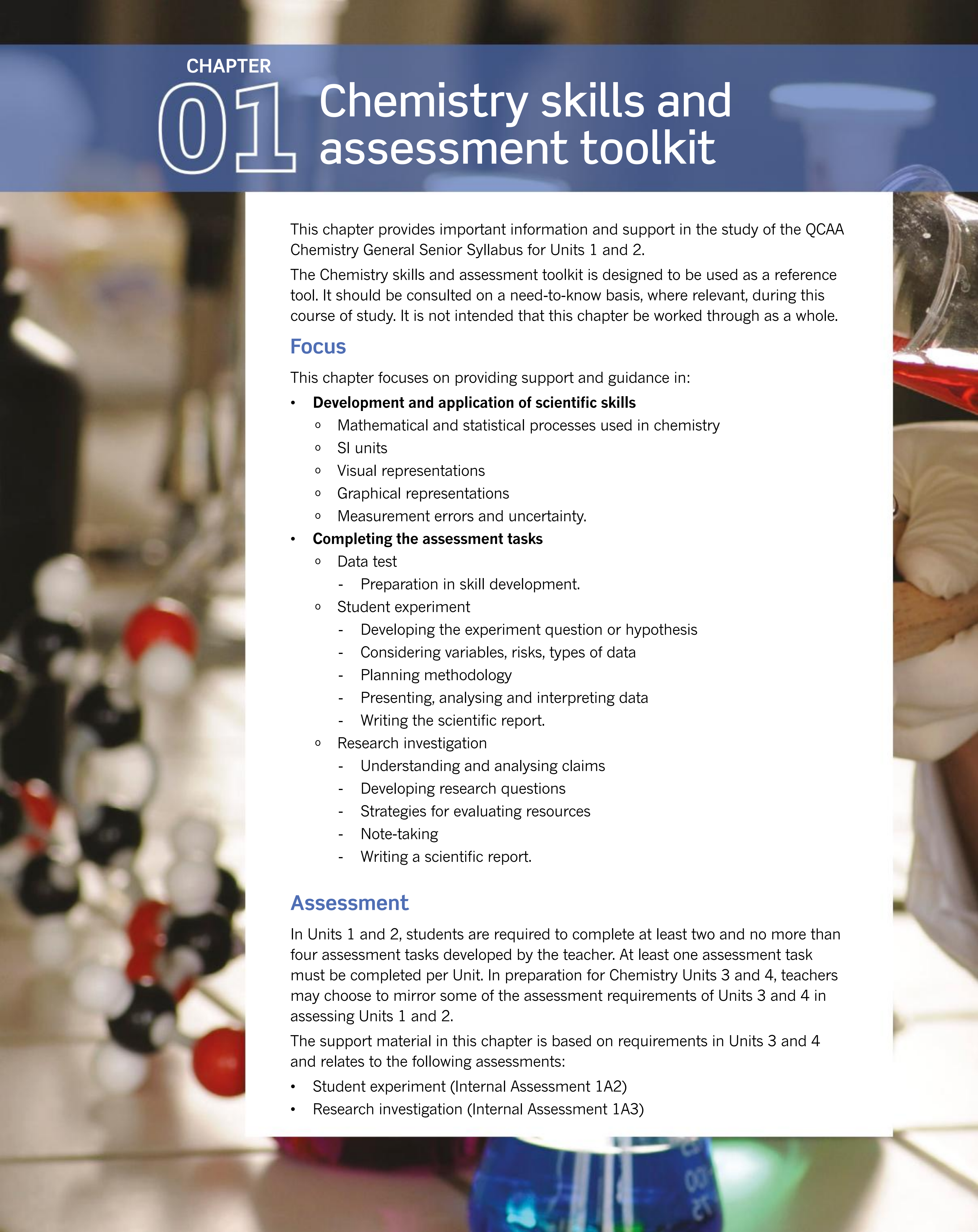
Analysis: **processes of matching, classifying, analysing errors, generalising, specifying** *(continued)*	infer	derive or conclude something from evidence and reasoning, rather than from explicit statements; listen or read beyond what has been literally expressed; imply or hint at
	interpret	use knowledge and understanding to recognise trends and draw conclusions from given information; make clear or explicit; elucidate or understand in a particular way; bring out the meaning of, e.g. a dramatic or musical work, by performance or execution; bring out the meaning of an artwork by artistic representation or performance; give one's own interpretation of; identify or draw meaning from, or give meaning to, information presented in various forms, such as words, symbols, pictures or graphs
	judge	form an opinion or conclusion about; apply both procedural and deliberative operations to make a determination
	organise/ sequence/ structure	arrange, order; form as or into a whole consisting of interdependent or coordinated parts, especially for harmonious or united action
	predict	give an expected result of an upcoming action or event; suggest what may happen based on available information
	reflect on	think about deeply and carefully
	scrutinise	to examine closely or critically (Macquarie 2015)
	sort	arrange in prescribed groupings or order

Increasing complexity of thinking

Level of thinking	Cognitive verb	Definition of cognitive verb
Knowledge utilisation: **processes of investigating, experimenting, decision-making, problem-solving**	adapt	modify or change something for a new purpose or use
	appraise	valuate the worth, significance or status of something; judge or consider a text or piece of work
	appreciate	recognise or make a judgement about the value or worth of something; understand fully; grasp the full implications of
	argue	give reasons for or against something; challenge or debate an issue or idea; persuade, prove or try to prove by giving reasons
	assess	measure, determine, evaluate, estimate or make a judgement about the value, quality, outcomes, results, size, significance, nature or extent of something
	comment	express an opinion, observation or reaction in speech or writing; give a judgement based on a given statement or result of a calculation
	conduct	direct an action or course; manage; organise; carry out
	conclude (conclusion)	a judgement based on evidence (ACARA 2015c)
	construct	create or pull together (e.g. an argument) by arranging ideas or items; display information in a diagrammatic or logical form; make; build
	convince (convincing)	persuade by argument or proof; leaving no margin of doubt; clear; capable of causing someone to believe that something is true or real; persuading or assuring by argument or evidence; appearing worthy of belief; credible or plausible
	create	bring something into being or existence; produce or evolve from one's own thought or imagination; reorganise or put elements together into a new pattern or structure or to form a coherent or functional whole
	decide	reach a resolution as a result of consideration; make a choice from a number of alternatives
	design	produce a plan, simulation, model or similar; plan, form or conceive in the mind; in English, select, organise and use particular elements in the process of text construction for particular purposes; these elements may be linguistic (words), visual (images), audio (sounds), gestural (body language), spatial (arrangement on the page or screen) and multimodal (a combination of more than one)
	determine	establish, conclude or ascertain after consideration, observation, investigation or calculation; decide or come to a resolution
	develop	elaborate, expand or enlarge in detail; add detail and fullness to; cause to become more complex or intricate

Increasing complexity of thinking

Knowledge utilisation: **processes of investigating, experimenting, decision-making, problem-solving** *(continued)*	devise	think out; plan; contrive; invent
	discuss/explore	examine by argument; sift the considerations for and against; debate; talk or write about a topic, including a range of arguments, factors or hypotheses; consider, taking into account different issues and ideas, points for and/or against, and supporting opinions or conclusions with evidence
	draw conclusions (conclusion)	a judgement based on evidence (ACARA 2015c)
	evaluate	make an appraisal by weighing up or assessing strengths, implications and limitations; make judgements about ideas, works, solutions or methods in relation to selected criteria; examine and determine the merit, value or significance of something, based on criteria
	experiment/test	try out or test new ideas or methods, especially in order to discover or prove something; undertake or perform a scientific procedure to test a hypothesis, make a discovery or demonstrate a known fact
	explore	inquire into something or discuss in detail
	generate	produce; create; bring into existence
	hypothesise	formulate a supposition to account for known facts or observed occurrences; conjecture, theorise, speculate, especially on uncertain or tentative grounds
	investigate	carry out an examination or formal inquiry in order to establish or obtain facts and reach new conclusions; search, inquire into, interpret and draw conclusions about data and information
	judge	form an opinion or conclusion about; apply both procedural and deliberative operations to make a determination
	justify	give reasons or evidence to support an answer, response or conclusion; show or prove how an argument, statement or conclusion is right or reasonable
	make decisions	select from available options; weigh up positives and negatives of each option and consider all the alternatives to arrive at a position
	manipulate	adapt or change to suit one's purpose
	modify	change the form or qualities of; make partial or minor changes to something
	persuade (persuasive)	capable of changing someone's ideas, opinions or beliefs; appearing worthy of approval or acceptance; (of an argument or statement); communicating reasonably or credibly
	predict	give an expected result of an upcoming action or event; suggest what may happen based on available information
	propose	put forward (e.g. a point of view, idea, argument, suggestion) for consideration or action
	prove	use a sequence of steps to obtain the required result in a formal way
	realise/resolve	in the Arts, consolidate and communicate intent through a synthesis of ideas and application of media to express meaning
	research	to locate, gather, record, attribute and analyse information in order to develop understanding (ACARA 2015c)
	solve	find an answer to, explanation for, or means of dealing with (e.g. a problem); work out the answer or solution to (e.g. a mathematical problem); obtain the answer/s using algebraic, numerical and/or graphical methods
	synthesise	combine different parts or elements
	test	take measures to check the quality, performance or reliability of something



CHAPTER 01

Chemistry skills and assessment toolkit

This chapter provides important information and support in the study of the QCAA Chemistry General Senior Syllabus for Units 1 and 2.

The Chemistry skills and assessment toolkit is designed to be used as a reference tool. It should be consulted on a need-to-know basis, where relevant, during this course of study. It is not intended that this chapter be worked through as a whole.

Focus

This chapter focuses on providing support and guidance in:

- **Development and application of scientific skills**
  - Mathematical and statistical processes used in chemistry
  - SI units
  - Visual representations
  - Graphical representations
  - Measurement errors and uncertainty.
- **Completing the assessment tasks**
  - Data test
    - Preparation in skill development.
  - Student experiment
    - Developing the experiment question or hypothesis
    - Considering variables, risks, types of data
    - Planning methodology
    - Presenting, analysing and interpreting data
    - Writing the scientific report.
  - Research investigation
    - Understanding and analysing claims
    - Developing research questions
    - Strategies for evaluating resources
    - Note-taking
    - Writing a scientific report.

Assessment

In Units 1 and 2, students are required to complete at least two and no more than four assessment tasks developed by the teacher. At least one assessment task must be completed per Unit. In preparation for Chemistry Units 3 and 4, teachers may choose to mirror some of the assessment requirements of Units 3 and 4 in assessing Units 1 and 2.

The support material in this chapter is based on requirements in Units 3 and 4 and relates to the following assessments:

- Student experiment (Internal Assessment 1A2)
- Research investigation (Internal Assessment 1A3)

Chapter organisation

The chapter is arranged in three parts:

- Part A: Working scientifically
- Part B: Student experiment
- Part C: Research investigation.

An outline of these parts is provided below.

QCAA Chemistry Syllabus objectives

- describe and explain scientific concepts, theories, models and systems and their limitations
- apply understanding of scientific concepts, theories, models and systems within their limitations
- analyse evidence
- interpret evidence
- investigate phenomena
- evaluate processes, claims and conclusions
- communicate understandings, findings, arguments and conclusions.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

Part A: Working scientifically

The focus of Part A is on basic mathematical skills and their applications. It features many Worked examples and opportunities to apply these examples yourself. These skills are engaged with throughout the course. They are therefore assessed, directly and indirectly, in assessment tasks and drawn on to analyse data in practical investigations. Explore Part A to help prepare yourself with the skills you will draw on for the Data test, and when undertaking a range of mandatory and suggested experiments. You will also find these skills useful when completing sections of the Student experiment, Research investigation and the Examination.

Refer to the following outline of Part A: Working scientifically, to learn, revise and practise the skills in the areas in which you need help.

Module	Look here for	eBook page
1.1 Chemical science	- areas of chemistry - the work of scientific research	e6
1.2 Orders of magnitude	- scientific notation - transforming decimal notations to scientific notation - converting values from decimal to scientific notation - converting between scientific notation and scientific units, orders of magnitude and prefixes - SI prefixes - examples of calculations	e11 e12
1.3 Mathematical basics for chemistry	- examples of maths used in chemistry - analysing straight line graphs	e17
1.4 Units	- measurement and units - correct use of unit symbols - SI prefixes - unit conversions	e20
1.5 Uncertainties in measurement and error	- explanations of the terms: uncertainty, error, accuracy, precision - causes of errors - calculations with uncertainty and error bars - writing measurements with uncertainty and error values - writing measurements and calculations to the correct significant figure - determining the absolute uncertainty of a value - converting between absolute and percentage uncertainties - calculating a value along with the associated uncertainty	e25 e28
1.6 Tables and graphing	- arranging and recording data in tables - converting data from tables to graphs - types of graphs—scatterplots, line graphs, bar and column graphs, pie charts - outliers and how they are represented on graphs - representing missing data on graphs	e39 e42
1.7 Statistics	- calculating mean, median, mode, range - statistical significance of standard deviation - regression, linear regression, coefficient of determination and *r*-values	e47

GO TO ➤ your R+ eBook to access Chapter 1 Chemistry Skills and Assessment Toolkit

Part B: Student experiment

The focus of Part B is on the student experiment. This internal assessment task requires students to follow the full scientific method over an extended and defined period of time. Students develop their own research question or hypothesis to investigate, based on an initial practical already completed.

Part B supports you through all aspects of the student experiment. The QCAA objectives and instrument-specific marking guide (ISMG) for the assessment are explained.

Explore Part B for examples of how to **modify** (**extend**, **refine** or **redirect)** the initial experiment to write an experiment research question or hypothesis. Be guided by the step-by-step guide to evaluate the quality of your hypothesis/question. Reinforce your knowledge and understanding of scientific methodology, delving into particular sections as needed. Refer to Part B for support on data types, data collection and analysis of data to draw valid conclusions. This includes how to identify errors in data, validity and relationships between data. Be guided in the write-up of your scientific report with support material on scientific writing style and the structure of the report.

Refer to the following outline of Part B: Student experiment, to learn, revise and practise your skills in the areas in which you need help.

Module	Look here for	eBook page
1.8 Research and planning	- types of variables - modify, refine, extend or redirect initial experiment - develop an experiment research question or hypothesis - evaluate the experiment research question or hypothesis - using a scientific journal to keep a record - validity and reliability - quantitative and qualitative data - nominal, ordinal, discrete, continuous data - risk assessments - chemical codes and symbols	e56 e58
1.9 Conducting and experimenting	- determine what data is relevant - determine what sufficient data is - collection of data - selecting appropriate equipment	e78
1.10 Results	- analysis of raw data - identifying errors; mistakes, systematic errors, random errors - analysing precision - analysing validity and theoretical relationships - interpreting results - relating findings to a chemical concept - evaluating data - evaluating the method	e83
1.11 Communicating and writing a scientific report	- the sections of a report - scientific writing style - write a scientific report - acknowledging sources - addressing the ISMG	e91

GO TO ➤ your R+ eBook to access Chapter 1 Chemistry Skills and Assessment Toolkit

Part C: Research investigation

The focus of Part C is on the research investigation. This internal assessment task requires students to gather secondary evidence on a research question over an extended and defined period of time. Students develop their own research question to investigate, based on a claim (provided by your teacher) related to the course.

Part C supports you through all aspects of the research investigation. The QCAA objectives and instrument-specific marking guide (ISMG) for the assessment are explained.

Engage with Part C to assist with writing your research question. Delve into the examples of claims, their analysis for context and elements, to guide you in developing your own research question. Be guided by the information about locating and evaluating suitable secondary sources for the research. This includes how to identify errors in data, validity and relationships between data. Be assisted in the writing of your scientific report with support material on the scientific writing style and the structure of the report. Part C provides a brief overview of different ways to present the report and provides details on the literature review format.

Refer to the following outline of Part C: Research investigation, to learn, revise and practise your skills in the areas in which you need help.

Module	Look here for	eBook page
1.12 Developing the research question from a claim	- analyse a claim - identify variables and measurable terms in the claim - examples of claims and research questions developed from them - guidelines for developing a question - develop a research question - refining a research question	e101
1.13 Finding and choosing suitable resources	- the difference between primary and secondary sources - locating resources - determine reliability and validity of resources	e107
1.14 Research: taking and organising notes	- recording notes in a scientific journal - paraphrasing information - different ways to record information - recording data and results - recording information about sources	e114
1.15 Writing a report for the research investigation	- different ways to present the report - presenting the report as a literature review - features of a literature review - structure of the literature review	e121

GO TO ➤ your R+ eBook to access Chapter 1 Chemistry Skills and Assessment Toolkit

1.1 Chemical science

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that chemistry is the study of materials and their properties and structure
- appreciate that chemistry uses mathematics
- explain the difference between scientific models, theories, laws, principles and facts
- understand that chemistry uses natural laws, theories, models and principles to describe real phenomena
- appreciate that a greater understanding of the theories and models that describe and explain the structure and properties of chemicals will help to make predictions and understand factors that affect and control chemical systems to produce desired products
- recall the major areas of chemistry
- explain how a vast number of professions use some aspect of chemistry
- appreciate that chemistry is a dynamic and evolving branch of science and helps to explain phenomena and solve problems encountered in an ever-changing world
- understand that science provides a systematic approach for comprehending the world.

Science is a human endeavour to understand the world around us and the known universe. It is one method of many used to develop knowledge. This knowledge is continually refined and redefined. Science follows a **methodology** to observe, test, measure, analyse and evaluate natural **phenomena** in a way humans can understand and construct meaning. There is no single **scientific method**; however, it is a cyclical process including observation of natural phenomena, questioning, hypothesising and predicting, experimentation and gathering **data**, refining and altering all previous elements, and theorising (Figure 1.1.1).

In the Chemistry course, students develop their understanding of chemical systems, and how **models** of matter and energy transfers and transformations can be used to describe, explain and predict chemical structures, properties and reactions.

FIGURE 1.1.1 A schematic outlining the processes involved in the scientific method. The scientific method is more of a cyclical process rather than linear.

Today, all scientific disciplines are expanding into new frontiers. The number of divisions and fields in chemistry is ever-increasing, and the relationship between disciplines is becoming more intertwined.

Innovations today are crossing traditional barriers between technology, chemistry, physics and biology of the past 100 or so years.

AREAS OF CHEMISTRY

Chemistry has been traditionally divided into four major areas—inorganic, organic, analytical and physical chemistry.

Organic chemistry deals with molecules formed predominantly from carbon, hydrogen, oxygen and nitrogen. The range and diversity of organic molecules is enormous—literally millions of molecules have been identified. Very long molecules can be made up of much smaller organic building blocks, forming polymers such as the plastics that have become so common over the past few decades. Because living organisms are made up primarily of organic molecules, organic chemistry forms the molecular foundation of the biological sciences and is intimately linked with pharmaceuticals and drug design.

Inorganic chemistry is the study of molecules composed of other elements and includes salts, metals, semiconductors and nanomaterials. Biologically important molecules, such as vitamins, haemoglobin and chlorophyll, contain metal atoms and organic sections.

Inorganic and organic chemistry often involve the design and synthesis of new materials with particular chemical and physical properties related to chemical reactivity, boiling point, strength, colour or magnetism.

Analytical chemistry involves the analysis and quantification of chemical components of a sample. This is of increasing importance surrounding the quality of manufactured goods, and in monitoring pollutant concentrations in our air, water and food. Even very low amounts of substances, such as mercury or dioxins, can be harmful to human health or the environment. Therefore, detecting and quantifying the amounts of these pollutants presents an ongoing intellectual and technical challenge.

Physical chemistry is the area of study of the physical properties, energies and spectra of all types of molecules, and the rate of chemical reactions. Physical chemists typically use spectroscopic or computational methods to study molecules, surfaces and chemical reactions. Sophisticated experimental techniques are required to probe complex or fast chemical processes, such as those that occur during combustion or the reactions taking place in the atmosphere.

Although traditional divisions in chemistry are still of vital importance, most of the exciting current scientific research is interdisciplinary and involves collaboration with many other scientific and technical fields including physics, biology, pharmacology, materials science, environmental science and engineering leading to developments in nanotechnology, materials, pharmaceutical chemistry and environmental research.

The role and work of a chemist

Those who study chemistry are typically known as chemists. However, as the world is changing, science, technology, engineering and maths (STEM) fields are integrating numerous disciplines into vocational roles. A recent report from a multinational business organisation indicates that jobs in as much as 75% of growing and future industries will require STEM qualifications.

A STEM qualification fosters skills in areas such as creativity and critical thinking, and is highly regarded in public and private businesses. According to this report, the safest future occupations are doctors, nurses and teachers. The Australian Government's chief scientist report states that of Australia's current STEM workforce, only 3.3% of STEM students graduated with a major in chemical sciences. There are numerous reports every year outlining that the majority of future professions will be STEM related.

The advances in scientific endeavours as well as innovations have seen a significant increase in multiple scientific specialities and disciplines working together. When more than one discipline is required to solve a problem or develop an innovation, it is said to be transdisciplinary. Transdisciplinary approaches to scientific endeavours form the heart of STEM.

The Royal Australian Chemical Institute identifies the jobs shown in Table 1.1.1 as being directly related to a chemistry qualification but there are many other jobs for which chemistry is a useful background subject.

TABLE 1.1.1 Jobs in chemistry

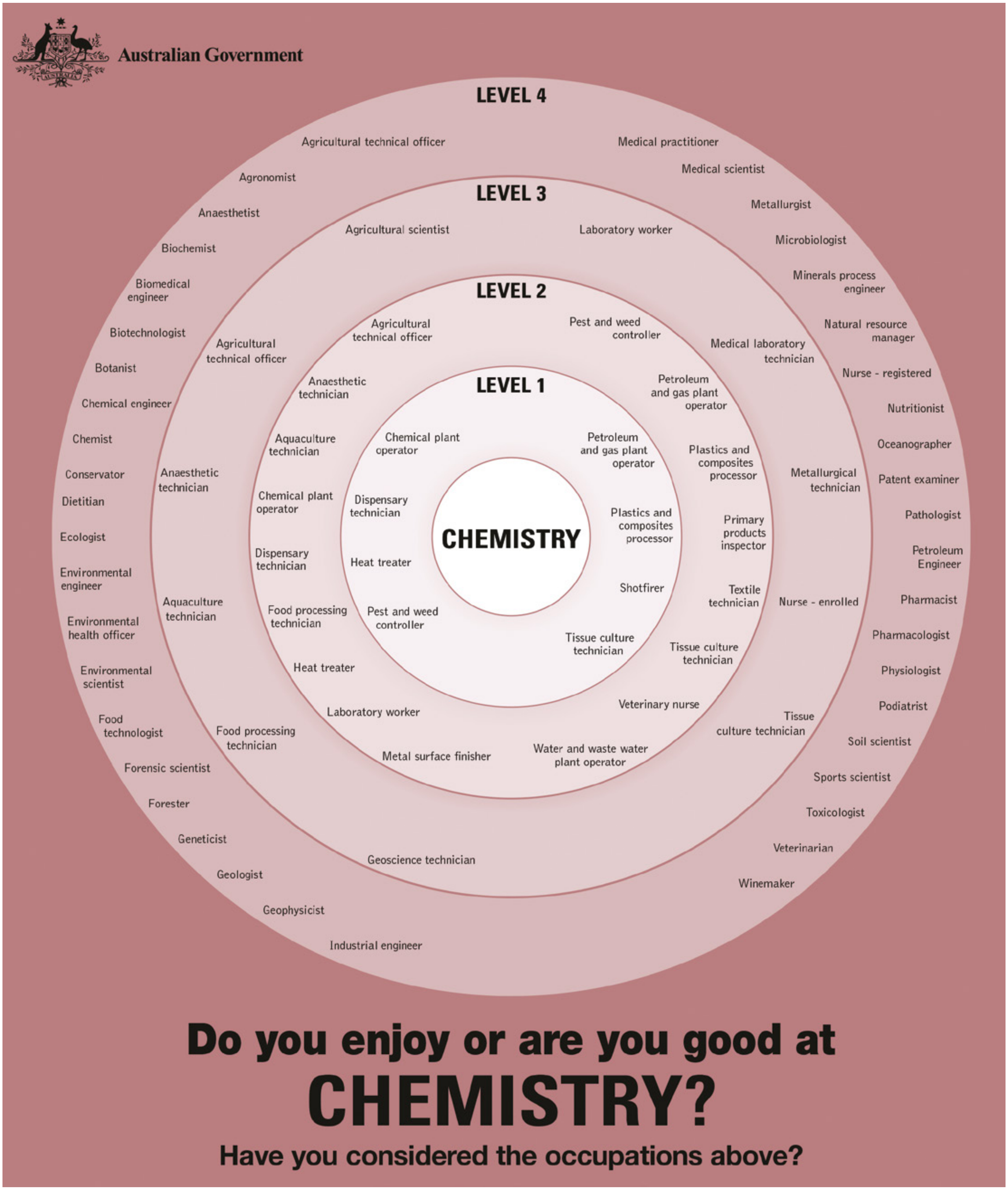
Job	Typical job description
analytical chemist	uses a diverse range of methods to investigate the chemical nature of substances, to identify and understand the substance and determine how it behaves in different conditions
clinical biochemist	analyses and interprets data relating to patients' samples to assist with the investigation, diagnosis and treatment of diseases
forensic scientist	provides impartial scientific evidence for use in courts of law to support the prosecution or defence in criminal and civil investigations
research scientist (physical sciences)	studies non-living systems to increase our understanding of how the physical world works; disciplines include astronomy, chemistry, earth science, materials science, meteorology and physics
chemical development engineer	involved in the development of industrial processes for the production of a diverse range of products, as well as in commodity and specialty chemicals
science writer	researches, writes and edits scientific news articles and features for business, trade and professional publications, specialist scientific and technical journals, and the general media
toxicologist	plans and carries out laboratory and field studies to identify, monitor and evaluate the impact of toxic materials and radiation on human and animal health, the environment, and the impact of future technology
process engineer	develops economical industrial processes to make the huge range of products on which modern society depends, including food and drink, fuel, artificial fibres, pharmaceuticals, chemicals, plastics, toiletries, energy and clean water

Table 1.1.2 lists jobs where a chemistry degree would be useful.

TABLE 1.1.2 A degree in chemistry would be beneficial in these jobs.

Job	Typical role description
patent agent	A patent agent obtains and enforces intellectual property rights on behalf of either individual inventors or organisations. Patents are granted by the government and give inventors the right to prevent the commercial use of inventions by other parties for a limited period. Chemists can specialise in pharmaceutical patent law.
environmental consultant	An environmental consultant works on commercial or government contracts to address a variety of environmental issues for their clients. This work covers a wide range of disciplines such as assessment of air, land and water contamination, environmental impact assessment, environmental audit, waste management, and the development of environmental policy and environmental management systems.
secondary school teacher or university lecturer	As well as teaching chemistry to students, lecturers, in particular, may also undertake academic or industry research and may be involved in school or departmental management.
chartered certified accountant	A chartered certified accountant is responsible for developing and maintaining financial and accounting systems, financial forecasting, auditing financial records and investigating financial anomalies. This field may appeal to chemists because of the numerical and analytical content of their degree.

A list of some of the careers related to chemistry can be seen in Figure 1.1.2.



Australian Government

LEVEL 4

Agricultural technical officer
Medical practitioner
Medical scientist
Metallurgist
Microbiologist
Minerals process engineer
Natural resource manager
Nurse - registered
Nutritionist
Oceanographer
Patent examiner
Pathologist
Petroleum Engineer
Pharmacist
Pharmacologist
Physiologist
Podiatrist
Soil scientist
Sports scientist
Toxicologist
Veterinarian
Winemaker
Industrial engineer
Geophysicist
Geologist
Geneticist
Forester
Forensic scientist
Food technologist
Environmental scientist
Environmental health officer
Environmental engineer
Ecologist
Dietitian
Conservator
Chemist
Chemical engineer
Botanist
Biotechnologist
Biomedical engineer
Biochemist
Anaesthetist
Agronomist

LEVEL 3

Agricultural scientist
Laboratory worker
Medical laboratory technician
Metallurgical technician
Nurse - enrolled
Tissue culture technician
Geoscience technician
Food processing technician
Aquaculture technician
Anaesthetic technician
Agricultural technical officer

LEVEL 2

Agricultural technical officer
Pest and weed controller
Petroleum and gas plant operator
Plastics and composites processor
Primary products inspector
Textile technician
Tissue culture technician
Veterinary nurse
Water and waste water plant operator
Metal surface finisher
Laboratory worker
Heat treater
Food processing technician
Dispensary technician
Chemical plant operator
Aquaculture technician
Anaesthetic technician

LEVEL 1

Chemical plant operator
Petroleum and gas plant operator
Plastics and composites processor
Shotfirer
Tissue culture technician
Pest and weed controller
Heat treater
Dispensary technician

CHEMISTRY

Do you enjoy or are you good at

CHEMISTRY?

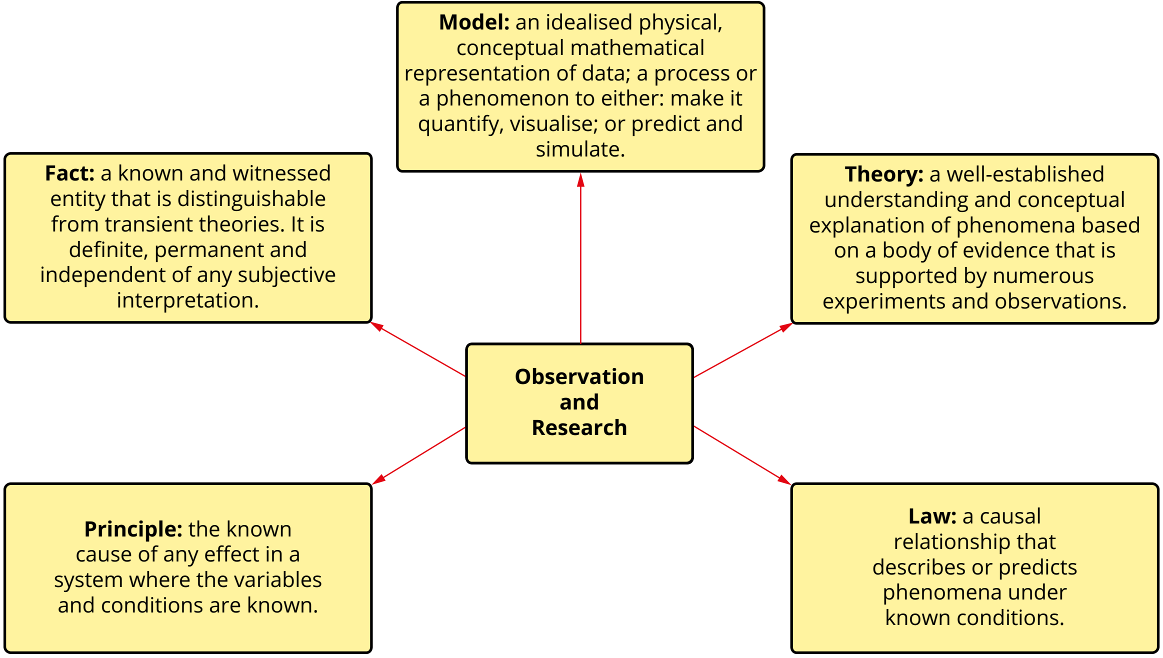
Have you considered the occupations above?

FIGURE 1.1.2 Examples of careers in the fields related to chemistry

THE WORK OF SCIENTIFIC RESEARCH

The foundation of all sciences, including chemistry, is scientific research. Scientific research works in the unknown to develop new understandings. To build evidence, it may take years and many experiments. In interpreting scientific evidence, scientists endeavour to establish models, **theories**, **laws**, **principles** and facts. These are illustrated in Figure 1.1.3.

It is important to note that all these deductions change over time as knowledge develops.



Model: an idealised physical, conceptual mathematical representation of data; a process or a phenomenon to either: make it quantify, visualise; or predict and simulate.

Fact: a known and witnessed entity that is distinguishable from transient theories. It is definite, permanent and independent of any subjective interpretation.

Theory: a well-established understanding and conceptual explanation of phenomena based on a body of evidence that is supported by numerous experiments and observations.

Observation and Research

Principle: the known cause of any effect in a system where the variables and conditions are known.

Law: a causal relationship that describes or predicts phenomena under known conditions.

FIGURE 1.1.3 Scientific research and observations are used to develop facts, principles, laws, theories and models.

1.1 Review

SUMMARY

- Science is known as a systematic approach to develop knowledge through testable explanations and repeatable means.
- Chemistry is the study of chemical systems, and how models of matter, energy transfer and transformations can be used to describe, explain and predict chemical reactions.
- There are four main areas of chemistry: organic, inorganic, physical and analytical.
- STEM (science, technology, engineering and mathematics) research is a transdisciplinary approach to scientific endeavours.
- Scientific evidence is used to formulate and establish models, theories, laws, principles and facts.
- A scientific model is an idealised representation of a process, a phenomenon or data.
- A scientific theory is a well-established understanding and conceptual explanation of phenomena based on a body of evidence.
- A scientific law is a causal relationship that describes or predicts phenomena under known conditions.
- A scientific principle is the known cause of any effect in a system where the variables and conditions are known.
- A scientific fact is a known and witnessed entity that is distinguishable from transient theories; facts are definite, permanent and independent of any subjective interpretation.

KEY QUESTIONS

Retrieval

1 Define chemistry.

Comprehension

2 Look at the careers model in Figure 1.1.2.
  a State two examples of careers in chemistry at level 4.
  b Identify a chemistry career in a growing STEM area.
  c Identify a chemistry career in a traditional area.

PART A WORKING SCIENTIFICALLY

Part A will consider the basic mathematical skills and applications that will be required for the Data test, a range of mandatory and suggested practicals as well as the student experiment.

1.2 Orders of magnitude

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- convert numbers to and from scientific notation
- estimate the value of a quantity using its order of magnitude
- estimate the value of a quantity of calculation.

SCIENTIFIC NOTATION

In scientific research and literature, measurements are often written in **scientific notation** to overcome confusion or ambiguity. Quantities are written as a number between one and ten and then multiplied by an appropriate power of ten. Note that 'scientific notation', 'standard notation' and 'standard form' all have the same meaning.

SKILLBUILDER 1.2.1

Transforming decimal notation to scientific notation

Scientists use scientific notation to handle very large and very small numbers. For example, instead of writing 0.000 000 035, scientists would write 3.5×10^{-8}.

A number in scientific notation (also called standard form or power of ten notation) is written as:

$a \times 10^n$

where:

- a is called the coefficient, and is a real number between 1 and 10, that is $1 \leq a < 10$
- n is called the exponent, and is an integer (a positive or negative whole number). n is the power that 10 is raised to.

Note that if n is positive, then the number is larger than one, and if n is negative, then the number is less than 1.

To transform a very large or very small number into scientific notation, perform these steps.

1. Write the original number as a decimal number greater than or equal to 1 but less than 10.
2. Multiply the decimal number by the appropriate power of 10.

The exponent is determined by counting the number of places the decimal point needs to be moved to form the original number again (Figure 1.2.1).

- If the decimal point is moved n places to the right, n will be a positive number; for example $510 = 5.1 \times 10^2$.
- If the decimal point is moved n places to the left, n will be a negative number; for example $0.051 = 5.1 \times 10^{-2}$.

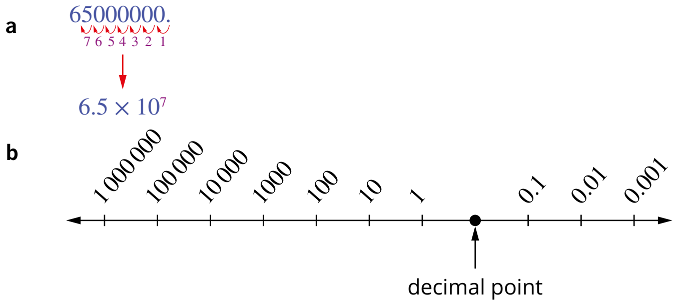


FIGURE 1.2.1 The exponential value is determined by counting the number of places the decimal point needs to be moved.

TABLE 1.2.1 Converting from scientific notation

Scientific notation	Number (expanded form)
1×10^{-9}	0.000 000 001
1×10^{-6}	0.000 001
1×10^{-3}	0.001
1×10^{-1}	0.1
1	1
1×10^{1}	10
1×10^{3}	1000
1×10^{6}	1 000 000
1×10^{9}	1 000 000 000

You will notice from Table 1.2.1 that when large numbers are written in scientific notation, the 10 has a positive exponential value. Very small numbers smaller than 1 are written with 10 to a negative exponent.

Test the skill

Convert 121 000 to scientific notation.

Examples of some values written in scientific notation are:
6.02×10^{23} particles
$25.25\,\text{mL} = 2.525 \times 10^{-2}\,\text{L}$
$0.00302\,\text{mol} = 3.02 \times 10^{-3}\,\text{mol}$

Values given in scientific notation show the significant numbers and indicate the precision of measurement during experimentation by the instrument. Uncertainties in measurement and error are covered in more detail in Module 1.5.

Worked example 1.2.1

CONVERTING TO SCIENTIFIC NOTATION

The radius of a sodium atom is 0.000 000 000 160 m. How many individual sodium atoms can fit across a typical pin head with a diameter of 1.5 mm?	
Thinking	**Working**
Calculate the diameter of a sodium atom (twice the radius).	0.000 000 000 320
Convert to scientific notation.	
Move the decimal place to the final digit.	 0.000 000 000 320 = 3.20
Count the number of places the decimal place was moved.	10
Use the number of decimal places moved as the power exponent to base 10.	3.20×10^{10}
Determine whether the decimal was moved left or right.	right
Place a negative symbol (–) in front of the power of 10.	$3.20 \times 10^{-10}\,\text{m}$
Convert 1.5 mm to m.	1.5 mm = 0.15 cm = 0.0015 m
Convert to scientific notation using the same steps as above.	$1.5 \times 10^{-3}\,\text{m}$
Now calculate the number of individual sodium atoms that can fit on a pin head by dividing the pin head diameter (1.5×10^{-3} m) by the sodium atom diameter (3.20×10^{-10} m).	number = $\frac{1.5\times10^{-3}}{3.20\times10^{-10}} = 4.7 \times 10^{6}$

➤ Try yourself 1.2.1

CONVERTING TO SCIENTIFIC NOTATION

The radius of a chloride ion is 181 pm. Calculate how many chloride ions could fit across a pin head with a diameter of 1.5 mm.

You should be routinely using scientific notation to express very large and very small numbers. This also involves learning to use your calculator intelligently. Scientific and graphics calculators can be put into a mode whereby all numbers are displayed in scientific notation. It is useful when doing calculations to use this mode rather than frequently attempting to convert to scientific notation by counting digits on the calculator display.

An important reason for using scientific notation is that it removes ambiguity about the precision of some measurements. For example, a measurement recorded as 240 g could be a measurement to the nearest gram; that is, somewhere between 239.5 g and 240.5 g, or it could have been rounded up to the nearest 10 g. By writing this quantity as 2.40×10^{2} g, it makes it clear the accuracy of the value (in this case, 3 **significant figures**).

Writing numbers using scientific notation

To write a number into long form from scientific notation, follow these rules.

- If the exponent is positive, write the coefficient without the decimal point. Continue to write zeros after the last digit in the coefficient so that the number of placeholders after the first digit is equal to the exponent.
- If the exponent is negative, write a zero in the units placeholder, then a decimal point. Continue to write zeros until the total number of zeros (including the zero in the units placeholder) is equal to the positive value of the exponent. Then write the coefficient without the decimal place.

Worked example 1.2.2

WRITING QUANTITIES FROM SCIENTIFIC NOTATION

The mass of the neutron is 1.675×10^{-27} kg. Write this value in long form.	
Thinking	**Working**
The exponent is negative, which means the number in long form will be less than 1. Write a zero in the units column and then a decimal point. After the decimal point, write another 26 zeros followed by the coefficient without the decimal place.	0.000 000 000 000 000 000 000 000 001 675 kg

➤ Try yourself 1.2.2

WRITING QUANTITIES FROM SCIENTIFIC NOTATION

Write the following values in long form.

a the diameter of a magnesium ion is 144 pm

b the number of hydrogen molecules in 1 mole of hydrogen gas is 6.02×10^{23}

Scientific notation is the preferred way of communicating very large and very small measurements and information in chemistry. Often for convenience and ease of use, quantities in chemistry problems should first be converted into scientific notation before any calculations can be carried out.

ORDERS OF MAGNITUDE AND PREFIXES

As stated above, scientists often work with very large or very small numbers because of the measurements made during experimentation. Physicists can measure astronomical distances between stars billions and billions of kilometres away; biologists measure the number of DNA bases in a genome in the millions or the size of biological molecules a millionth or a billionth the size of a metre, and chemists can measure the number of atoms in solution to the millions or billions of molecules. It is important to understand the numbers in scientific notation, their order of magnitude and how to convert them into other units.

Scientific notation is common for most sciences, displaying numbers with magnitude, for example 2.56×10^{-6}. The other form of communicating measured results is to use units that represent the order of magnitude. If 2.56×10^{-6} m was measured, it could be written either in the scientific notation as it is already, or as 256 μm. The symbol μ represents the prefix *micro* (Table 1.2.2), which means 'one millionth of the measured unit'. So when *micro* is placed in front of metre, it becomes micrometre, meaning 'one millionth of a metre', therefore 256 μm becomes 256 millionths of a metre. Again, these are very small measurements and numbers.

Each prefix in Table 1.2.2 represents an order of magnitude. You should become familiar with both scientific notation and prefixes for units of measurement. You should also be able to interchange them. Table 1.2.2 will help.

TABLE 1.2.2 The official SI prefixes used in chemistry

Prefix	Symbol	Multiplier	Description	Decimal	Example
tera	T	10^{12}	trillion	1 000 000 000 000	1 TByte hard drives are now common.
giga	G	10^{9}	billion	1 000 000 000	3.16 Gs = 1 century.
mega	M	10^{6}	million	1 000 000	1 MHz is close to the frequencies at which TV stations broadcast.
kilo	k	10^{3}	thousand	1000	1 kg is 1000 grams.
–	–	1	unit	1	–
deci	d	10^{-1}	tenth	0.1	1 L = 1 dm^3.
centi	c	10^{-2}	hundredth	0.01	1 cm is 1/100th of a m or 0.01 m.
milli	m	10^{-3}	thousandth	0.001	There are 1000 mm in 1 m.
micro	μ	10^{-6}	millionth	0.000 001	μg is commonly used as the unit for medicine dosage.
nano	n	10^{-9}	billionth	0.000 000 001	The wavelength of the yellow sodium streetlights is 550 nm.
pico	p	10^{-12}	trillionth	0.000 000 000 001	Blood concentration of insulin is measured in picomoles (pmol).

For example, when ionic radii are measured in X-ray crystallography, the length of each edge of the unit cell of sodium chloride is found to be 564.02 pm. Do not put spaces between the prefixes and the unit symbols. It is important to give a symbol the correct case (upper- or lowercase). There is a big difference between 1 mm and 1 Mm.

SKILLBUILDER 1.2.2

Converting between units

Often in science you may need to convert from one unit to another to complete a calculation. Knowing how to change between units is an important skill; however, conversion factors should be used carefully. As some units of measurement are difficult to visualise, knowing the size of different units in relation to each other will aid your understanding and help you avoid errors in your calculations. The most common mistake made with conversion factors is multiplying rather than dividing.

Cortisol is a natural hormone found in the body; however, it is used by some athletes to increase performance. Normal levels of cortisol in the blood are around 400 nmol L^{-1}. To put this into context, you can convert this value into a unit you can easily visualise, like grams per litre (g L^{-1}). To convert 40 nmol L^{-1} to mol L^{-1} (before converting to g L^{-1}), you need to move the decimal seven places to the left, so it becomes 0.000 000 4 mol L^{-1}. This can also be written as 4.0×10^{-7} mol L^{-1}. One mole of cortisol weighs 362.46 g, therefore 0.000 000 4 mol L^{-1} is equal to 0.000 145 g L^{-1} (or 1.45×10^{-4} g L^{-1})—very small indeed!

Worked example 1.2.3

CONVERTING BETWEEN UNITS

Convert 60 700 grams to kilograms.	
Thinking	**Working**
Consider the units from grams to kilograms.	$=\frac{\text{g}}{}\times\frac{\text{kg}}{\text{g}}$ $=\frac{\cancel{\text{g}}}{}\times\frac{\text{kg}}{\cancel{\text{g}}}$ $=\text{kg}$
To convert grams to kilograms, multiply 60 700 by kg per g. Since 1 kg = 1000 g, the kg is replaced by 1 and the g is replaced by 1000.	$=\frac{60700}{}\times\frac{\text{kg}}{\text{g}}$ $=\frac{60700}{}\times\frac{1}{1000}$ $=60.7\text{ kg}$

➤ Try yourself 1.2.3

CONVERTING BETWEEN UNITS

The rate of combustion of propane has been determined by measuring the production of CO_2 in mL. If CO_2 was generated at a rate of 25 μL s^{-1}, determine what this is in mL s^{-1}. Calculate how many litres of CO_2 would be produced in one hour.

When converting from one unit to another it is essential that you check that the answer makes sense. For instance, if converting from μm to km, you would expect that there are many μm in a km. The inversion of prefix conversions is one of the major sources of mistakes in general chemistry.

1.2 Review

SUMMARY

- 'Scientific notation', 'standard notation' and 'standard form' all have the same meaning.
- A number in scientific notation (also called standard form or power of ten notation) is written as $a \times 10^n$, where:
  - a is called the coefficient, and is a real number between 1 and 10; that is, $1 \leq a < 10$
  - n is called the exponent, and is an integer (positive or negative).
- A range of scientific symbols is used to represent 10^n.

KEY QUESTIONS

Retrieval

1 State why chemists need to write numbers in scientific notation instead of long form.

2 Identify the two numbers between which the coefficient of a number expressed in scientific notation has to lie.

3 State what the unit Tg means.

Comprehension

4 Represent the following values in scientific notation.

a 637 nm

b 319 s

5 Represent the following numbers in long form.

a 1.1×10^{-10} m

b 7.031×10^{6} ms^{-2}

6 Determine the order of magnitude of the following values.

a 6.41×10^{9} m

b 1.201×10^{-3} m^{3}

7 Match each number with its scientific notation.

Number	Scientific notation
0.002	2×10^{3}
2000	1.234×10^{-1}
0.1234	2×10^{-3}
123.4	1.234×10^{1}
12.34	1.234×10^{2}

Analysis

8 Convert the following lengths into the unit in brackets, presenting the information in scientific notation.

a 1.56 mm (μm)

b 8.35 nm (cm)

9 A human hair is about 50 μm wide, whereas a virus is 2×10^{-8} m wide. Determine how many times wider the width of a human hair is than the width of the virus.

10 Calculate the surface area, volume and surface area : volume ratio for the following shapes.

Shape	Surface area (cm^2)	Volume (cm^3)	Surface area : volume (cm^{-1})
cube (3 cm)			
sphere (radius 1.5 cm)			
cylinder (radius 1.2 cm; height 2 cm)			

1.3 Mathematical basics for chemistry

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that the rules of mathematics apply for solving mathematical operations or calculations in chemistry
- perform basic mathematical operations using percentages and fractions
- rearrange and substitute formulas using algebraic rules
- use the equation of a straight line to determine its gradient and *y*-intercept.

Science uses mathematics in working with data. The evidence can be quantified, measured and analysed. This is known as statistical analysis (a branch of mathematics).

When conducting mathematical operations or calculations, remember all the rules of mathematics apply, including:

- order of operations
- rules for transposing formulas and working with fractions
- substitutions
- approximations.

EXAMPLES OF MATHEMATICAL BASICS FOR CHEMISTRY

Worked example 1.3.1 demonstrates examples of mathematics in chemistry.

Worked example 1.3.1

MATHEMATICAL BASICS IN CHEMISTRY

The rate of the reaction between two chemicals is given by the expression:

rate = $k[A][B]$

where k is the rate constant.

Determine the rate when $k = 2.503\ (\text{mol dm}^{-3})^{-1}\ \text{s}^{-1}$, $[A] = 0.100\ \mu\text{mol L}^{-1}$ and $[B] = 180.0\ \mu\text{mol dm}^{-3}$.

Thinking	**Working**
Convert the units and prefixes into SI units and scientific notation (if required).	$[A] = 1.00 \times 10^{-4}\ \text{mol dm}^{-3}$ $[B] = 180.0 \times 10^{-6}\ \text{mol dm}^{-3}$ (or $1.800 \times 10^{-4}\ \text{mol dm}^{-3}$)
Write down the equation.	rate = $k[A][B]$
Substitute in the values.	rate = $2.503 \times 10^{-3} \times 1.00 \times 10^{-4} \times 1.800 \times 10^{-4}$ $= 4.5054 \times 10^{-11}\ \text{mol dm}^{-3}\ \text{s}^{-1}$
Express to the lowest number of significant figures in the data.	rate = $4.51 \times 10^{-11}\ \text{mol dm}^{-3}\ \text{s}^{-1}$

Try yourself 1.3.1

MATHEMATICAL BASICS IN CHEMISTRY

A hydrogen atom has a diameter of 64 pm. If the atom is represented as a sphere, calculate:

a the volume of a hydrogen atom

b the surface area of a hydrogen atom.

ANALYSING STRAIGHT LINES

A straight line is the simplest relationship between two variables. The **dependent variable** is plotted on the *y*-axis and the **independent variable** is on the *x*-axis then the relationship between the two variables is represented by the gradient. The *y*-intercept and gradient can be easily calculated by hand or with technology. Figure 1.3.1 shows a typical straight line relationship.

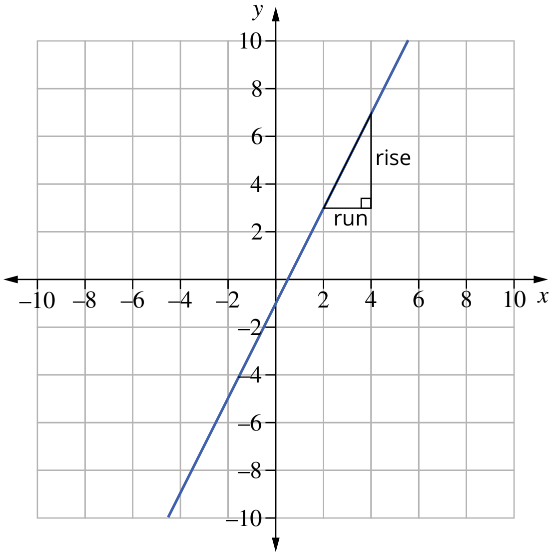


FIGURE 1.3.1 A graph representing a simple relationship between two variables

The gradient of any straight line is calculated using:

$$\text{gradient, } m = \frac{\text{rise}}{\text{run}} = \frac{\Delta y}{\Delta x}$$

The *y*-intercept, *c*, is read off the *y*-axis.

The equation of any straight line is determined by the gradient and *y*-intercept:

$$y = mx + c$$

In the example in Figure 1.3.1, the equation of the straight line is $y = 2x - 1$ with a gradient of 2 and *y*-intercept of –1.

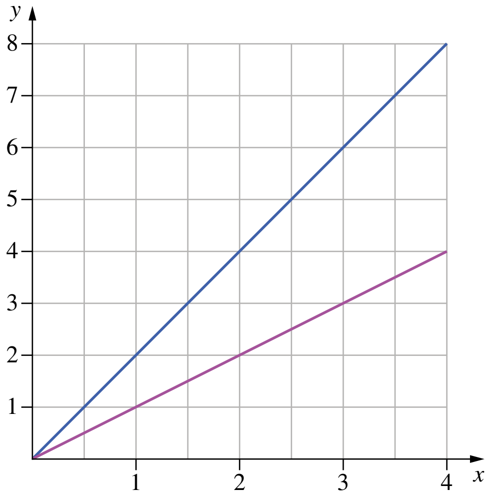
SKILLBUILDER 1.3.1

Interpreting the slope of a linear graph

Scientists often represent a relationship between two variables as a graph. For directly proportional relationships, the variables are related to each other by a straight line, where the slope (or gradient) of the line represents the constant of proportionality between the two variables.

The slope or gradient of the line is defined as the ratio of change between two points in the vertical axis (Δy), divided by the change between two points in the horizontal axis (Δx). In other words, it measures the rate at which one variable (the dependent variable) changes with respect to the other (the independent variable).

The graph shown has two straight lines with different slopes. The steeper slope (blue line) indicates that the rate of change is higher. This means the change is happening more quickly. On the other hand, the flatter slope (red line) indicates that the rate of change is lower. This means the change is happening more slowly.



Worked example 1.3.2

INTERPRETING THE SLOPE OF A LINEAR GRAPH

On a graph with two sloped lines, identify what the steeper sloped line indicates.

A a faster rate of change
B a slower rate of change
C the same rate of change
D a much slower rate of change

Thinking	Working
The slope or gradient of the line is defined as the ratio of change between two points in the vertical axis (Δy), divided by the change between two points in the horizontal axis (Δx).	A is correct. The steepness of the slope indicates the rate of change. A line with a steeper slope indicates a faster rate of change.

➤ Try yourself 1.3.2

INTERPRETING THE SLOPE OF A LINEAR GRAPH

Identify which of the following options indicates the rate of change of a straight line on a graph.

A area under the graph
B *y*-intercept
C *x*-intercept
D gradient

1.3 Review

SUMMARY

- When conducting mathematical operations or calculations, remember all the rules of mathematics apply, including:
  - order of operations
  - rules for transposing formulas and working with fractions
  - substitutions
  - approximations.

KEY QUESTIONS

Retrieval

1 State the formula for calculating the gradient of a straight line.

Comprehension

2 Determine these quantities.
  a the mass of 75 mL of ethanol (density of ethanol is $0.789\,\text{g}\,\text{mL}^{-1}$)
  b the volume of 24.5 g of ethanol

Analysis

3 A student recorded the following temperatures of water.

Time (min)	Temperature (°C)
0	22
2	34
4	43
5	50
8	64
10	78

  a Calculate the average temperature change per minute during the experiment.
  b Determine whether the increase was constant throughout the experiment.

1.4 Units

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- appreciate that chemistry uses the International System (SI) of units
- recall fundamental units, their symbols and definitions
- understand that the SI prefixes are multipliers of powers of ten of the base units
- convert between various units of the same type using SI and some non-SI units.

Science uses a specific system of units to describe measurements that are made. In chemistry, measurements are made using the **International System (SI)** of units. Most countries around the world use and understand these units as they are internationally accepted in science, commerce, sport and everyday life.

MEASUREMENT AND UNITS

When using measurements, scientists use conventional ways of writing numerical measurements, such as using SI units. Make sure the units you measure or observe are accurate and remember to use SI units when recording your evidence. Table 1.4.1 presents common quantities and SI units used in chemistry.

TABLE 1.4.1 Common quantities and SI units used in chemistry

Quantity	Symbol for quantity	Unit
mass	m	gram (g) kilogram (kg)
volume	V	litre (L) millilitre (mL)
amount of substance	n	mole (mol)
molar mass	M	grams per mole (g mol^{-1})
relative molecular or formula mass	M_r	no units; relative to one atom of ^{12}C exactly
relative atomic mass	A_r	no units; relative to one atom of ^{12}C exactly
relative isotopic mass	I_r	no units; relative to one atom of ^{12}C exactly
energy	E	joule (J) kilojoule (kJ)
concentration (molarity)	c	moles per litre (mol L^{-1}, M or mol dm^{-3})
density	d	g mL^{-1} or kg L^{-1}
temperature	T	°C, K
pressure	P	atm, Pa

Correct use of unit symbols

The correct use of unit symbols removes ambiguity, as symbols are recognised internationally. Some important points to keep in mind when writing units and recording measurements and numbers are given in Table 1.4.2.

Uppercase letters are only used for the *symbols* of the units that are named after people. For example, the unit of energy is joule and the symbol is J. The joule was named after James Joule who was famous for studies into energy conversions. The exception to this rule is ‘L’ for litre. Uppercase ‘L’ is used because a lowercase ‘l’ looks like the numeral ‘1’.

The product of a number of units is shown by separating the symbol for each unit with a space. The division or ratio of two or more units can be shown in fraction form, using a slash, or using negative indices. Prefixes should not be separated by a space.

TABLE 1.4.2 Examples of correct and incorrect use of symbols for derived units

Rule	✓	✗
Numbers and symbols should not be mixed.	thirty grams 30 g	thirty g 30 grams
Do not put spaces between prefixes and unit symbols.	kg mL kPa	k g m L k Pa
Write the correct case for symbols (upper- or lower-case).	1 mm 84 kg 5.2 mL	1 Mm 84 KG 5.2 m L
SI units should not be followed by a full stop unless they are at the end of a sentence. The symbols for units are not abbreviations.		
Only use upper-case letters for unit symbols named after people. The exception is ‘L’ for litre, to avoid confusion because ‘l’ looks like the numeral ‘1’.	J (the unit of energy is joule, named after James Joule)	
Units named after people can take the plural form by adding an ‘s’ when used with numbers greater than one. Never do this with the unit symbols.	two kilojoules	2 kJs
The product of a number of units is shown by separating the symbol for each unit with a space.	0.1 mol dm^{-3}	0.1 moldm^{-3}
The division or ratio of two or more units can be shown in fraction form, using a slash, or using negative indices.	0.1 mol dm^{-3} J g^{-1} °C^{-1}	0.1 moldm^{-3} Jg^{-1} °C^{-1}
When units are displayed with a negative exponent such as g mL^{-1}, the negative value of the exponent signifies the division of the units, or per unit. In this case, g mL^{-1} is the same as g/mL, or grams per decilitre.	1000 kJ mol^{-1} 13 °C/g	1000 kJ/mol^{-1} 13 °C/g^{-1}

Table 1.4.3 provides some examples of correct and incorrect use of symbols for derived units.

TABLE 1.4.3 Examples of the use of symbols for derived units

Correct symbol ✓	Incorrect symbol ✗
J g^{-1} °C^{-1}	Jg^{-1} °C^{-1}
g mol^{-1}	gmol^{-}1
kPa	K Pa

When units are displayed with a negative exponent, such as mol L^{-1}, the negative value of the exponent signifies the division of the units; in other words, per unit. In this case, mol L^{-1} is the same as mol/L, or moles per litre.

SI PREFIXES

In order to write down a physical measurement, two methods are used: prefixes and scientific notation. Scientific notation was discussed in Module 1.2.

An SI **prefix** is a symbol written at the front of a unit to indicate that the measurement is a power of ten higher or lower than the unit itself.

There are twenty SI prefixes in use and most of them denote units that are one thousand (10^3) times larger or smaller than the previous prefix. Table 1.4.4 lists the SI prefixes commonly used in chemistry.

TABLE 1.4.4 SI prefixes commonly used in chemistry

Prefix	Symbol	Multiplier	Description	Decimal	Example
tera	T	10^{12}	trillion	1 000 000 000 000	Electricity generation provides over 20 000 terawatt-hours of energy a year.
giga	G	10^{9}	billion	1 000 000 000	Over 10 gigatonnes of carbon dioxide is produced from burning coal.
mega	M	10^{6}	million	1 000 000	1 MHz is close to the frequencies at which TV stations broadcast.
kilo	k	10^{3}	thousand	1000	1 kg is 1000 g.
hecto	h	10^{2}	hundred	100	1 hPa is a unit used in meteorology used to describe atmospheric pressure.
–	–	1	unit	1	–
deci	d	10^{-1}	tenth	0.1	1 L = 1 dm^3.
centi	c	10^{-2}	hundredth	0.01	1 cm is the same as 1/100th of a m or 0.01 m.
milli	m	10^{-3}	thousandth	0.001	There are 1000 mm in 1 m.
micro	μ	10^{-6}	millionth	0.000 001	μg is commonly used as the unit for medicine dosage.
nano	n	10^{-9}	billionth	0.000 000 001	The wavelength of the yellow sodium streetlights is 550 nm.
pico	p	10^{-12}	trillionth	0.000 000 000 001	The diameter of the hydrogen atom is about 100 pm.
femto	f	10^{-15}	quadrillionth	0.000 000 000 000 001	Some lasers emit a pulse of light that lasts for 1 fs.

SI prefixes are used to avoid writing very small or very large numbers with lots of zeros. The prefix is added directly to the name of a unit, and a prefix symbol attaches directly to the symbol for a unit.

The larger prefixes are written as a capital letter, while the smaller prefixes are written using a lowercase letter.

Note that the fundamental unit kilogram has itself a prefix, but all multiples of mass are written with the base unit grams. As an example, one million kilograms is not written as one megakilogram, but as one gigagram (Gg).

Before any measurements with a prefix can be substituted into a formula, they must be converted into SI units without the prefix, i.e. in scientific notation.

Worked example 1.4.1

UNIT CONVERSIONS

Calculate how many litres are in one cubic kilometre.	
Thinking	**Working**
Write each unit in terms of m^3. Remember that a km^3 is a cube that has sides of 1000 m by 1000 m by 1000 m.	$1\ L = 0.001\ m^3$ and $1\ km^3 = 1000\ m \times 1000\ m \times 1000\ m = 10^9\ m^3$
Take the inverse of each equation to find the equivalent volume in m^3.	$1\ m^3 = \frac{1}{10^9} = 10^{-9}\ km^3$ and $1\ m^3 = \frac{1}{0.001} = 10^3\ L$
Now equate them since they are both equal to $1\ m^3$.	$10^{-9}\ km^3 = 10^3\ L$ $\therefore \frac{10^3\ L}{10^{-9}\ km^3} = 10^{12}\ L/km^3$
State the answer.	There are 10^{12} or one trillion litres in 1 cubic kilometre.

➤ Try yourself 1.4.1

UNIT CONVERSIONS

If a person weighs 6.0×10^1 kg and the mass of a single carbon atom is 1.99×10^{-23} g, calculate how many carbon atoms would be needed to equal the mass of the person.

1.4 Review

SUMMARY

- Scientists use conventional ways of writing numerical measurements such as using International System (SI) units.
- Units with a prefix need to be converted into scientific notation before they can be used in a chemistry formula.
- When converting units, it is very useful to write the conversions as fractions and then cross-cancel out to ensure that the conversion is done correctly.
- When units are displayed with a negative exponent, the negative value of the exponent signifies the division of the units. For example, g mL^{-1} is the same as g/mL or grams per millilitre.
- When writing and recording measurements, remember:
  - numbers and symbols should not be mixed
  - do not put spaces between prefixes and unit symbols
  - write the correct case for symbols (upper- or lowercase).

KEY QUESTIONS

Retrieval

1 State the standard unit of measurement, and the symbol, for the following quantities.

  a time
  b thermodynamic temperature
  c length
  d mass

2 List the following units from largest to smallest: centi, milli, deci, mega, kilo, nano, micro.

Comprehension

3 Explain why it is important that general society, not just the scientific community, needs a systematic and consistent set of units for measurement.

Analysis

4 Convert 30.00 mL into L.

5 Identify any errors in the recordings below, and correct them.

  a 23 mL/g h^{-1}
  b 79 Kl s^{-1}
  c four litres
  d 1.56 grams

6 Determine how many metres there are in the 656 nanometres wavelength of red light emitted from hydrogen.

7 Determine how many atoms there are in 5.3 g of helium.

1.5 Uncertainties in measurement and error

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain the difference between accuracy and precision
- explain the difference between uncertainty and error
- calculate uncertainty and error values for measurements
- write measurements with uncertainty and error values
- write measurements and calculations to the correct number of significant figures
- perform calculations that contain uncertainty or error values
- understand that all measurements made in chemistry experiments are subject to a level of uncertainty
- differentiate between accuracy and precision, and know how to express these quantitatively using significant figures
- perform calculations using significant figures
- explain the difference between mistakes, systematic uncertainties and random uncertainties
- determine the absolute uncertainty of a value based on the precision of the instrument
- convert between absolute and percentage uncertainties
- calculate a value along with the associated uncertainty.

All measurements are limited by the instrument used and the observer. Accuracy of the measurements and results are therefore affected. Instruments have a prescribed level of precision and cannot measure within the smallest increment of the device. In addition, a scientist's technique and experience with the instrument influences the precision with which the reading can be made.

However, in science it is impossible to measure the true value, as there will always be some degree of **uncertainty**. This results in **error** and uncertainty in the values measured or observed. Therefore, science can be thought of as developing evidence-based results that have some limitations. Measurements with uncertainty and errors are normal in science and so results are always presented with a **range** (e.g. ±0.5) indicating possible precision, or processed through statistical calculations.

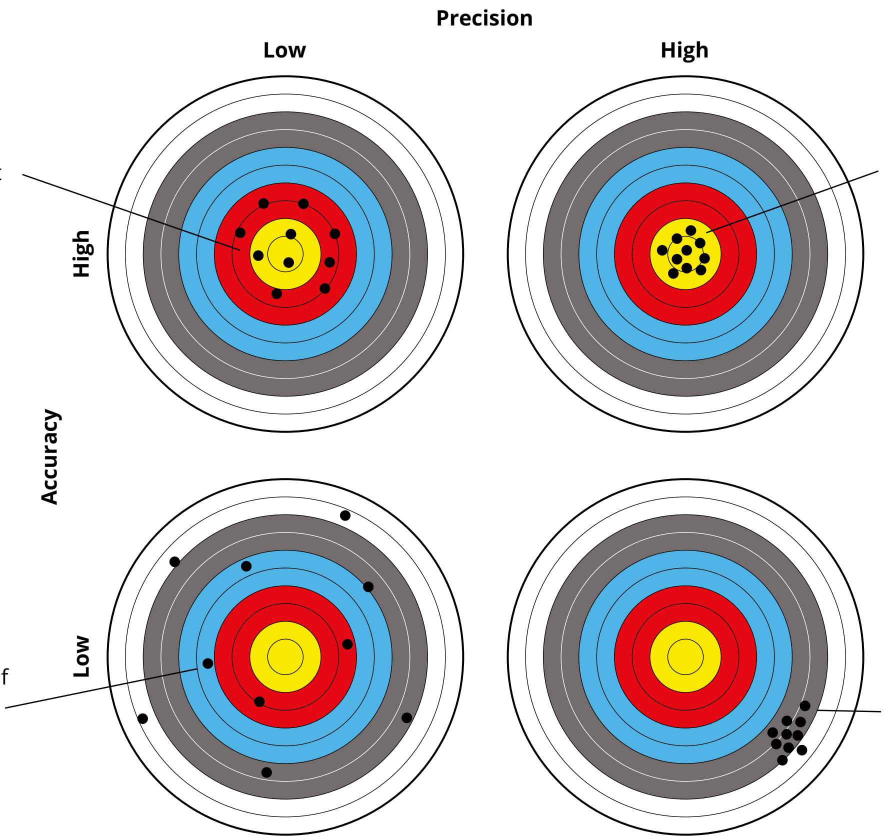
ACCURACY AND PRECISION

In science and statistics, the terms *accuracy* and *precision* have very specific and different meanings.

- **Accuracy** is the ability to obtain the correct measurement. To obtain accurate results, you must minimise **systematic errors**.
- **Precision** is the ability to consistently obtain the same measurement. To obtain precise results, you must minimise **random errors**.

To understand more clearly the difference between accuracy and precision, think about firing arrows at an archery target (Figure 1.5.1 on page e26). Accuracy is being able to hit the bullseye, whereas precision is being able to hit the same spot every time you shoot.

ℹ Any measurement made in chemistry needs to be both accurate and precise.



If you hit the area around the bullseye each time but do not always hit the bullseye, you are accurate

If you hit the bullseye every time you shoot, you are both accurate and precise

If you hit a different part of the target every time you shoot, you are neither accurate nor precise

If you hit the same area of the target every time but not the bullseye, you are precise but not accurate

FIGURE 1.5.1 Using a dartboard with darts illustrates the difference between accuracy and precision. In summary, measurements should be both precise and as accurate as you can make them.

Recording numerical data

When using instruments to measure data, the number of significant figures (or digits) and decimal places you use is determined by how precise your measurements are.

The overall precision of the measurement depends on the scale, accuracy and precision of the instrument and technique you are using (Figure 1.5.2). For example, a beaker may be used to measure volumes approximately and has limited accuracy of, for example, ±5%; however, a graduated pipette is much more accurate, with accuracies of ±0.1% or ±0.2%. Your pipette may be accurate but if your technique using the pipette is variable, the overall accuracy and precision will be limited.

When you record **raw data** and report processed data, you must use the number of significant figures available from your equipment or **observation**. (This is discussed in more detail in Module 1.10.) Using either a greater or smaller number of significant figures can be misleading. For example, Table 1.5.1 shows measurements of five mass samples taken using an electronic balance accurate to 2 decimal places. The data was entered into a spreadsheet to calculate the mean, which was displayed with 4 decimal places. You would record the mean as 20.83 g, not 20.8260 g, because 2 decimal places is the precision limit of the instrument. Recording 20.8260 g would be an example of false precision.

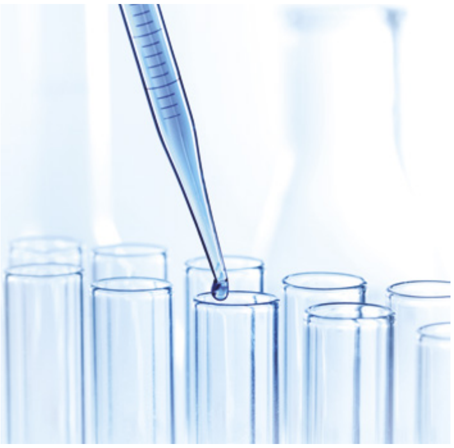


FIGURE 1.5.2 A 5 mL graduated pipette can measure volumes to an accuracy of one-tenth of a millilitre, or 5 mL ± 0.1 mL. The pipette has major divisions of 1 mL and minor divisions of 0.1 mL. You can estimate to 0.05 mL and record volumes to 2 decimal places, for example, 3.80 mL or 4.55 mL.

TABLE 1.5.1 An example of false precision in a data calculation

Sample	1	2	3	4	5	Mean
Mass (g)	20.13	20.62	21.22	20.99	21.17	20.8260

UNCERTAINTY AND ERROR

Uncertainty and errors arise from various different sources. Uncertainty is a range of measurements from the experiment, within which the true value is expected to lie. Due to variations, an estimation is made to suggest where the true value should be found among the measurements. Uncertainty is one indicator of a limitation.

Error is the difference between a measurement and the true value. This is not an indication of human error or mistakes. It is due to an inability to achieve the true value. Error cannot be eliminated, and therefore it is important to understand and record the error and report it. Two types of errors can occur—systematic error and random error (Table 1.5.2).

TABLE 1.5.2 Systemic and random errors

Error	Definition	Cause	Example
systematic	An error in measurements caused by the design of a system (e.g. methodology) or instrument (uncalibrated), which results in the measurements shifting in a systemic direction. The mean may be displaced or varied in a predictable way affecting the accuracy of the measurement.	• instrument calibration • choice of instrument • precision of instrument (increments) • precision of instrument (sample rate—per second) • sampling (the choice of population or representatives) • heat loss to the surroundings	The balance used for weighing out samples reads 0.05 g too high for all your masses (because it was improperly tared (set to zero) at the start of the experiment). Using a standard hydrochloric acid solution for a series of titrations that is not 1.0 M. As long as the actual concentration of the acidic solution is determined, the number of moles of base required to neutralise it in each titration can be corrected for.
random	An error that affects the measurement in an unpredictable way, resulting in an even variation (fluctuation) in measurements above and below the expected true value. This affects the precision of the measurement.	• estimation (measuring between increments) • sampling (frequency, spread, increments) • parallax • instrument sensitivity • noise	Using a universal indicator to determine the pH of a solution, in which the pH must be estimated by the shade and colour of the paper (e.g. yellow/orange versus bright orange). Measuring the time taken for a reaction to occur has a random error associated with the identification of when the reaction is judged to be complete.

Uncertainties and errors should always be calculated or measured and then reported. They then form part of the analysis and interpretation (see Module 1.7 Statistics and Module 1.10 Results) as well as the evaluation. Also, it is a good idea to consider uncertainty and error when planning experiments (see Module 1.8 Research and planning and Module 1.9 Conducting and experimenting later in this chapter).

Calculating uncertainty in measurement

The goal when taking measurements is to get as close as possible to the 'true' or 'correct' value. For example when using a measuring cylinder for the volume of solution, the reading may be between two increments on the instrument. The uncertainty can be calculated, or estimated, and is often represented as a percentage or expressed based on the unit of measurement. If the uncertainty was described as ±1%, it means that the measurement taken is likely to be no more than 1% above or below the 'true' value of the quantity that is being measured.

When averaging repeat measurements, the uncertainty should be reported alongside your average. **Uncertainty** represents a realistic range within which the true value is likely to be. The following equation is a simple way to calculate the uncertainty.

$$\text{uncertainty} = \pm \frac{\text{upper value} - \text{lowest value}}{2}$$

Worked example 1.5.1

CALCULATING UNCERTAINTY

A titration was performed three times and the volume of acid required to neutralise the solution was 20.8, 21.4 and 20.2 mL. The average volume of acid used was 20.8 mL. Calculate the uncertainty.	
Thinking	**Working**
Write down the formula.	$\text{uncertainty} = \pm \frac{\text{upper value} - \text{lowest value}}{2}$
Insert the values then calculate.	$\text{uncertainty} = \pm \frac{21.4 - 20.2}{2}$ $= \pm \frac{(1.2)}{2}$ $= 0.6\,\text{mL}$
Combine measurement with uncertainty.	20.8 mL ± 0.6 mL

➤ Try yourself 1.5.1

CALCULATING UNCERTAINTY

The rate of the reaction between sodium thiosulfate and sulfuric acid was measured by timing how long it took for a cross to disappear. The reaction was performed three times at each concentration of sulfuric acid and the average calculated. For the 1.0 M solution, the time for the cross to disappear was 51, 47 and 58 s. The average time was 52 s. Calculate the uncertainty.

Another example of uncertainty is instrumental error or uncertainty in measurement. This is the precision of an instrument and is limited by the smallest increment the instrument can measure. The uncertainty of an instrument is half the smallest increment. This is known as **absolute uncertainty**. The following equation is a simple way to calculate the absolute uncertainty of a measurement.

$$\text{absolute uncertainty} = \pm \frac{\text{smallest increment}}{2}$$

SKILLBUILDER 1.5.1

Calculating absolute uncertainty

Calculate the absolute uncertainty for instrumental error during an experiment where the temperature was measured at 32°C using a glass thermometer with increments of 1 degree Celsius and also a digital thermometer with increments of 0.1 degree Celsius.

For the glass thermometer

1 Write down the formula.

$$\text{uncertainty} = \pm \frac{\text{smallest increment}}{2}$$

2 Insert the values and calculate.

$$\text{uncertainty} = \pm \frac{1}{2}$$
$$= \pm 0.5°\text{C}$$

3 Combine the measurement with the uncertainty.

32°C ± 0.5°C

For the digital thermometer

1 Write down the formula.

uncertainty = smallest increment

2 Insert the values and calculate.

uncertainty = 0.1

3 Combine the measurement with the uncertainty.

32.0°C ± 0.1°C

Uncertainty can also be reported as a percentage of the measurement; this is known as **relative uncertainty**. To calculate the relative uncertainty, apply this formula:

$$\text{relative uncertainty} = \pm\left(\frac{\text{absolute uncertainty}}{\text{measurement}}\right)\times 100$$

Worked example 1.5.2

CALCULATING RELATIVE UNCERTAINTY FOR INSTRUMENTAL ERROR

The temperature during an experiment was measured and then reported at 32.00°C ± 0.05°C. Calculate the relative uncertainty.

Thinking	**Working**
Write the formula for relative uncertainty.	$\text{relative uncertainty} = \pm\left(\frac{\text{absolute uncertainty}}{\text{measurement}}\right)\times 100$
Insert the values and calculate.	$\text{relative uncertainty} = \pm\left(\frac{0.05}{32}\right)\times 100$ $= 0.16\ \%$
Combine the measurement with the relative uncertainty and express to the appropriate number of significant figures.	32.00°C ± 0.16%

➤ Try yourself 1.5.2

CALCULATING RELATIVE UNCERTAINTY FOR INSTRUMENTAL ERROR

The pH of a solution was determined to be 11.4 ± 0.6. Calculate the relative uncertainty.

Significant figures

When measuring phenomena during experimentation, the precision of the instruments determines the number of significant figures and decimal places you record in the data.

Significant figures convey meaning and precision and the number of significant figures used depends on the scale of the instrument. You should record data using the number of significant figures available from the equipment or observation. Using either a greater or smaller number of significant figures is misleading.

When writing an answer to the correct number of significant figures, you may need to use rounding if the initial answer has more figures (digits) than you need. These extra figures are called 'non-significant figures'. If the first non-significant figure is ≥5, round up the number. If the first non-significant figure is <5, then do not round up. For example, if the initial answer for a calculation was 2.1259 but you only needed an answer to 3 significant figures, you would round it to 2.13. If the initial answer was 2.1241, then to 3 significant figures, this becomes 2.12.

SKILLBUILDER 1.5.2

Identifying significant figures

When giving an answer to a calculation, it is important to take note of the number of significant figures that you use.

You should give an answer that is as accurate as possible. However, an answer cannot be more accurate than the data or the measuring device used to calculate it. For example, if an electronic balance that measures to the nearest gram shows that an object has a mass of 56 g, then the mass should be recorded as 56 g, not 56.0 g. This is because you do not know whether it is 56.0 g, or 56.1 g, or 56.2 g or 55.8 g.

The number 56 has 2 significant figures. Recording to 3 significant figures (e.g. 56.0 g or 55.8 g) would not be scientifically 'honest'.

If this mass of 56 g is used to calculate another value, it would also not be 'honest' to give an answer that has more than 2 significant figures.

Determining the number of significant figures to give an answer to depends on the kind of calculation you are doing.

- If you are multiplying or dividing, use the smallest number of significant figures provided in the initial values.
- If you are adding or subtracting, use the smallest number of decimal places provided in the initial values.

Working out the number of significant figures

Use the following rules to avoid confusion in determining how many significant figures are in a number.

1. All non-zero digits are always significant; for example, 21.7 has 3 significant figures.
2. All zeros between two non-zero digits are significant; for example, 3015 has 4 significant figures.
3. A zero to the right of a decimal point and following a non-zero digit is significant; for example, 0.5700 has 4 significant figures.
4. Any other zero is not significant, as it will be used only for locating decimal places; for example, 0.005 has just 1 significant figure.

Worked example 1.5.3

IDENTIFYING SIGNIFICANT FIGURES

a Identify which of the following is written to 2 significant figures.

A 30.1
B 0.000 40
C 0.5
D 5.12

b George multiplied 1.22 by 1.364. Identify which of the options below shows the result of this multiplication with the correct number of significant figures.

A 1.66
B 1.664
C 1.65
D 1.7

Thinking	Working
a The zeros that come before the integer 4 are not significant, whereas those that follow the integer are significant.	B is correct. A has 3 significant figures. C has 1 significant figure. D has 3 significant figures.
b 1.22×1.364 $= 1.66408$	A is correct. When two numbers are multiplied, use the smallest number of significant figures in the initial values to report your answer. In George's multiplication, the answer is 1.66408, but since 1.22 has 3 significant figures, the correct answer is 1.66.

➤ Try yourself 1.5.3

IDENTIFYING SIGNIFICANT FIGURES

Identify which of the following shows 41 written to 4 significant figures.

A 0.410
B 4.100
C 41.00
D 4100

Worked example 1.5.4

ACCURATELY RECORDING MEASUREMENTS FROM INSTRUMENTS

When measuring the amount of solution for an acid/base experiment, you need to record the results accurately and be able to identify the appropriate number of significant figures to use.

a State the measurement shown on this plastic pipette, and identify the number of significant figures in this result.

b State the measurement shown on this glass pipette, and identify the number of significant figures in this result.

Thinking	**Working**
a The increments shown on the plastic pipette determine the accuracy with which the result can be recorded. If you add extra numbers to the result it becomes inaccurate.	The plastic pipette is marked with increments of 0.05 mL, therefore all results must be recorded to the nearest 0.01 mL.
Determine where the meniscus (the lowest point) of the liquid falls.	The meniscus of the liquid in the pipette is closest to the 2.00 mL marking, therefore the result must be recorded as 2.00 mL.
Determine the number of significant figures included in this result.	The measurement of 2.00 mL has 3 significant figures.
b The increments shown on the glass pipette determine the accuracy with which the result can be recorded. If you add extra numbers to the result it becomes inaccurate.	The glass pipette is marked with increments of 0.5 mL, therefore all results must be recorded to the nearest 0.01 mL.
Determine where the meniscus (the lowest point) of the liquid falls.	The meniscus of the liquid in the pipette is closest to the 2.00 mL marking, therefore the result must be recorded as 2.00 mL.
Determine the number of significant figures included in this result.	The measurement of 2.00 mL has 3 significant figures.

➤ Try yourself 1.5.4

ACCURATELY RECORDING MEASUREMENTS FROM INSTRUMENTS

When measuring the temperature for a calorimeter experiment, you need to record the results accurately and be able to identify the appropriate number of significant figures to use.

a State the measurement shown on this glass alcohol thermometer, and identify the number of significant figures in this result.

b State the measurement shown on this digital thermometer, and identify the number of significant figures in this result.

Significant figures are the exact digits observed when recording the results of an observation or experiment. For example, if the volume of O_2 gas produced and measured from the electrolysis of water in a 20-minute experiment was 6.0 mL because exactly 6.0 mL was observed, then the measurement is 6.0 mL with 2 significant figures, not 6 mL.

Table 1.5.3 shows some examples of measurements and how many significant figures they have.

TABLE 1.5.3 Measurements and significant figures

Measurement	Number of significant figures
15	2
3.5	2
3.50	3
0.0**37**	2
1401	4
1400	2

ℹ It is important to record the exact numbers measured during an experiment. This determines the significant figures, and later the accuracy and precision of the analysis. If a zero is observed, record it, as it is the exact amount measured. This includes all zeros before and after decimal places.

Using either a greater or smaller number of significant figures can be misleading. For example, Table 1.5.4 shows three measurements to 2 decimal places (4 significant figures). The data was entered into a spreadsheet to calculate the **mean**, which was displayed with 3 decimal places. You would record the mean as 19.52 mL, not 19.517 mL, because 2 decimal places is the precision limit of the instrument. Recording 19.517 g would be an example of false precision.

TABLE 1.5.4 An example of false precision in a data calculation

Trial	1	2	3	Mean
Volume (mL)	19.45	19.55	19.55	19.517

Calculations with uncertainties

There are a few rules when conducting calculations with numbers (or measurements) that have uncertainties. It is important to record the correct uncertainty once all the calculations are done so that the appropriate accuracy is reported. Table 1.5.5 outlines the rules for undertaking calculations with data and its associated uncertainty.

Keep in mind that the rules for significant figures apply to all numerical reporting, including the uncertainties.

TABLE 1.5.5 Rules for calculations involving uncertainty

Calculation	Rule	Examples
addition and subtraction	add the absolute uncertainties	$5.4 \pm 0.05 + 3.2 \pm 0.05 = 8.6 \pm 0.1$ $5.4 \pm 0.05 + 3.2 \pm 0.1 = 8.6 \pm 0.15$ $5.4 \pm 0.05 - 3.2 \pm 0.05 = 2.2 \pm 0.1$ $5.4 \pm 0.05 - 3.2 \pm 0.1 = 8.6 \pm 0.15$
multiplication and division	add the relative uncertainties	$3.0 \pm 0.05 \times 1.5 \pm 0.05$ $= 3.0 \pm 1.6\% \times 1.5 \pm 5\% = 4.5 \pm 6.6\%$ $3.0 \pm 0.05 \div 1.5 \pm 0.05$ $= 3.0 \pm 1.6\% \div 1.5 \pm 5\% = 2.0 \pm 6.6\%$
multiplying by a constant (absolute uncertainty)	multiply the uncertainty	$3.14 \times (4.1 \pm 0.05) = 12.9 \pm 0.16$
multiplying by a constant (relative uncertainty)	maintain the uncertainty	$3.14 \times (4.1 \pm 1.6\%) = 12.9 \pm 1.6\%$

Determining an uncertainty in a single measurement

When using an instrument to measure a property of an object, there will always be a limit as to how precise the reading can be.

The precision that can be read off a single instrument is called the limit of reading. This is the smallest visible graduation on an analogue instrument, or if a digital instrument is used, then the limit of reading is usually the smallest decimal place given on the screen.

For most analog instruments, the **absolute uncertainty**, Δx, in a measurement of x is exactly half of the limit of reading for each part of the measured object that has to be lined up. This means that the actual measurement could be anywhere from *half* of the smallest graduation too big to *half* of the smallest graduation too small.

Some analog instruments have a zero mark and a reading mark, such as a ruler shown in Figure 1.5.3a. The object to be measured has to be lined up at *both* ends, thus doubling the absolute uncertainty. For instance, when using rulers or burettes, the absolute uncertainty is equal to the limit of reading, rather than half of it.

In a digital instrument, Δx is typically the limit of reading, or the smallest decimal place given on the screen. An example of this is a high-precision electronic balance as shown in Figure 1.5.3b.

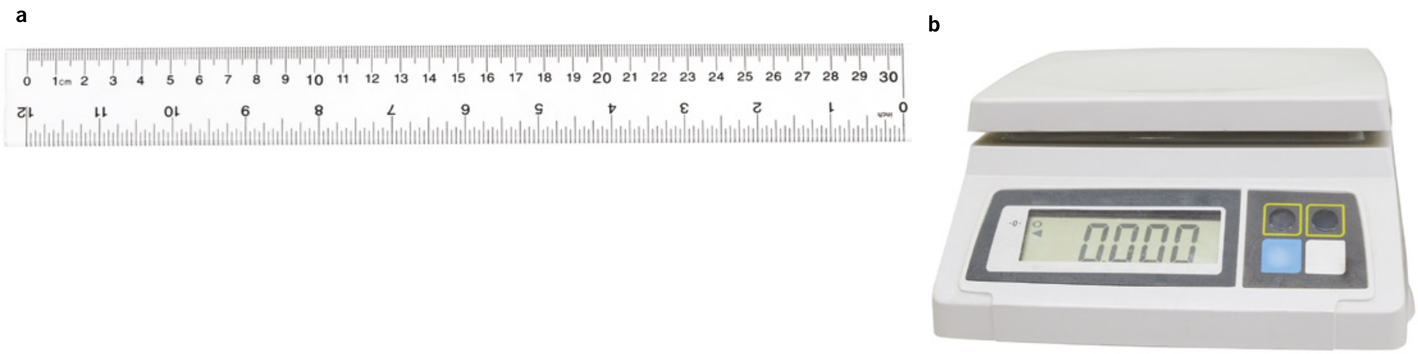


FIGURE 1.5.3 (a) a ruler with a limit of reading of 1 millimetre, (b) an electronic balance with a limit of reading of 0.001 g

For analog instruments that have one scale to line up (such as a measuring cylinder, thermometer or voltmeter), the absolute uncertainty is half of the smallest graduation scale, or limit of reading.

For analog instruments that have two scales to line up (such as rulers and burettes), the absolute uncertainty is equal to the smallest graduation scale, or limit of reading.

For digital instruments, the absolute uncertainty is equal to the smallest decimal place displayed.

Usually, the absolute uncertainty in a measurement will have one significant figure, or two significant figures if the first digit is 1. This is done because the uncertainty of a measurement is meant to be an estimate and writing more than two significant figures is therefore meaningless.

When the absolute uncertainty has been determined, the measurement should be rounded so that it and the absolute uncertainty have the same number of *decimal places*. This is to keep the measurement and its uncertainty to the same precision. Examples of measurements and their absolute uncertainties are shown in Table 1.5.6.

TABLE 1.5.6 Examples of writing a measurement and its absolute uncertainty

Measurement and absolute uncertainty	Comments
48.2 ± 1.0°C	The absolute uncertainty is written to 2 significant figures because the first digit is 1. The measurement and uncertainty then have the same number of decimal places.
99.99 ± 0.01 s	The measurement and absolute uncertainty are written to 2 decimal places, with one significant figure for the uncertainty. This measurement of time may have been recorded using a digital device, such as a data logger, since the precision is quite high.
0.000000031 ± 0.000000005 kg	This is technically a correct way of writing a measurement and its absolute uncertainty, but it is normally written in scientific notation so that the measurement and uncertainty have the same power of ten, i.e. $(3.1 \pm 0.5) \times 10^{-8}$ kg

The limit of reading and absolute uncertainties of common equipment used in a chemistry laboratory are shown in Table 1.5.7.

TABLE 1.5.7 A summary of common instruments and their uncertainties

Instrument	Type of instrument	Limit of reading	Absolute uncertainty	Example	Notes
mm ruler	analog	1 mm	± 1.0 mm	123.8 ± 1.0 mm	Half of the limit of reading is 0.50 mm, but since the object to be measured has to be lined up at both ends, the uncertainty is double this, or ±1.0 mm.
cm ruler	analog	1 cm	± 1.0 cm	31.4 ± 1.0 cm	Half of the limit of reading is 0.50 cm, but since the object to be measured has to be lined up at both ends, the uncertainty is double this, or ±1.0 cm.
thermometer	analog	1.0 °C	± 0.5 °C	24.4 ± 0.5 °C	Only the top part of the liquid in the thermometer has an uncertainty, so the uncertainty is exactly half of the limit of reading.
measuring cylinder	analog	1.0 mL	± 0.5 mL	15.7 ± 0.5 mL	The limit of reading and absolute uncertainty depend on the size of the measuring cylinder.
burette	analog	0.1 mL	± 0.05 mL	20.10 ± 0.05 mL	Half of the limit of reading is 0.05 mL, but since the object to be measured has to be lined up at both ends, the uncertainty is double this, or ±0.1 mL.
volumetric pipette	analog	varies	varies ± 0.03 mL	25.00 ± 0.03 mL	The uncertainty for the pipette is based on the precision given on the pipette. The precision varies based on the size and quality of the pipette.
stopwatch	digital	0.01 s	from ± 0.1 to ± 0.4 s	9.6 ± 0.2 s	The limit of reading is not the absolute uncertainty for a stopwatch because the user's reaction time (usually between 0.1 and 0.4 s) needs to be taken into consideration.
electronic balance	digital	typically 0.01 g	± 0.01 g	7.03 ± 0.01 g	The limit of reading is usually the smallest decimal place given on the screen.

These steps will help you write a measurement and its absolute uncertainty.

1 Determine the absolute uncertainty from the type of instrument used to measure the quantity.
2 Measure the quantity, including an estimate of the final significant figure.
3 If the first digit of the absolute uncertainty is 1, then one more significant figure can be quoted. Usually this extra digit is zero.
4 The measurement is then rounded up or down to the same number of decimal places, or place value holders, as the uncertainty; however, the uncertainty should not be presented to more than 2 significant figures.
5 Do not forget the unit of measurement for the quantity.

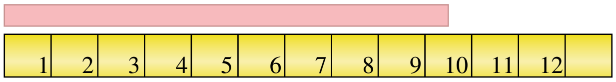
The absolute uncertainty should be written to 1 or a maximum of 2 significant figures.

Worked example 1.5.5

READING MEASUREMENTS

For each of the following objects, determine its measurement and uncertainty.

a the length of the pink rectangle measured with a ruler calibrated in centimetres



b the temperature of tap water measured using an analog thermometer

c the temperature of tap water measured using a digital thermometer

Thinking	Working
a The pink rectangle is between 9 cm and 10 cm in length, so the best estimate for this is 9.6 cm.	length, L = 9.6 cm
The limit of reading is 1.0 cm, so this is the uncertainty in a centimetre ruler.	$\Delta L = \pm 1.0$ cm
Write the measurement and uncertainty.	length = 9.6 ±1.0 cm
b The temperature is between 20°C and 21°C, with the best estimate of 20.2°C.	temperature, T = 20.2°C

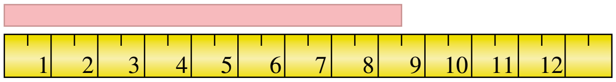
The limit of reading is 1.0°C, so the uncertainty is half of this, i.e. 0.5°C.	$\Delta T = \pm 0.5$°C
Write the measurement and uncertainty.	temperature = 20.2 ± 0.5°C
c This temperature measurement is done using a digital thermometer, so the temperature is simply read off the screen at 31.2°C.	temperature, $T = 31.2$°C
The uncertainty in this measurement is then 1 of the smallest decimal value, i.e. ±0.1°C.	$\Delta T = \pm 0.1$°C
Write the measurement and uncertainty.	temperature = 31.2 ± 0.1°C

➤ Try yourself 1.5.5

READING MEASUREMENTS

For each of the following objects, determine its measurement and uncertainty.

a the length of the pink rectangle measured with a ruler calibrated in centimetres



b a red liquid in an analog measuring cylinder

c the mass of this doughnut measured on an electronic balance

1.5 Review

SUMMARY

- Accuracy is the ability to obtain the correct measurement. To obtain accurate results, you must minimise systematic errors.
- Precision is the ability to consistently obtain the same measurement. To obtain precise results, you must minimise random errors.
- Uncertainty is a range of measurements from the experiment in which the true value is thought to be.
- Error is the difference between a measurement and the true value; this is *not* an indication of human error or a mistake.
- Uncertainty and error values are presented as a range (e.g. ±0.3).
- A simple way to calculate the uncertainty of a range of data is:

  $$\text{uncertainty} = \pm\frac{(\text{upper value} - \text{lowest value})}{2}$$

- A simple way to calculate the absolute uncertainty of an analogue measurement is:

  $$\text{absolute uncertainty} = \pm\frac{\text{smallest increment}}{2}$$

- The relative uncertainty can be calculated using this formula:

  $$\text{relative uncertainty} = \pm\left(\frac{\text{absolute uncertainty}}{\text{measurement}}\right)\times 100$$

- The rules for calculations involving uncertainty are listed in the following table.

Calculation	Rule
addition and subtraction	add the uncertainties
multiplication and division	add the relative uncertainties
multiplying with a constant (absolute uncertainty)	multiply the uncertainty
multiplying with a constant (relative uncertainty)	maintain the uncertainty

- Significant figures are a way of communicating how precise a measurement is.
- The number of significant figures used depends on the scale of the instrument.
- There are specific rules governing how many significant figures there can be in any measurement.
  - All non-zero digits are significant.
  - All zeros between non-zero digits are significant.
  - Trailing zeros (either decimal or not) if measured are significant.
  - Leading zeros are not significant (in numbers with decimals, they only inform where the decimal is placed).
- All measurements made in chemistry experiments are subject to a level of uncertainty or doubt to the precision of that measurement.
- A measurement without an uncertainty is meaningless.
- Accuracy is a description of how close a measurement is to an accepted value.
- Precision is a description of how many digits can be written down for a single measurement, or how close together a group of measurements are.
- Systematic errors arise due to mistakes in calibrating or using equipment the wrong way for every measurement.
- Random errors are naturally occurring and cannot be eliminated.
- The limit of reading of an instrument is the smallest graduation on the device.
- The absolute uncertainty of a measurement is determined by the limit of reading of an instrument.
- Uncertainties are typically written as an absolute value or percentage.
- Propagation of uncertainties involves the mathematical combination of uncertainties from a range of variables.
- When adding or subtracting measurements, the individual absolute uncertainties always add.
- When multiplying or dividing measurements, the individual percentage uncertainties always add.
- When raising a measurement to a power, the percentage uncertainty is multiplied by the absolute value of the power.
- When using measurements in a formula, perform any addition and subtraction propagations first, followed by multiplication, division and indices.

1.5 Review *continued*

KEY QUESTIONS

Retrieval

1 Describe how values of uncertainty and error are presented.

2 State the absolute uncertainty of the following instruments.

- **a** a cm ruler
- **b** a protractor graduated in degree intervals
- **c** an electronic balance giving a reading to the nearest 0.001 g
- **d** an analog needle voltmeter graduated in 10 volt intervals

Comprehension

3 Explain the difference between accuracy and precision.

4 Explain why it would be unreasonable to state that the absolute uncertainty in a hand-held stopwatch is ± half of the limit of reading or ± 0.005 s.

Analysis

5 Calculate the average and uncertainty of the two sets of data below.

Data set A: 11.4, 10.9, 11.8, 10.6, 11.5, 11.1
Data set B: 25.3, 27.4, 22.0, 26.1, 28.4, 23.1

6 Determine the absolute uncertainty of the following instruments.

- **a** a garden pH meter graduated in 0.2 intervals
- **b** a burette graduated in 0.1 mL intervals
- **c** an electronic balance giving a reading to the nearest 0.001 g

1.6 Tables and graphing

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- create tables to record and organise data
- present data from tables into graphs
- distinguish between types of graphs
- select appropriate graphs to represent data
- represent missing data in graphs
- explain what an outlier is, and know how to represent them in graphs.

PRESENTING DATA IN TABLES

Tables record number values and allow you to organise your data. In general, tables provide more detailed data than graphs, but it is easier to observe trends and patterns in data when they are in graph form than in table form.

Presenting raw data in tables

Tables organise data into rows and columns, and can vary in complexity according to the nature of your data. Tables can be used to organise raw data and **processed data**, or to summarise results.

The simplest form of a table is a two-column chart. The first column should contain the independent variable (the one you control) and the second column should contain the dependent variable (the one that may change in response to a change in the independent variable).

As you can see in Figure 1.6.1, tables should have the following features:

- a descriptive title
- column headings (including the units)
- aligned figures (align the decimal points)
- the independent variable placed in the left column.

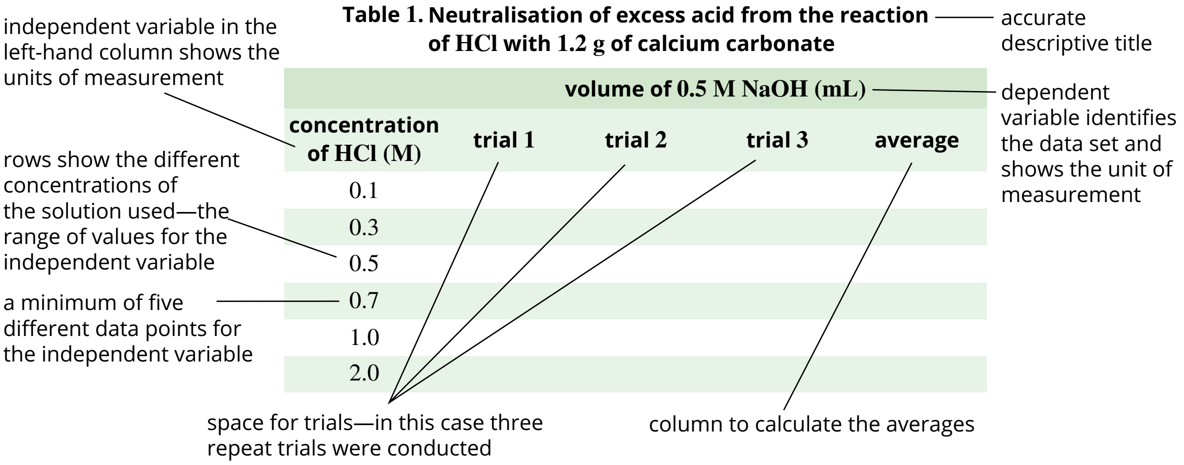


Table 1. Neutralisation of excess acid from the reaction of HCl with 1.2 g of calcium carbonate

	volume of 0.5 M NaOH (mL)			
concentration of HCl (M)	**trial 1**	**trial 2**	**trial 3**	**average**
0.1				
0.3				
0.5				
0.7				
1.0				
2.0				

FIGURE 1.6.1 Features of a good table

Presenting processed data in tables

Table 1.6.1 shows the relationship between temperature and average reaction rate. It displays reaction rate data in a processed format, because it displays the mean of several values for each temperature. To report processed data accurately the uncertainty is also displayed, which will be a requirement for assessments.

TABLE 1.6.1 Effect of temperature on mean reaction rate, including uncertainty

Temperature (°C)	Average reaction rate (mol dm^{-3} s^{-1})
15	0.038 ± 0.002
25	0.043 ± 0.001
35	0.059 ± 0.001
45	0.074 ± 0.002

PRESENTING DATA IN GRAPHS

A **graph** is an incredibly useful tool. In chemistry, data is plotted using a line or curved graph and it is sometimes called a scatterplot.

Plotting data in graphs has many advantages.

- Data can be represented visually.
- Data points that are incorrect can be ignored.
- Constants can be determined from the properties of the line or curve.
- Uncertainties can be represented as **error bars**.
- A general **line of best fit** that matches the trend of the majority of the data points can be determined.
- Any known relationship between the dependent and independent variables can be supported or rejected, especially by the use of uncertainties.

Tables often provide more detailed data than graphs, but graphs can suggest a possible relationship between the dependent and independent variables, and highlight if none exists. From the shape of the curve, it is easier to observe trends, patterns and gaps in data in graphical form than in tabular form.

Graphs, like that shown in Figure 1.6.2, are used when two variables are being considered and one variable is dependent on the other. The graph shows the relationship between the variables.

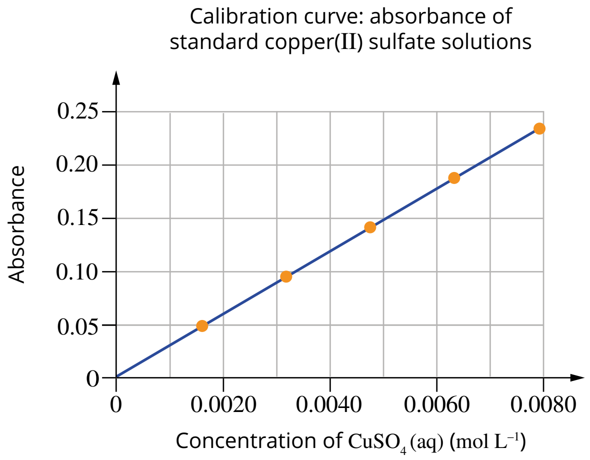


FIGURE 1.6.2 A graph showing the relationship between two variables: absorbance (independent variable) and concentration of $CuSO_4$(aq) (dependent variable)

DRAWING GRAPHS IN CHEMISTRY

For use in chemistry, every graph must have the following features.

1. The graph must have a suitable set of axes for the independent and dependent variables.

   Most chemistry graphs are only drawn in the first quadrant, but some graphs use all four quadrants (i.e. have negative values for the independent and dependent variables).

Remember that:

- the independent variable is the variable that you deliberately change and always goes along the x-axis
- the dependent variable is the variable that changes and always goes on the y-axis.

2 The graph must have a suitable scale that allows information to be seen easily and quickly.

For most graphs, it is important to choose a scale that is linear and has tick marks at common intervals (e.g. 2, 5, 10, 100 units).

3 The title and axes must have labels. Make sure that the axis labels show information about the quantity and units. Give a descriptive title that is not just a repeat of the labels of the axes.

4 Data points should be plotted as a small circle or small cross indicating their position.

5 Error bars should be drawn for each point. These indicate the absolute uncertainties in the independent (a horizontal line) and dependent (a vertical line) values. They are drawn as a horizontal line centred on the data point indicating the independent variable's uncertainty left and right of the value, and as a vertical line centred on the data point above and below the value, indicating the dependent variable's uncertainty. Combining the x and y error bars forms an error rectangle that the true measurement is located within (Figure 1.6.3).

The smaller the errors bars for a given point, the more precise the measurement is.

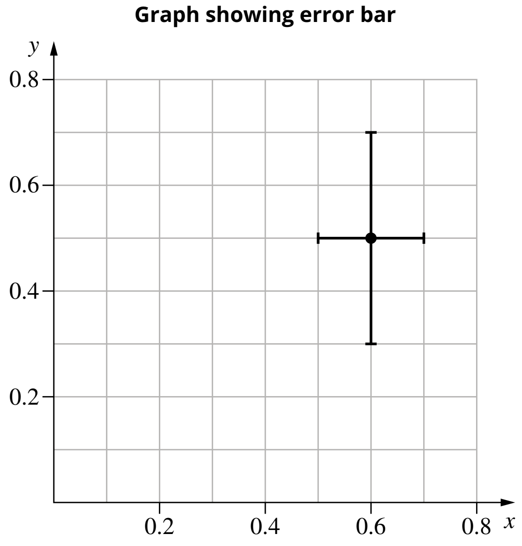


FIGURE 1.6.3 This error bar has an x value of 0.6 ± 0.1 and a y value of 0.5 ± 0.2.

Error bars are the graphical representation of the uncertainties in the independent and dependent variables.

The line of best fit

This is the straight line or a curve that best summarises the data. This line does not have to pass through every point or the origin, but does have to be one smooth continuous line or curve and *not* a 'join the dots' line (Figure 1.6.4). Typically, the line of best fit should have as many points on the line as possible, but if this is not possible then it should have as many points above the line as there are below the line.

The line of best fit is a straight line or curved line that best summarises the general pattern of the data on the graph.

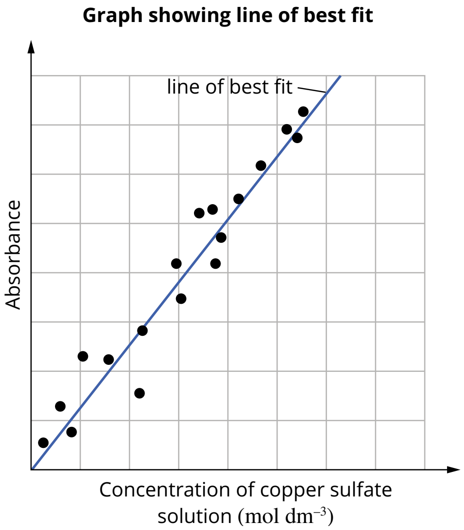


FIGURE 1.6.4 This graph shows that the line of best fit should normally pass through the error bars of each point and be extrapolated so that it intersects the y-axis.

These are some of the general rules to follow when making a graph.

- Keep the graph simple and uncluttered.
- Use a descriptive title.
- Represent the independent variable on the x-axis and the dependent variable on the y-axis.
- Start each axis at zero.
- Match the length of the axes to the data.
- Clearly label axes with both the variable and the unit in which it is measured.
- Use small symbols such as circles or crosses for data points.
- Use different symbols for different data sets.

TYPES OF GRAPHS SEEN IN CHEMISTRY

There are several types of graphs, including line graphs, bar graphs and pie charts. The best one to use will depend on the nature of the data. In chemistry, many relationships in the form of formulas and graphs are seen that exist between two variables. Each graph below shows some particular relationship between the independent and dependent variables that occur often in chemistry.

You will encounter all of these graphs at some stage during your Chemistry course.

Scatterplots and line graphs

Scatterplots are commonly used to display data in the form of a graph, and can be used to plot raw or processed data. They are used to show the relationship between two variables when one variable is dependent on the other.

The independent variable, which is set by the experimenter, is always shown on the x-axis. The dependent variable, which is the variable measured in the experiment, is always shown on the y-axis. The data is plotted on the graph as a series of points. Each point should be drawn in pencil as a small circle or cross. Alternatively, you can use a computer program to generate your graphs.

A **line graph** is a good way of representing **continuous quantitative data**. In a line graph, the values are plotted as a series of points on the graph. A line can then be drawn from each point to the next, as shown in Figure 1.6.5. This line shows the change in data from one point to the next but does not predict the value of a point between the plotted data.

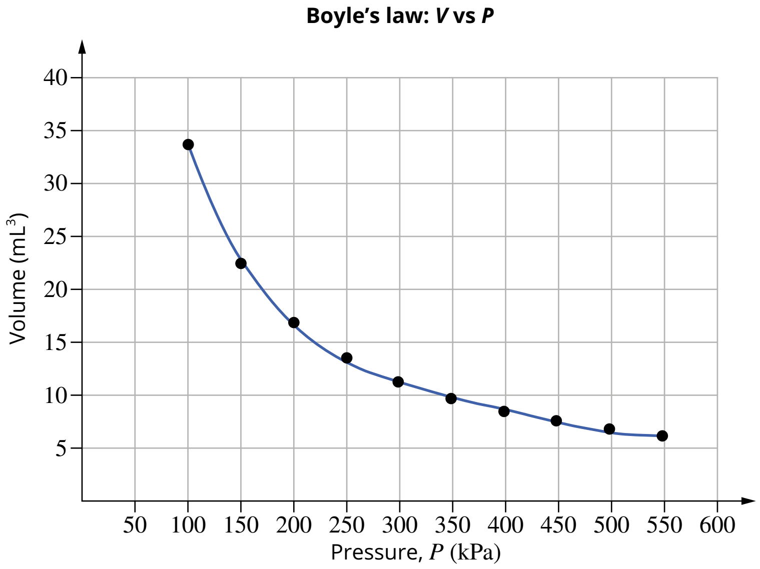


FIGURE 1.6.5 A line graph showing the change in volume of a gas in a syringe against the pressure of the gas inside

Alternatively, a single straight or curved line can be drawn, as shown in Figure 1.6.6, which is a **trend line**, or a line of best fit. It is used to show the overall trend in the data, and can be used to predict values between the data points. A line of best fit usually does not pass through every data point. Its position can be estimated by eye (ensuring an equal number of data points above and below the line when drawn), but mathematical or graphing programs can be used to determine the line of best fit more accurately. This will be covered in more detail in Module 1.8.

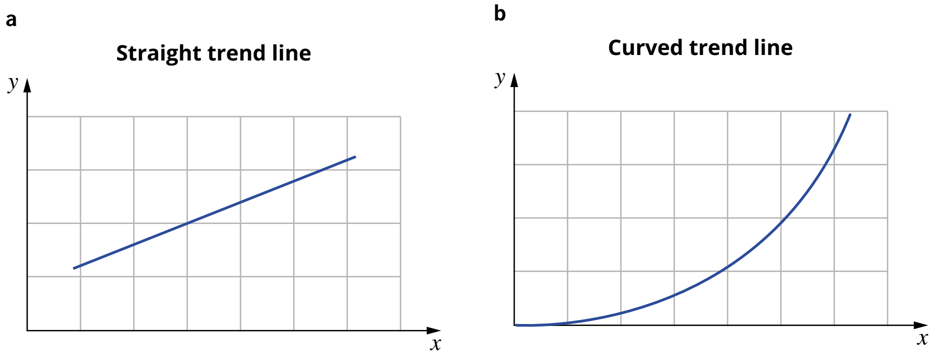


FIGURE 1.6.6 Graphs showing (a) straight and (b) curved trend lines

Bar and column graphs

Bar and column graphs are used to show categories and **discrete (discontinuous) data** (see Module 1.8 Research and planning).

- A **column graph** shows the value of the dependent variable by the height of the column; the categories are labelled across the *x*-axis.
- A **bar graph** shows the value of the dependent variable by the length of the horizontal bar; the categories are labelled up the *y*-axis.

Bar and column graphs are commonly used when the independent variable is categorical rather than numerical, or when the numerical data is discrete. The bars or columns are always the same width and the same distance apart.

Bar and column graphs are very useful for graphing qualitative and discrete data, such as the water quality data from various locations along a waterway (Figure 1.6.7).

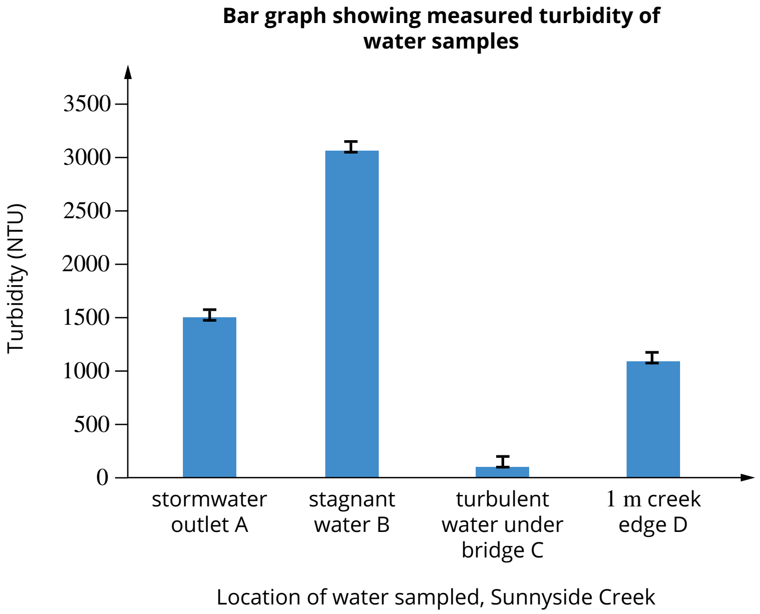


FIGURE 1.6.7 A bar graph can be used to compare data. This graph shows the measured turbidity of water samples taken from various locations along Sunnyside Creek.

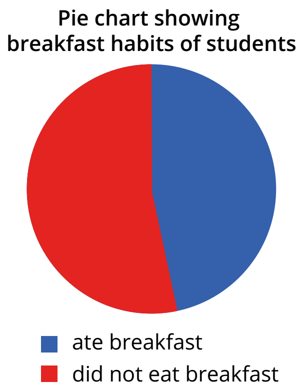


FIGURE 1.6.8 A pie chart presenting data on the breakfast habits of students

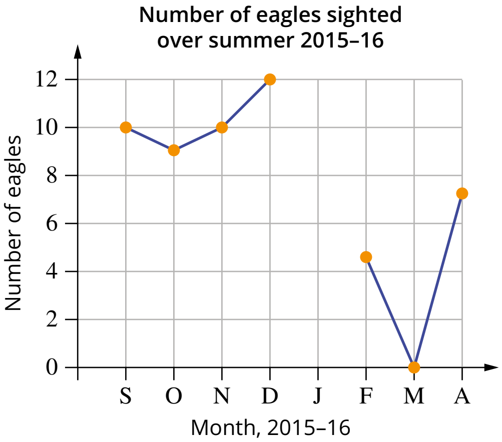


FIGURE 1.6.9 A line graph with data missing (for January)

Pie charts

A **pie chart** is a way of presenting qualitative or discrete quantitative data (see Module 1.8 Research and planning). It shows each category of data as a proportion of the total data. The chart is a circle divided into sections according to the proportions of each category, like slices of a pie (Figure 1.6.8). Each category is coloured or shaded differently so that it can be distinguished clearly from the other categories. It is recommended that pie charts are only used when there are few categories.

A circle is equal to 360°. To draw a pie chart you must find how many degrees are needed for each category.

1. Add the amounts in each category to find the total.
2. Divide 360° by the total (this will tell you how many degrees of the circle one value is worth).
3. Multiply the answer by the amount in the first category. Your answer will be in degrees, which can then be marked for the first category using a protractor on the circle.
4. Repeat for each category.

Missing data

When you have missing data, leave a gap for it, as shown in Figure 1.6.9. Ensure that the axes are complete (do not skip values) and do not join the data points that have missing data points between them.

Outliers

Sometimes when you collect data, there may be one point that does not fit the trend and may be an error. This is called an **outlier**. An outlier is often caused by a mistake made in measuring or recording data, or from a random error in the measuring equipment. This is covered in more detail later in this module. If you have an outlier in your results, you include it in your graph but ignore it when drawing the line of best fit as in Figure 1.6.10 where the point (0.48, 0.08) is an outlier.

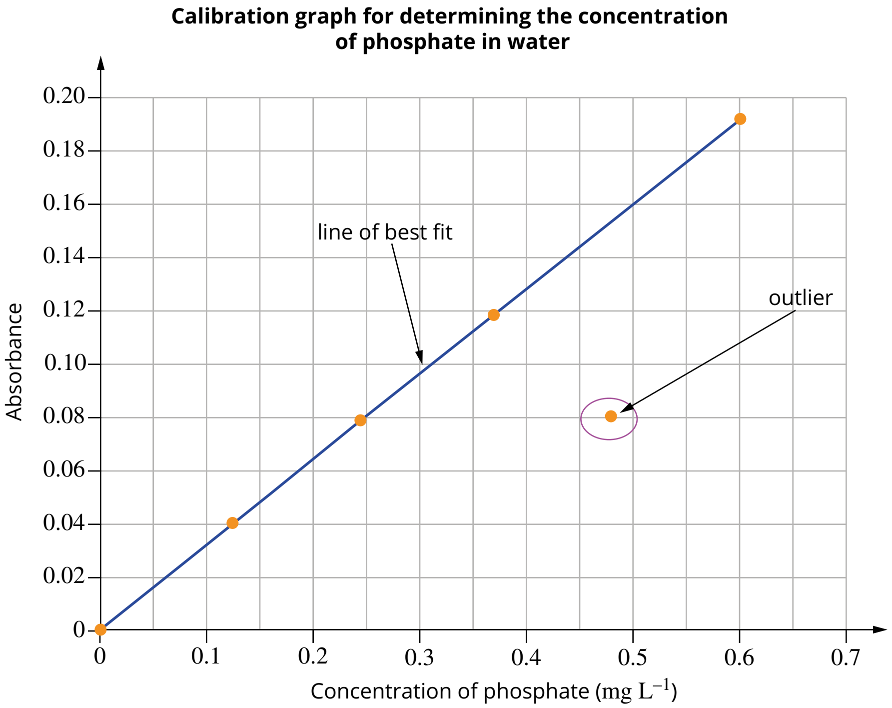


FIGURE 1.6.10 The graph is a calibration graph for determining the concentration of phosphate in water. A line graph shows an outlier, which has been ignored when adding the line of best fit.

Distorting the truth

Poorly constructed graphs can distort the truth. For example, in Figure 1.6.11, you can see two graphs that show the same data—the test results of two groups of students. One group of students did not eat breakfast before doing the test and scored an average of 42 marks out of 50. The other group of students did eat breakfast and scored an average of 48 marks out of 50. One graph distorts the difference in marks between the two groups by using a scale of only 40 to 50 marks on the *y*-axis. It is important to make sure the graphs you create do not distort your data. You should also be wary of distorted data when interpreting graphs in other publications.

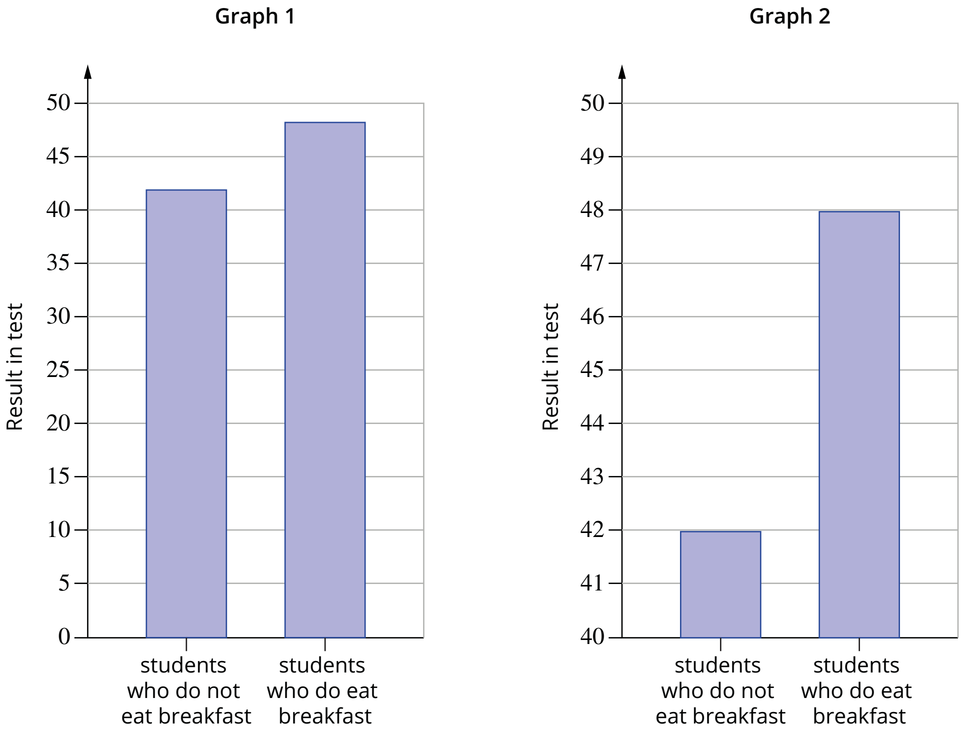


FIGURE 1.6.11 Graph 1 shows the difference between two groups of students out of the total 50 marks on the *y*-axis. Graph 2 shows the difference between the two groups within only a narrow range of marks on the *y*-axis, which distorts the difference and makes it appear larger than it really is.

1.6 Review

SUMMARY

- Tables and graphs are used to represent scientific data.
- In general, tables provide more detailed data than graphs, but it is easier to observe trends and patterns in data in graph form than in table form.
- Tables organise data into rows and columns, and can vary in complexity according to the nature of your data.
- Tables should have the following features:
  - a descriptive title
  - column headings (including the units)
  - aligned figures (align the decimal points)
  - the independent variable placed in the left column
  - the dependent variable placed in the right column.
- Graphs are used to display relationships, trends and patterns between two variables with one proposed to be dependent on the other.
- There are several types of graphs, including line graphs, bar graphs and pie charts. The best one to use will depend on the nature of the data.
- Scatterplots are commonly used to display data in the form of a graph, and can be used to plot raw or processed data.
- A trend line (line of best fit) is used to show the overall trend in the data, and can be used to predict values between the data points.
- A line graph is a good way of representing continuous quantitative data.
- Bar and column graphs are used to show categories and discrete (discontinuous) data.
  - A column graph shows the value of the dependent variable by the height of the column; the categories are labelled across the *x*-axis.
  - A bar graph shows the value of the dependent variable by the length of the horizontal bar; the categories are labelled up the *y*-axis.
- A pie chart is a way of presenting qualitative or discrete quantitative data.
- Missing data is represented in graphs by leaving a gap.
- An outlier is a data point that does not fit the trend, and may be caused by an error.

KEY QUESTIONS

Retrieval

1 State the purpose of a trend line.

Comprehension

2 Explain what outliers are.

3 Identify the relationship between the variables indicated by a sloping linear graph.

4 Identify the relationship that exists if one variable decreases as the other increases.

5 Explain how the general pattern (trend) of a graph can be represented once the points are plotted.

6 Describe the relationship that exists if both variables increase or both decrease at the same rate.

Analysis

7 Antonia draws a straight line graph of data, including error bars, from an experiment she has just completed. She finds that it is not possible to draw a straight line of best fit through every error bar. Analyse the problem and advise Antonia on what to do.

8 If theory suggests that the *y*-intercept of a line is zero, deduce what this should look like on a linear graph.

9 Assess when is it more appropriate to use a graph showing a line of best fit rather than a ruled line graph joining the line from point to point.

10 Consider the data below and determine the relationship between atmospheric pressure and altitude.
Describe the trend and draw a line of best fit through the data.

Pressure (atm)	1.00	0.79	0.52	0.23	0.11	4.5×10^{-2}	1.3×10^{-2}	6.0×10^{-3}	2.1×10^{-3}
Altitude (km)	0	2.0	5.0	10.0	15.0	20.0	30.0	40.0	50.0

1.7 Statistics

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- calculate and interpret the range, mean, median and mode of a set of data
- evaluate the statistical significance of standard deviation
- use the R^2 values to evaluate the fit of a trend to a set of data.

One of the difficulties in chemistry is that direct measurements of phenomena or measurements within a closed system are challenging to obtain. Therefore, the limitations and uncertainty in the data include uncertainties in the accuracy and precision of the measurements that were recorded as well as errors related to reliability and validity of the experimental procedure. To address this problem, chemists perform statistical analysis to help understand the quality of the measurements, the results and their meaning.

This module will outline a basic understanding of common statistical analysis used in chemistry relevant to the Chemistry General Senior Syllabus.

STATISTICAL MEASURES

There are a number of common statistical measures that help describe data accurately.

Range

The range is the difference between the highest and lowest values in a data set. Table 1.7.1 shows the measurements taken for a reaction between sodium thiosulfate and sulfuric acid.

TABLE 1.7.1 Time for the cross to disappear in the reaction between sodium thiosulfate and sulfuric acid at 20°C and 50°C

Temperature	1	2	3	4	5	Mean	Range
20°C	253	158	201	169	187	187.6	253 – 158 = 95
50°C	69	85	70	91	75	78	91 – 69 = 22

To determine the range for values in Table 1.7.1, you would subtract the smallest value from the largest value. Notice how an abnormally large or abnormally small value in the data set makes the variability appear high. If one value appears way out of range, such as the time taken for the cross to disappear the first time the reaction was performed at 20°C, it may be an outlier. This illustrates the importance of observation during experiments and recording all details in a journal as experimentation and research is conducted.

The experiment was conducted with a smaller number of measurements taken at each temperature, and the results seen in Table 1.7.2.

TABLE 1.7.2 Time for the cross to disappear in the reaction between sodium thiosulfate and sulfuric acid at 20°C and 50°C

Temperature	1	2	3	Mean	Range
20°C	253	158	201	194	253 – 158 = 95
50°C	69	85	70	75	85 – 69 = 16

A larger **sample size** will strengthen the **validity** of the range as described above with results that may be errors.

Even though the range has not changed much between Tables 1.7.1 and 1.7.2, the results at 20°C seem more random. In Table 1.7.2 at 50°C, the second reaction could be an outlier as the other two results show excellent agreement. The range simply outlines the total variation in measured results; no other analysis or interpretation can be achieved.

Measures of central tendency

Measures of central tendency are single values that allow you to describe the central position in a set of data. Measures of central tendency are sometimes also called measures of central location. The mean, median and mode are all measures of central tendency.

Consider the following data set: 3, 5, 7, 8, 8, 8, 10.

- The **mean** (or average) is the sum of the values divided by the number of values, which in this case is $(3 + 5 + 7 + 8 + 8 + 8 + 10) \div 7 = 7$.

$$\text{mean} = \frac{\sum(x_1 + x_2 + x_3 + \ldots)}{n}$$

where $\sum$ = sum of

$x^1 + x^2 + x^3 = \text{measurement}_1 + \text{measurement}_2 + \text{measurement}_3$

n = total number of measurements

- The **median** is the 'middle' value in an ordered list of values, which in this case is the fourth value, which is 8.
- The **mode** is the value that occurs most often in a list of values, which in this case is 8. This measure is particularly useful for describing qualitative or discrete data.

The appropriate measure of central tendency to use depends on the type of data you are working with (Table 1.7.3).

TABLE 1.7.3 When to use the different measures of central tendency

Type of data	Mean	Median	Mode
nominal (qualitative)	✓	✗	✗
ordinal (qualitative)	✓	✓	maybe
discrete or continuous (quantitative)	✓	✓	✓

Graphically representing the data using central tendency provides a clear and succinct display of results. As there are many measured values and results in raw data, it can be difficult to analyse and interpret. It is much easier and less complicated to analyse and interpret processed data compared to raw data. Raw data includes all the measured results and usually a large number of measurements or observations.

Processed evidence is usually a smaller number of results that represent all the measured raw data; these representative results are calculated using statistical methods such as mean, median and mode. Therefore, an analysis can be conducted to compare results using three results (e.g. three means) rather than nine results (e.g. nine individual measurements in the raw data). Interpreting a relationship is also much easier and less complicated when using fewer numbers.

It is important to realise that statistical calculations do not produce the true value, but rather a single value (or smaller number of values) that is meant to represent many. Analyse the mass of two alcohols burnt in a calorimetry experiment, shown in Table 1.7.4. The average has been calculated for each alcohol; however, you will notice that the mean value was not measured during the experiment and is not in the raw data.

TABLE 1.7.4 Mass of ethanol and hexanol burnt

Ethanol (g)	Hexanol (g)
0.77	1.60
0.78	1.82
0.72	1.67
0.73	1.72
0.76	1.71
0.74	1.80
Mean 0.75	Mean 1.72

Standard deviation

Standard deviation specifically estimates the spread of data in a population or set of values. It calculates the spread or dispersion of data and its distance from the mean.

The standard deviation statistical calculation assumes that the data being used is from a set in which a normal curve (or normal distribution) exists. A normal distribution of data forms a 'bell' curve or a normal curve as seen in Figure 1.7.1.

A normal curve on a graph has a characteristic distribution of data.

- Approximately 68.2% of all the data lies within 1 standard deviation (1σ) of the mean. That is, 34.1% of data lies above the mean and 34.1% below the mean.
- Approximately 27.2% of all the data lies between the 1st and 2nd standard deviation (between 1σ and 2σ). That is, 13.6% of data lies above 1σ and 13.6% below 1σ.
- Approximately 4.2% of all data lies between the 2nd and 3rd standard deviation (between 2σ and 3σ). That is, 2.1% of data lies above 2σ and 2.1% below 2σ.
- The final 0.2% of data lies above and below 3σ, 0.1% on either side.

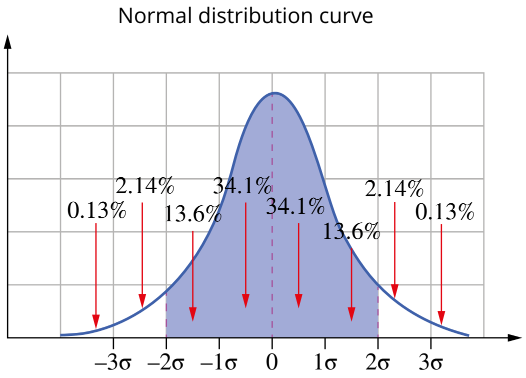


FIGURE 1.7.1 Normal distribution curve showing how far each standard deviation (σ) is from the mean (0)

When using standard deviation to analyse data, it is important to understand what has been calculated. If a measured phenomenon is expected to be equivalent from one test to another, then it would be expected that all measurements would be very close to the mean, except for unusual results like an error. You could decide that all accurate measurements should be within two standard deviations (2σ) of the mean. Therefore, any measurement outside (greater than) 2σ from the mean would be considered an outlier (measurements that have a value outside of the expected calculation).

If standard deviation was used for a set of data in which variation is expected, it could be used to demonstrate the variation. For example, according to the Australian Bureau of Statistics, the average height of an adult female in Australia is 161.8 cm. If the height of all the females in the Chemistry class were measured with the following results (Table 1.7.5), you could use standard deviation to estimate where individuals fit into the assumed normal curve for the height of adult females in Australia.

Using the formula to calculate standard deviation:

$$\mathrm{SD} = \sqrt{\frac{\sum\left(x_1-\bar{x}\right)^2+\left(x_2-\bar{x}\right)^2+\left(x_3-\bar{x}\right)^2}{n-1}}$$

$\sum$ = sum of

x_1 = measurement$_1$

$\bar{x}$ = mean

$n - 1$ = number of measurements − 1

TABLE 1.7.5 Height of female students in the Chemistry class

Sample	Height (cm)
student 1	164.5
student 2	156.7
student 3	172.2
student 4	180.6
student 5	163.4
student 6	158.8
student 7	153.9
student 8	150.7
student 9	160.4
student 10	156.8
mean	161.8

The standard deviation for the set of data is ± 8.5 cm. Student 4 at 180.6 cm is more than 2σ away from the mean, therefore it can be said that this student is estimated to be in the top 3% of the adult female population for height. This is not an outlier because variation is expected, so standard deviation was used to understand the variance of female students in the class.

REGRESSION

Regression is the term used to state that the independent variable causes the result in the dependent variable during a scientific investigation. Some relationships between variables are causal, meaning that there is a cause and effect relationship. The independent variable is the cause and the change dependent variable is the effect.

There are numerous statistical calculations that estimate the causal relationship between variables; some calculations provide a value that indicates how directly an independent variable causes the dependent variable result. This attempts to show the strength of the relationship, and to identify if the change in the dependent variable is entirely due to the independent variable.

In chemistry, not all relationships are direct or straightforward cause and effect. Often there are many variables that influence the dependent variable and it is helpful to try and predict how much of an effect the independent has on the dependent, while other variables are also influential.

Coefficient of determination (R^2)

The **coefficient of determination** (R^2) can calculate a value that indicates the predictability between the variables based on how much the results vary. The coefficient of determination produces a value that is an estimate as to how well you can predict the dependent variable based on the independent variable, considering the variations in the results. In the examples in Figure 1.7.2, the data of technique 1 does not follow the linear trend line exactly and therefore as the concentration changes (the independent variable), the absorbance (dependent variable) does not always change consistently. Sometimes it changes by 5, sometimes by 2 or even by 10 or more. The coefficient of determination (R^2) is a calculation that considers the variation in the results to estimate the predictability.

For example, a laboratory has used two different analytical techniques to determine the concentration of a performance-enhancing drug in a blood sample from an athlete suspected of cheating. The calibration curves for each technique and the R^2 value for each are shown in Figure 1.7.2.

Based on the lines of best fit and R^2 values, the second method was chosen to analyse the athlete's blood sample.

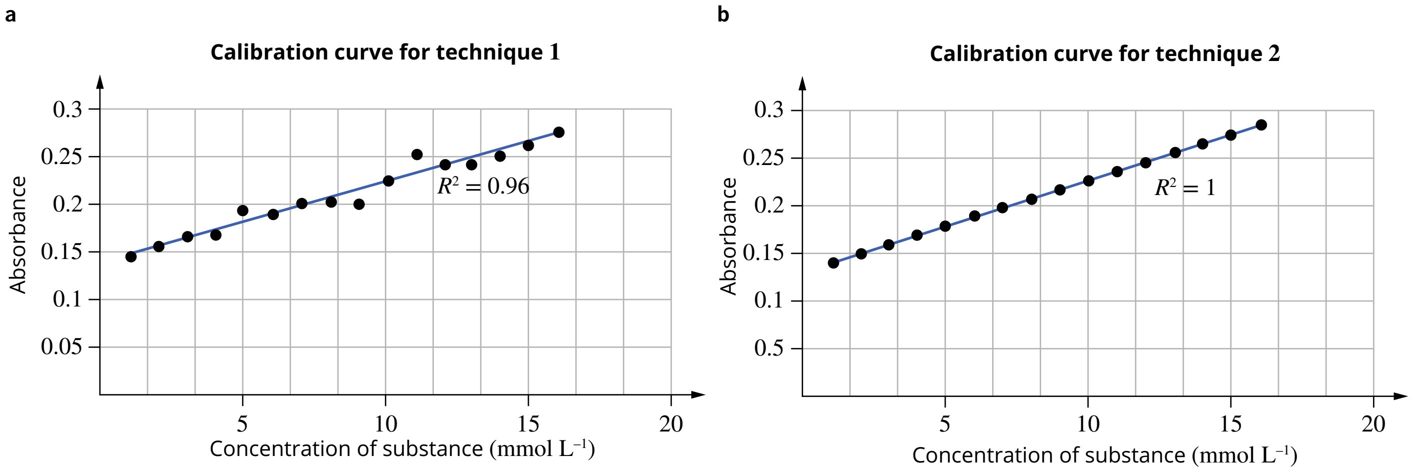


FIGURE 1.7.2 (a) The calibration curve for technique 1: linear regression was used to establish a line of best fit. (b) The calibration curve for technique 2: linear regression was used to establish a line of best fit.

Even though both techniques measured the same samples, the data is different. As can be seen in the graphs, the linear trend line is very similar with a similar slope; however, technique 1 has an R^2 value of 0.96 and technique 2 is 1.00.

The coefficient of determination produces a value between 0 and 1. If the value is 0, it is unable to predict the change in the dependent variable from the independent variable due to so much variation in the results. If the value is 1, then it can predict the change in the dependent variable 100% of the time without error. Any other number between 0 and 1 is considered a percentage, for example 0.76 means that the coefficient of determination can predict 76% of the dependent variable results due to the independent variable. It is interpreted as 76% of the variation in the dependent variable can be explained by the independent variable.

In Figure 1.7.2b, the absorbance of the sample was perfectly correlated without error with the concentration of the species. This suggests that the results from technique 2, including error, **bias** and instrumental precision, are highly reliable.

Figure 1.7.2a shows that the absorbance of the sample was not precisely correlated with the concentration. The R^2 value for technique 1 is lower than in technique 2, meaning the results are less predictable due to more variation. Technique 1 was less precise compared to technique 2; therefore, technique 2 is the preferred technique to determine the concentration in the actual sample.

1.7 Review

SUMMARY

- The range is the difference between the highest and lowest values in a data set.
- The mean, $\bar{x}$, is the average of a data set. It is calculated using the formula:

  $\bar{x} = \frac{\sum x_1 + x_2 + x_3 \ldots}{n}$, where

  $\sum$ = sum of

  $x_1 + x_2 + x_3 =$ measurement$_1$ + measurement$_2$ + measurement$_3$

  n = total number of measurements.
- The median is the 'middle' value in an ordered list of values.
- The mode is the value that occurs most often in a data set. The mode is particularly useful for describing qualitative or discrete data.
- Standard deviation is a measure of the spread or dispersion of data from the mean.
- Regression is the term used to state that the independent variable causes the result in the dependent variable during an experiment.
- Linear regression is a calculation that estimates a direct linear relationship between the independent and dependent variables.
- The coefficient of determination estimates how well you can predict the dependent variable based on the change in the independent variable.
- The coefficient of determination produces an R^2 value between 0 and 1.
  - The closer R^2 is to 0 the less predictable the dependent result is based on the independent value.
  - The closer the R^2 is to 1 the more predictable the dependent result is based on the independent value.

KEY QUESTIONS

Retrieval

1 Define the terms 'mean', 'median' and 'mode'.

Comprehension

2 Describe what the standard deviation informs scientists about a data set.

Analysis

3 For the data set {21, 28, 19, 19, 25, 24, 20}, calculate the:

a range
b median
c mode
d mean
e standard deviation

4 Identify which of the statistical measurements of mean, mode and median is generally most affected by an outlier.

PART B STUDENT EXPERIMENT

QCAA makes the following statement relating to the student experiment in Units 1 and 2 Chemistry: 'It is suggested that student performance on Unit 1 is assessed using a student experiment and data test modelled on the techniques used in Unit 3.'

The QCAA requires students to complete a student experiment in Unit 3 Chemistry. In preparation for Unit 3, teachers may choose to assign a similar assessment task in Units 1 or 2, as preparation for Unit 3.

The student experiment assessment task requires students to research a question or **hypothesis**. Students use research conventions to investigate the question or hypothesis by collecting, analysing and synthesising **primary data**. The experiment requires students to locate and use information beyond the scope of their knowledge and what they have been given.

The student experiment requires you to undertake the full scientific method. The Queensland Chemistry Syllabus states that this process begins with a practical conducted during class, either a mandatory or suggested practical. This in-class practical will be altered to conduct your own experiment. It is recommended that during the class practical you record your observations, queries and thoughts in a logbook. These notes can be used to lead to a research question or hypothesis for the student experiment.

The student experiment constitutes 20% of the total assessment in Unit 3 Chemistry.

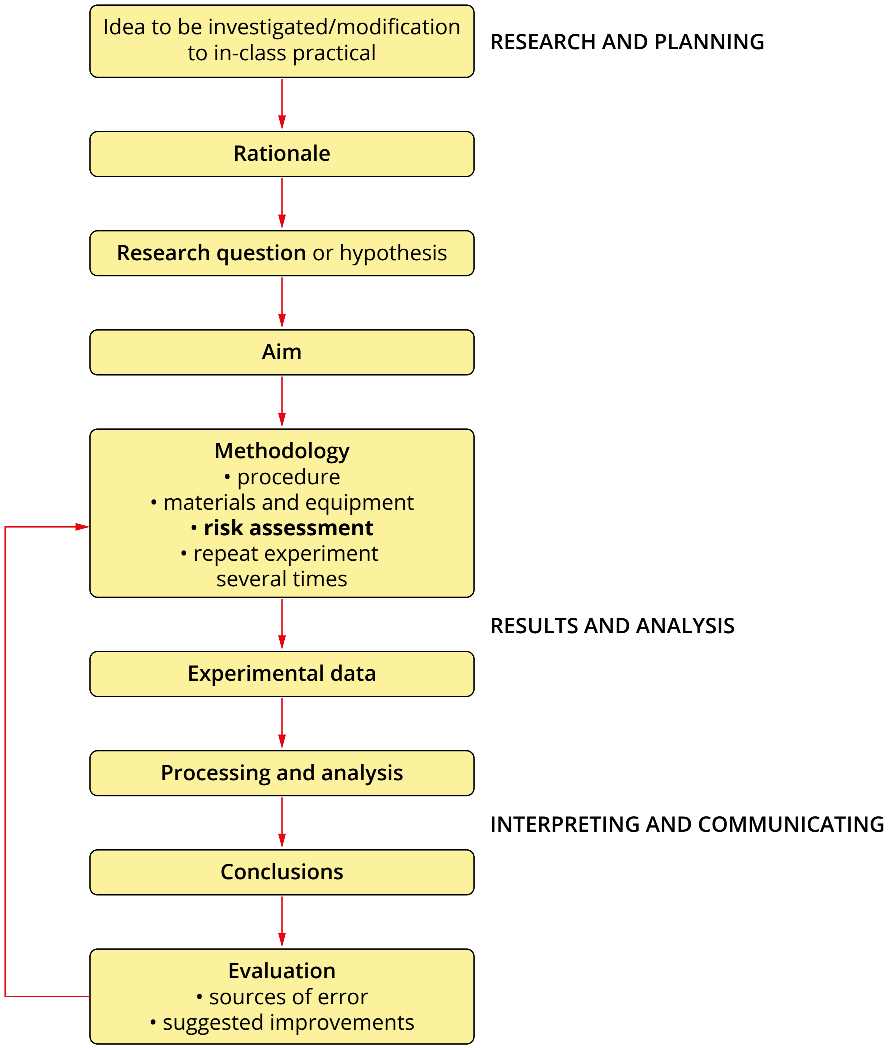
A summary of the objectives and marking for the summative internal assessment: Student experiment, IA2, (Units 3 and 4) is provided below.

The student experiment may be presented in:

- written form (e.g. scientific report), 1500–2000 words, or
- multimodal presentation form (e.g. poster presentation), 9–11 minutes.

Criteria	Assessment objectives	Demonstrated by	Marks
Research and planning	• Apply understanding • Investigate	• a considered rationale for the experiment • justifications for the experiment • a research question that is specific and relevant • collected data that is sufficient and relevant • considered risks and issues (ethical and environmental) and their management	6
Analysis of evidence	• Apply understanding • Analyse evidence • Investigate through experimentation	• use of relevant algorithms and correct data processing • detailed and careful coverage of relevant trends, patterns and relationships in the evidence • detailed and careful coverage of uncertainty and limitations of evidence • collection of relevant raw data and sufficient data	6
Interpretation and evaluation	• Interpret experimental evidence • Evaluate experimental processes and conclusions	• a conclusion that is justified and addresses the research question • a discussion about the reliability and validity of the experiment that is supported by evidence • providing possible improvements and extensions to the experiment based on examination of evidence	6
Communication	• Present the experimental findings, including methodology, conclusions, evaluation	• scientific language and representations that are concise and fluent • suitable use of genre conventions • appropriate referencing conventions to acknowledge sources	2
Total			20

The scientific inquiry is not a linear process. Scientists will not necessarily complete these steps in the stated order and some steps may need to be repeated or altered in order to more accurately address the research question, as demonstrated below.



INSTRUMENT-SPECIFIC MARKING GUIDE

Student responses are assessed against an instrument-specific marking guide (ISMG). In developing your experiment and planning your response, it is important to always have in mind the assessment objectives, and in particular the characteristics that are described in the performance level descriptors.

The major features of ISMG are outlined below and shown for the communication criterion.

- The ISMG is differentiated into four criteria (criterion): research and planning, analysis of evidence, interpretation and evaluation, and communication.
- Criteria are differentiated into performance levels, against which the qualities of the response are assessed.
- A performance level is comprised of a performance level mark, which may be a single mark or two-mark range, and a performance level descriptor.
- The performance level descriptor describes the characteristics that are demonstrated by a response at this quality.

In the modules that follow, you will find a guide to a scientific method (Modules 1.8–1.10) followed by an outline for producing a **scientific report** (Module 1.11).

TYPES OF INVESTIGATIONS

There are many different types of investigations that can be conducted in chemistry. You are probably most familiar with practical investigations or experiments. Other types of investigations involve researching and evaluating primary and/or secondary sources of information to answer a question. Examples of the types of different investigations are listed in Tables 1.7.6 and 1.7.7.

TABLE 1.7.6 Examples of primary research

Examples of primary research	Example tasks
conducting experiments in a laboratory	• planning a valid experiment • conducting a risk assessment • working safely • recording observations and results • analysing and evaluating data and information
conducting field work	• conducting a risk assessment • working safely • recording observations and results • analysing and evaluating data and information
conducting surveys	• writing questions • collecting data • analysing data and information
designing a model	• identifying a problem to be modelled • summarising key findings • identifying advantages and limitations of the model

TABLE 1.7.7 Examples of secondary research

Examples of secondary research	Example tasks
researching published data from primary and secondary sources	• finding published information in scientific magazines and journals, books, databases, media texts, labels of commercially available products • analysing and evaluating data and information

Before you are able to start the practical side of your investigation, you first need to understand the theory behind it. This section is a guide to some of the key steps that should be undertaken when first developing your research questions and hypotheses.

Taking the time to carefully plan and design a practical investigation before you begin will help you to maintain a clear and concise focus throughout. Preparation is essential. It is important to have both a solid understanding of the theory behind your investigation and a detailed plan for the practical components of your investigation. In this module, you will learn about the key steps you should take when planning and designing a practical investigation.

1.8 Research and planning

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- identify and explain the difference between controlled, measured, independent and dependent variables
- develop a research question or hypothesis
- evaluate research questions and hypotheses
- use a scientific journal to record experiments and experimental data
- plan, evaluate and refine scientific experiments
- explain what validity and reliability mean in relation to experimentation
- explain the difference between, and identify, qualitative and quantitative data
- characterise qualitative data as either nominal or ordinal
- characterise quantitative data as either discrete or continuous
- explain the difference between replication and repeat trials
- conduct risk assessments for planned experiments
- recognise common chemical GHS codes and symbols
- understand the criteria against which research and planning will be assessed.

All scientific work begins with research and planning; to conduct an experiment, as much as possible must be known about the relationship between variables. This includes understanding the relationship between controlled or measured variables as well as the independent and dependent variables (see Module 1.7). If variables are not controlled appropriately then errors and unknown quantities will influence the results. Furthermore, relationships cannot be established or observed and the experiment becomes invalid or unreliable (or both). Research and planning is the foundation of the scientific method and is always recorded in a journal.

The scientific method is non-linear and experimenters should always return to previous stages of the method with new ideas or knowledge. The journal will show a chronological record of ideas, development of knowledge and understanding, planning and refinement. Even though the journal will be in chronological order it most likely will not be entirely in a conceptually logical order. The journal is an ongoing draft of scientific work from which the final scientific report is written.

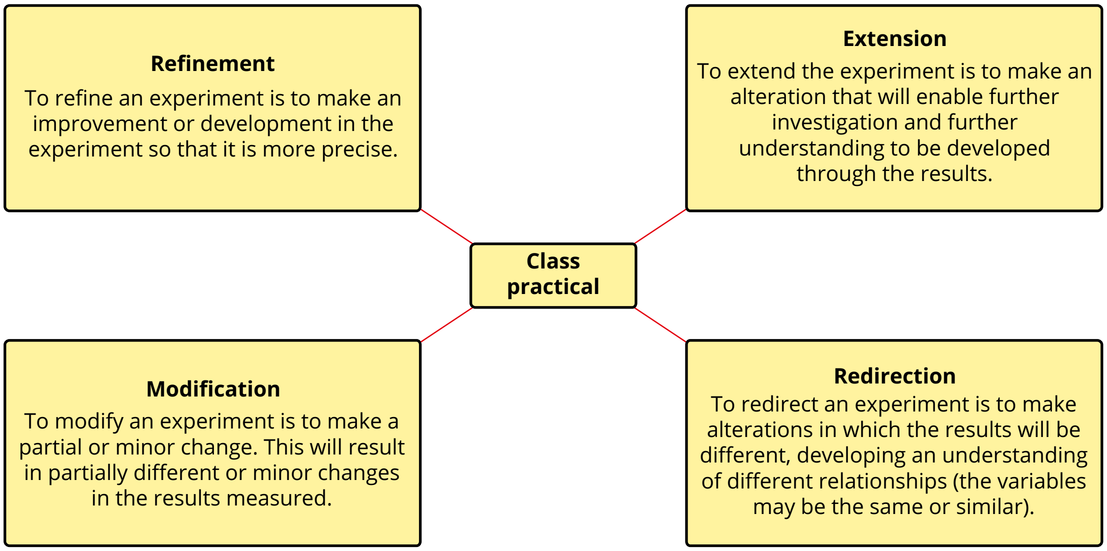
IDENTIFY AN EXPERIMENT AND DEVELOP A RESEARCH QUESTION OR HYPOTHESIS

Identifying an experiment for the Student experiment only requires you to modify, refine, extend or redirect a practical undertaken in class. Therefore, the experiment will be similar to the class practical with an alteration to investigate something slightly different, as shown in Figure 1.8.1 on page e57.

Identify an experiment

When choosing a topic or observation to alter, consider the following:

- Start with a topic/observation of which you already have some background knowledge or understanding, or some idea about how to perform the experiment.
- Choose a topic/observation you find interesting.



Refinement
To refine an experiment is to make an improvement or development in the experiment so that it is more precise.

Extension
To extend the experiment is to make an alteration that will enable further investigation and further understanding to be developed through the results.

Class practical

Modification
To modify an experiment is to make a partial or minor change. This will result in partially different or minor changes in the results measured.

Redirection
To redirect an experiment is to make alterations in which the results will be different, developing an understanding of different relationships (the variables may be the same or similar).

FIGURE 1.8.1 Changes that can be made to a class practical to develop the student experiment

- Check that your school laboratory has the resources for you to perform the experiment or investigate the topic.
- Choose a topic that can provide clear measurable data that shows how variable *y* depends on variable *x*.

Table 1.8.1 shows examples of modifying, refining, extending or redirecting practicals into a topic of interest for a student experiment.

TABLE 1.8.1 Examples of extending a practical into a student experiment

Practical	Modification, refinement, extension or redirection	Student experiment
Investigate the rate of chemical reactions. How does changing the temperature of the solution affect the rate of the reaction between HCl and Mg?	modification and refinement	How does changing the concentration of the hydrochloric acid affect the rate of the reaction between hydrochloric acid and magnesium?
	modification and refinement	How does changing the size of the magnesium strips affect the rate of reaction between hydrochloric acid and magnesium?
Conduct a calorimetry experiment to measure the enthalpy of a reaction. Measure the enthalpy of combustion of hexanol.	extension	Does using less water in the calorimeter increase the accuracy in determination of the combustion of hexanol?
	extension	Is the accuracy of the molar heat of combustion of hexanol improved by using multiple volumes of hexanol?

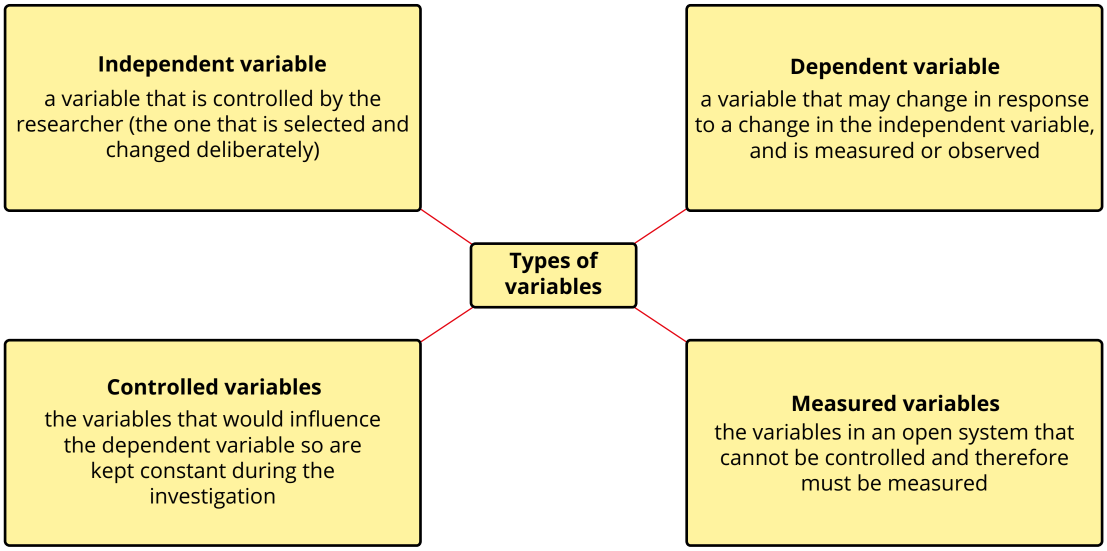
A justification of the alteration to the methodology of the student experiment is required by the QCAA Chemistry General Senior Syllabus. Module 1.9 addresses this in the section 'Refining the methodology'.

When altering the class practical to identify a student experiment, it is best to think of a single variable that may influence the outcome (independent variable). This may require some research. The more variables that are altered (including **measured variables** and **controlled variables**), the more research required and the more complex the task becomes. Note that some alterations of variables may require the alteration of other variables.

If only one variable is changed (the independent variable) then the class practical can be used as the control and the data collected can be used to compare results. If both the independent and dependent variable are altered then the data between the class practical and student experiment is not comparable. The student experiment will need to determine its own control.

Defining the variables

The factors that can change during your experiment or investigation are called the variables. An experiment or investigation determines the relationship between variables by measuring the results. There are four categories of variables (Figure 1.8.2). You should have only one independent variable. Otherwise you could not be sure which independent variable was responsible for changes in the dependent variable (the results).



Types of variables

Independent variable
a variable that is controlled by the researcher (the one that is selected and changed deliberately)

Dependent variable
a variable that may change in response to a change in the independent variable, and is measured or observed

Controlled variables
the variables that would influence the dependent variable so are kept constant during the investigation

Measured variables
the variables in an open system that cannot be controlled and therefore must be measured

FIGURE 1.8.2 The four types of variables in experiments

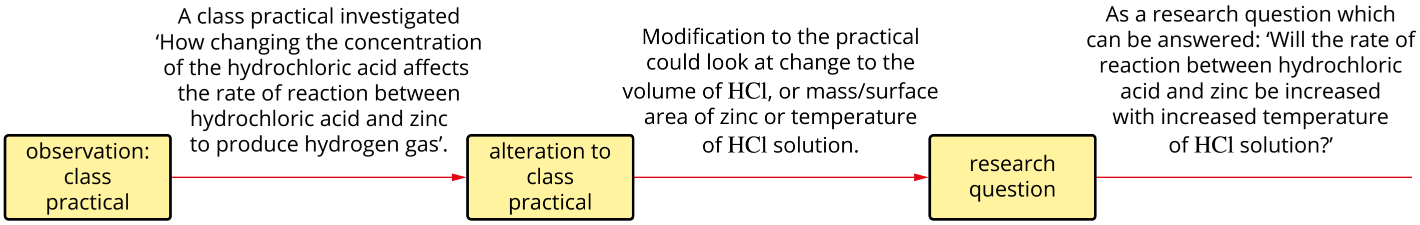
Developing a research question

A research question is defined as a question that directs the scientific inquiry activity. Its purpose is to focus the research investigation or student experiment, inform the direction of the research, and guide all stages of inquiry, analysis, interpretation and evaluation.

The question determines the experiment and the experiment is testing the question. A research question should:

- be specific and relevant to the class experiment
- clearly identify the subject matter of the experiment
- specify the scope or conditions of the inquiry
- aim to find trends or patterns of relationships between two variables.

Consider the example in Figure 1.8.3.



observation: class practical

A class practical investigated 'How changing the concentration of the hydrochloric acid affects the rate of reaction between hydrochloric acid and zinc to produce hydrogen gas'.

alteration to class practical

Modification to the practical could look at change to the volume of HCl, or mass/surface area of zinc or temperature of HCl solution.

research question

As a research question which can be answered: 'Will the rate of reaction between hydrochloric acid and zinc be increased with increased temperature of HCl solution?'

FIGURE 1.8.3 The process from class observation of a practical to altering the practical and developing a possible research question

Background research for the student experiment would follow to refine the question. The research could include:

- information about the variables
- correlations between variables
- ideas for refining the question—do not reject ideas that might seem improbable at this stage.

Figure 1.8.4 demonstrates the process of research question refinement and the resultant research question that will guide the student experiment.

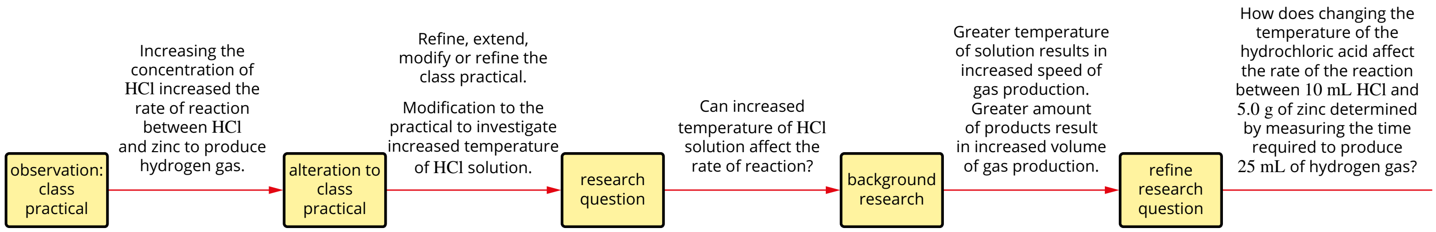


FIGURE 1.8.4 The process of refinement of the research question

The structure of a research question

A research question should state the independent and the dependent variables. It should ask if the independent variable will affect the dependent variable. The research question must specifically outline variables in a way that is measurable; this also allows the research question to be answered.

Look at this example below.

independent variable

'How does changing the concentration of the HCl affect the rate of the reaction between HCl and Mg?'

dependent variable

A well-constructed research question should have the following characteristics.

- It should include measurable variables—the independent variable and the dependent variable.
- It should have a guiding word, such as *who*, *what*, *why* or *will*.
- It should be phrased so that a definitive answer can be developed.
- You should be able to link the guiding word to command verbs such as *identify*, *describe*, *compare*, *contrast*, *distinguish*, *analyse*, *evaluate* or *create* so that a task can be determined.

Table 1.8.2 on page e60 gives examples of constructed research questions.

TABLE 1.8.2 Structuring a research question

Guiding word(s)	Example research questions	What are you being asked to do? What are the command verbs?
what	• What difference can nanomaterials make to society and the environment? • What are electrons, protons and neutrons made of?	Identify and describe specific evidence, reasons and examples from a variety of possibilities. *identify, describe*
where	• Where would an element with an atomic number of 130 be placed in the modern periodic table, what properties would it have and how likely is it to be discovered?	Identify and describe giving reasons for a place or location. *identify, describe*
will	• Will the use of a catalyst affect the rate of the reaction between zinc and hydrochloric acid?	Identify and describe giving reasons for effectiveness. *identify, describe*
how	• How are atoms seen? • How can lead be transformed into gold? • How does particle size affect the rate at which sugar dissolves in water?	Identify and describe in detail a process or mechanism. Give examples using evidence and reasons. *identify, describe*
why	• Why are the ten most abundant elements in the universe not the same as the ten most abundant elements on Earth? • Why do transition metals have multiple oxidation states? • Why does the composition of crude oil vary between different oil wells?	Explain in detail the causes, reasons, mechanisms and evidence for. *identify, explain*
would	• Would there be life if elements did not form compounds?	Evaluate. Justify, giving reasons for and against (using evidence, analogies and comparisons). *evaluate, assess, justify*
is/are	• Are there more elements to be discovered? • Is it an advantage or disadvantage for elements to be unreactive? • Is it worth sending people to the moon to mine for lanthanoids and actinoids?	Evaluate. Justify, giving reasons and evidence. *evaluate, assess, justify*
on what basis	• On what basis are alternative forms of the periodic table constructed?	Evaluate. Justify using reasons and evidence. *evaluate, assess, justify*
can	• Can enzyme concentration increase the rate of enzyme activity? • Can we live without lanthanoids and actinoids?	Evaluate and assess. Is it possible? Give reasons, suggest possible alternatives. *evaluate, assess, justify, create*
do/does	• Does surface area or volume affect the rate of diffusion? • Does surfactant biodegradability affect performance? • Do we need crude oil?	Evaluate. Justify using reasons and evidence for and against. *evaluate, assess, justify*
should	• Should cars be made from shape-memory metals?	Evaluate advantages and disadvantages, implications and limitations. Make a judgement. *evaluate, assess, justify, create*
might	• What might be substituted for oil products when crude oil supplies run out?	Evaluate. Justify, giving reasons for and against (using evidence, analogies and comparisons). *evaluate, assess, justify, compare, contrast, create*

Formulating a hypothesis

From the research question, a hypothesis can be developed. A hypothesis is a statement that proposes a relationship between variables, as it is based on some level of understanding. This statement *must* be testable, meaning it must specifically and clearly state a change in variables that can be tested through measurement.

The Student experiment does not necessarily require a hypothesis as it may not always be appropriate or beneficial. A hypothesis, if it is suitable, requires the controlled variables to be more stringent during the experiment, resulting in an analysis of raw data that can specifically address the original inquiry of the observation. With more stringent controls, it may be possible for errors and uncertainties to be reduced; also the interpretation of results may be more straightforward (though not always).

The writing of a hypothesis is an optional inclusion that might support the development of your research question. A hypothesis is not an assessed component of the ISMG, and not every research question will lead to the development of a hypothesis.

Scientists use **literature reviews** and background research to develop an understanding of an observation and then infer a reason for their observation. The inference is then tested using experimentation to determine if it is true (verified or supported) or false (falsified or refuted).

Because the hypothesis proposes a specific relationship between the independent and dependent variables, the hypothesis can either be supported or refuted by the results. To be able to propose a specific relationship the scientist must have some knowledge and understanding of the variables.

To develop a hypothesis, similar steps are undertaken to developing a question (Figure 1.8.5).

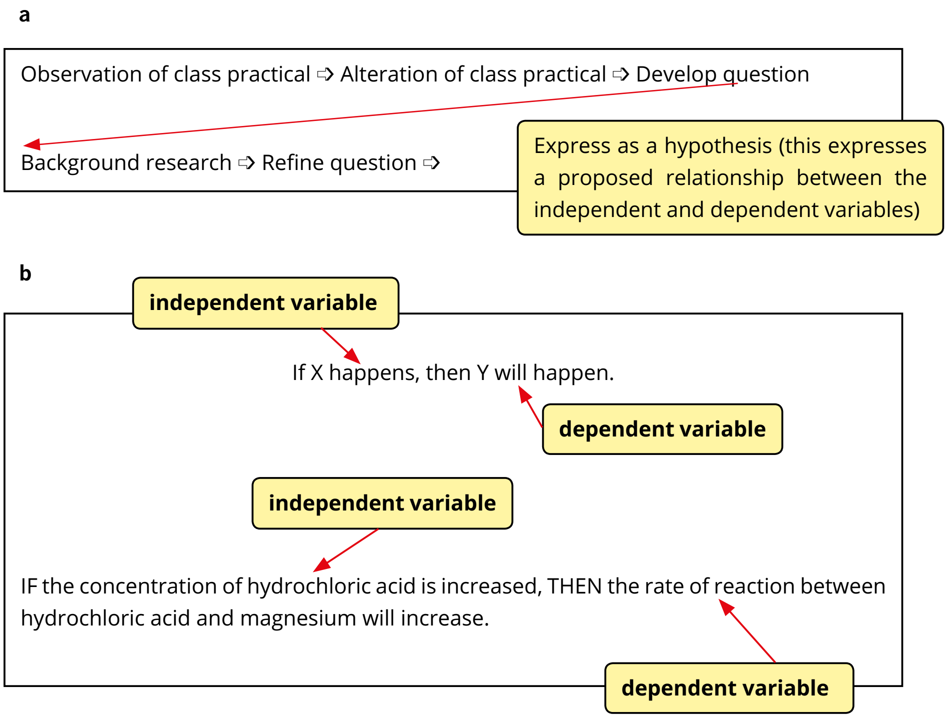


FIGURE 1.8.5 (a) Refining a hypothesis/question. (b) Formulating a hypothesis—write in terms of the dependent and independent variables.

The 'IF' part of the hypothesis refers to the independent variable—the variable you alter in the experiment. The 'THEN' part relates to the dependent variable, which is the variable you measure or observe.

A hypothesis does not need to include 'if' and 'then' in its wording. For example, the previous hypothesis could also be worded in the following way:

'The rate of the reaction between hydrochloric acid and magnesium will increase as the concentration of the hydrochloric acid is increased.'

A good hypothesis can be tested to be true (verified or supported) or false (falsified or refuted) by investigation.

A well-constructed hypothesis provides guidance and structure to the experiment in the following areas:

- methodology—provides specific limitations and guidance (due to the proposed relationship) directing the planning of the experiment; suggests specific variables to control and measure
- results—offers guidance for analysis and processing of data to determine if the proposed relationship exists; gives direction for the display of results to be specific to the hypothesised relationship between the variables
- evaluation—specifically directs what should be considered for adjustments in the methodology to improve or extend the experiment.

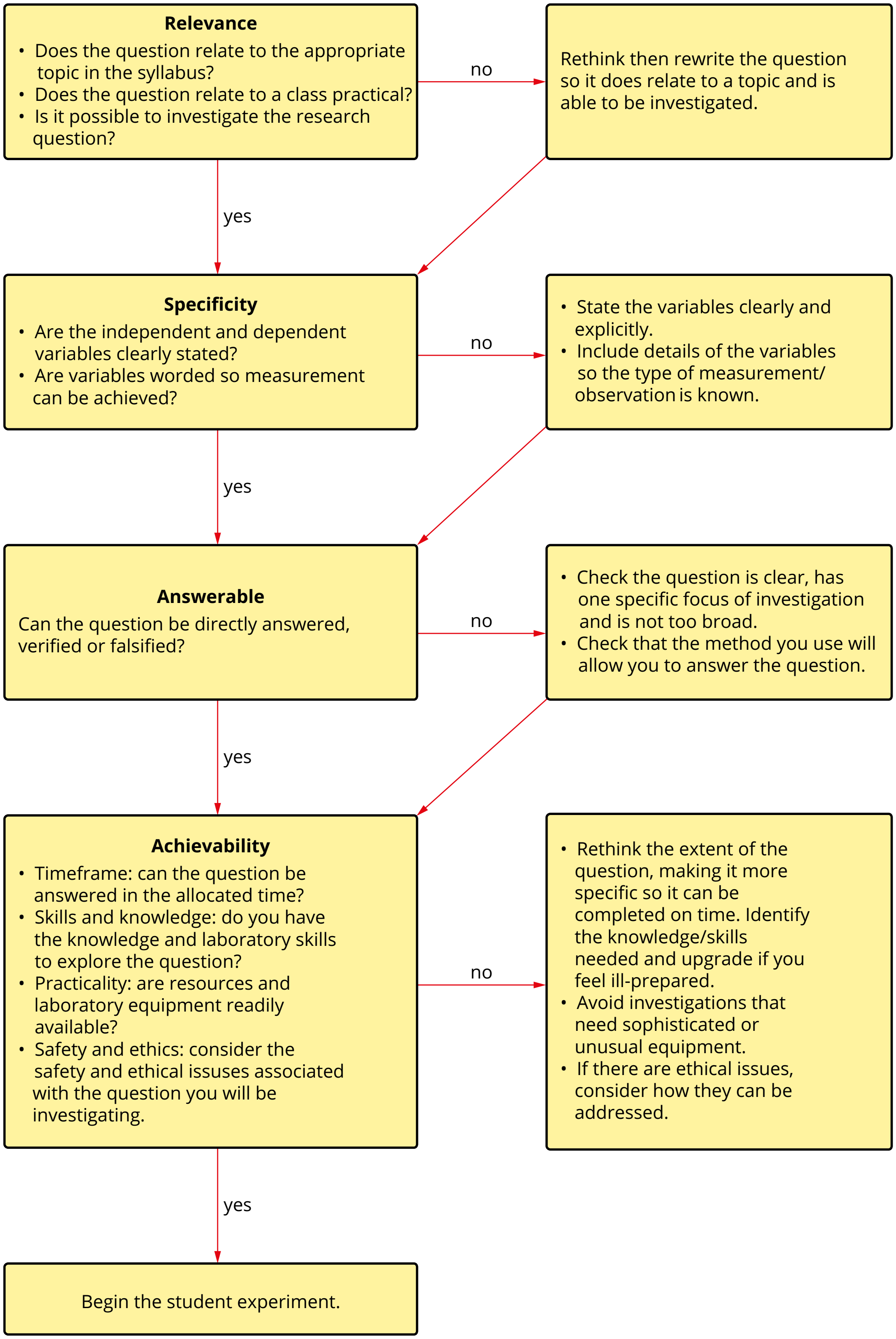
ℹ The writing of a hypothesis is an optional feature that you may wish to write to support the development of your research question.

Although a hypothesis is not an assessed component of the ISMG, some students may wish to write a hypothesis to support their development of a research question. This is an optional inclusion, since it may not always be possible to write a hypothesis for every research question.

Evaluating your research question or hypothesis

The quality of the question or hypothesis is vital to the quality of the response that can be written.

Once a research question or hypothesis has been developed, stop to evaluate it before progressing. Follow the prompts in the chart below (Figure 1.8.6) to refine and improve the research question or hypothesis.



Relevance

- Does the question relate to the appropriate topic in the syllabus?
- Does the question relate to a class practical?
- Is it possible to investigate the research question?

Rethink then rewrite the question so it does relate to a topic and is able to be investigated.

Specificity

- Are the independent and dependent variables clearly stated?
- Are variables worded so measurement can be achieved?

- State the variables clearly and explicitly.
- Include details of the variables so the type of measurement/observation is known.

Answerable

Can the question be directly answered, verified or falsified?

- Check the question is clear, has one specific focus of investigation and is not too broad.
- Check that the method you use will allow you to answer the question.

Achievability

- Timeframe: can the question be answered in the allocated time?
- Skills and knowledge: do you have the knowledge and laboratory skills to explore the question?
- Practicality: are resources and laboratory equipment readily available?
- Safety and ethics: consider the safety and ethical issuses associated with the question you will be investigating.

- Rethink the extent of the question, making it more specific so it can be completed on time. Identify the knowledge/skills needed and upgrade if you feel ill-prepared.
- Avoid investigations that need sophisticated or unusual equipment.
- If there are ethical issues, consider how they can be addressed.

Begin the student experiment.

FIGURE 1.8.6 Steps to refine and improve the hypothesis or research question

Examples of hypotheses and research questions

The following examples demonstrate that the student experiment may be expressed as either a hypothesis or a research question to be investigated.

Hypothesis: As the temperature of the gas increases, the volume of the gas increases.

Research question: How does the volume of air in a syringe change as the temperature of the air is increased from 20 to 80 degrees Celsius?

Hypothesis: The rate of reaction increases as the temperature of the reaction mixture increases.

Research question: Does the time taken for the production of 25 mL of hydrogen gas decrease as the temperature of the solution of 50 mL 0.5 M hydrochloric acid and 5 g zinc is increased?

Hypothesis: The accuracy of the molar enthalpy of combustion of ethanol can be increased by performing multiple calorimetry experiments.

Research question: Does burning different masses of ethanol in multiple calorimetry experiments improve the accuracy of the determination of the molar heat of combustion?

DEVELOPING THE RATIONALE

Once you have decided upon an experiment to modify (i.e. refine, extend or redirect), and you have created your own related hypothesis or question, a rationale needs to be developed. The **rationale** is where you explain the scientific concepts appropriate to the research question.

Research relevant scientific information

The Student experiment in the Chemistry Syllabus requires students to:

- research what is currently known about the relationship between the dependent and independent variables
- create a methodology that allows sufficient, relevant data to be collected that enables the research question to be answered
- manage the risks and issues associated with the experiment.

The student experiment ISMG (IA2) outlines the characteristics to be demonstrated by students related to the background research as:

- informed application to modify experimental methodologies demonstrated by
  - a considered rationale for the experiment
  - justified modifications to the methodology
- effective and efficient investigation demonstrated by
  - a considered methodology that enables the collection of sufficient, relevant data
  - considered management of risks and ethical or environmental issues.

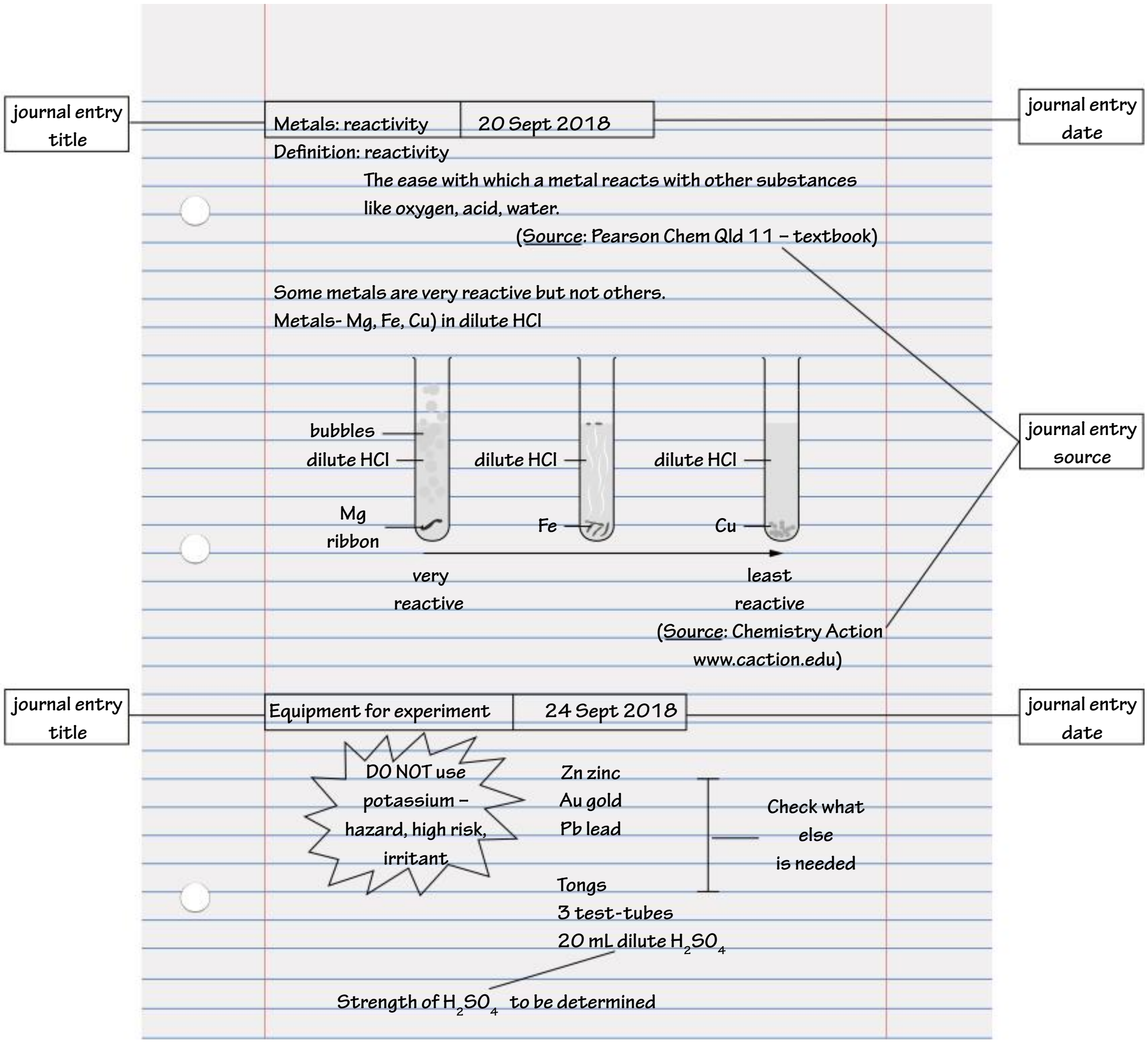
Furthermore, the Chemistry Syllabus expects background scientific information to be used in a rationale for the experiment to:

- explicitly justify the modifications to the methodology (alterations to the class practical)
- explain how the methodology will enable the research question to be answered through the collection of the data
- inform risk. This relates to identifying how the risks associated with the experiment will be reduced through use of personal protective equipment or specific features of the methodology.

The rationale is also expected to inform the interpretation of the evidence (results) and conclusion.

There is a lot of information to record during the student experiment so it would be best to keep it all in a single document. Scientists always document all their ideas, questions, background research and literature reviews, methodology drafts and revisions, results, refinements etc. in a single document. It is called a **scientific journal**.

Taking notes and recording your thoughts in the scientific journal keeps a record of all information that arose during the process of the scientific method (Figure 1.8.7). It is unknown how vital some information will become throughout the process until it is complete and summarised into the scientific report. It will be used in a rationale for the experiment; to explicitly justify the modifications to the methodology (alterations to the class practical); to provide a reason for collecting data; and to inform risk, ethical and environmental management. It is also expected to inform the interpretation of the evidence (results) and conclusion.



Metals: reactivity | 20 Sept 2018

Definition: reactivity

The ease with which a metal reacts with other substances like oxygen, acid, water.

(Source: Pearson Chem Qld 11 – textbook)

Some metals are very reactive but not others.

Metals- Mg, Fe, Cu) in dilute HCl

very reactive → least reactive

(Source: Chemistry Action www.caction.edu)

Equipment for experiment | 24 Sept 2018

DO NOT use potassium – hazard, high risk, irritant

Zn zinc
Au gold
Pb lead

Check what else is needed

Tongs
3 test-tubes
20 mL dilute H_2SO_4

Strength of H_2SO_4 to be determined

FIGURE 1.8.7 A journal entry with the journal entry date, titles and apparatus used in preparation for an experiment. There will be numerous journal entries on any given date, and they should have their own titles. Also, the same titles will appear on different entry dates as research continues on specific topics.

The purpose of researching background scientific information in the scientific method is to develop understanding and knowledge. It must be relevant to the independent and dependent variables in the research question. As the variables become known, this will direct the research. The scientific method specifically requires scientists to demonstrate understanding and knowledge of the direct, and possibly indirect, relationship between the independent and dependent variables and perhaps controlled or measured variables. The background research is essential to achieve all this.

PLANNING AND REFINING METHODOLOGY

This section is a guide to some of the key steps that should be taken when planning and refining an investigation.

Planning experiments

Once you have formulated your research question or hypothesis, defined the variables, and developed knowledge and understanding of concepts and relationships, you will need to develop your experiment. You will also need to consider the ethical and safety implications of the testing during the experiment.

Create a work schedule that outlines the time frame of your experiment (including all trials and/or samples), being sure to include sufficient time to repeat experiments if necessary. Check with your teacher that your protocol and schedule (methodology) are appropriate, and that others will be able to repeat your experiment exactly by following the methodology you have written. If you have planned well you will be able to test your methodology and run trials. You need to be able to perform your experiment independently, in the time available in the school laboratory, and with minimal support from your teachers and school laboratory staff.

The methodology of your experiment is a specific step-by-step procedure. Though, when written in the final scientific report, it may be written in paragraph form. You must ensure that the methodology is valid, specific, reliable and accurate. All of these factors need to be considered when planning.

Validity

Validity refers to whether an experiment or investigation is in fact testing the set research question or hypothesis. Is the experiment obtaining data that is relevant to the research question?

Factors influencing validity include:

- whether your experiment measures what it **claims** to measure. In other words, your experiment should test your hypothesis and/or research question.
- the certainty that something observed in your experiment was the result of your experimental conditions and not some other cause that you did not consider. In other words, whether the independent variable influenced the dependent variable in the way you have concluded.
- the degree, or scope, to which your findings can be generalised to more broad chemistry phenomena from which your experiment is taken, i.e. how does your experiment fit in to the big ideas in chemistry.

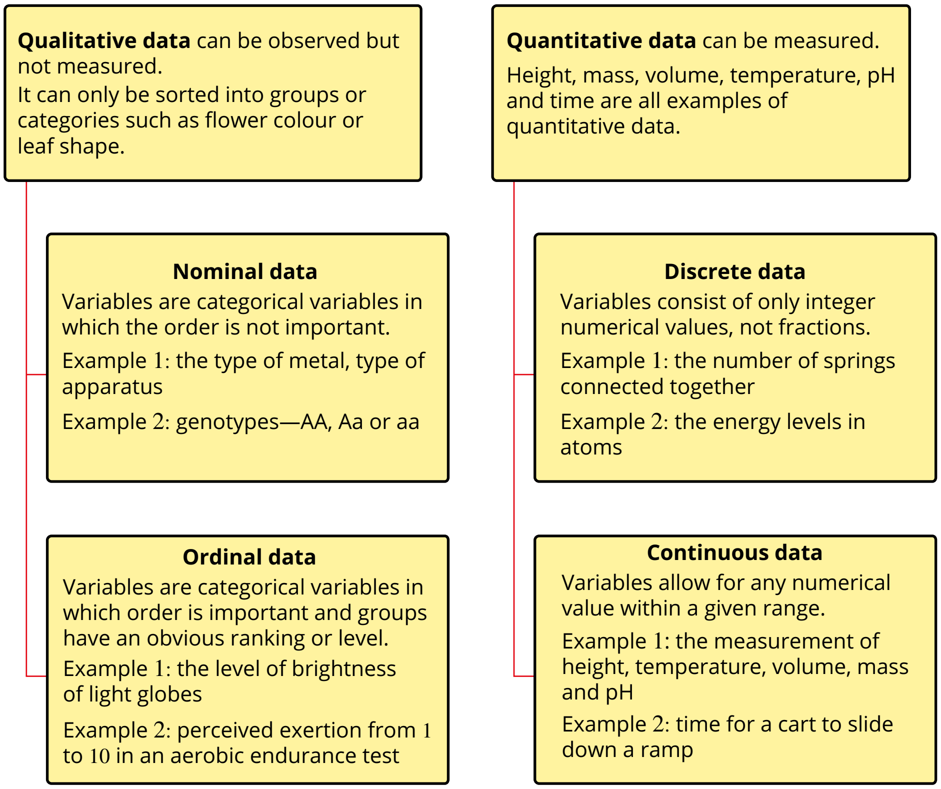
To ensure an investigation is valid, it should be designed so that only one variable is being changed at a time. The remaining variables must remain constant so that meaningful conclusions can be drawn about the relationship between variables.

Also, the raw data collected during the experiment must be appropriate to ensure the data is valid. To ensure validity, carefully determine:

- the independent variable
- the dependent variable
- the controlled and/or measured variables
- the appropriate raw data that will be collected (quantitative versus qualitative), and that each will be measured, collected or controlled appropriately.

Data can either be qualitative or quantitative. **Qualitative data** is descriptive and unmeasurable and uses descriptions or adjectives to record observations. **Quantitative data** is empirically measurable and uses instruments to record observations as numbers with units.

Qualitative and quantitative data have further subsets in each category (Figure 1.8.8 on page e66).



Qualitative data can be observed but not measured.
It can only be sorted into groups or categories such as flower colour or leaf shape.

Nominal data
Variables are categorical variables in which the order is not important.
Example 1: the type of metal, type of apparatus
Example 2: genotypes—AA, Aa or aa

Ordinal data
Variables are categorical variables in which order is important and groups have an obvious ranking or level.
Example 1: the level of brightness of light globes
Example 2: perceived exertion from 1 to 10 in an aerobic endurance test

Quantitative data can be measured.
Height, mass, volume, temperature, pH and time are all examples of quantitative data.

Discrete data
Variables consist of only integer numerical values, not fractions.
Example 1: the number of springs connected together
Example 2: the energy levels in atoms

Continuous data
Variables allow for any numerical value within a given range.
Example 1: the measurement of height, temperature, volume, mass and pH
Example 2: time for a cart to slide down a ramp

FIGURE 1.8.8 Quantitative and qualitative variables

Measurement of temperature requires an instrument and provides quantitative data (Figure 1.8.9).

FIGURE 1.8.9 (a) Using a thermometer or a wireless temperature sensor measures temperature empirically and provides quantitative data that can be analysed statistically. Processing empirical data can produce discrete, explicit and comparative analysis. (b) A liquid thermometer for measuring water temperature. (c) Feeling heat radiating from a heater is an example of qualitative data as it is based on personal observation. Qualitative data cannot be statistically analysed.

Depending on the experiment, it may be appropriate to record the qualitative observation (e.g. pattern of turtle shell), or perhaps measure the quantitative results (e.g. blood glucose level). See Figure 1.8.10.

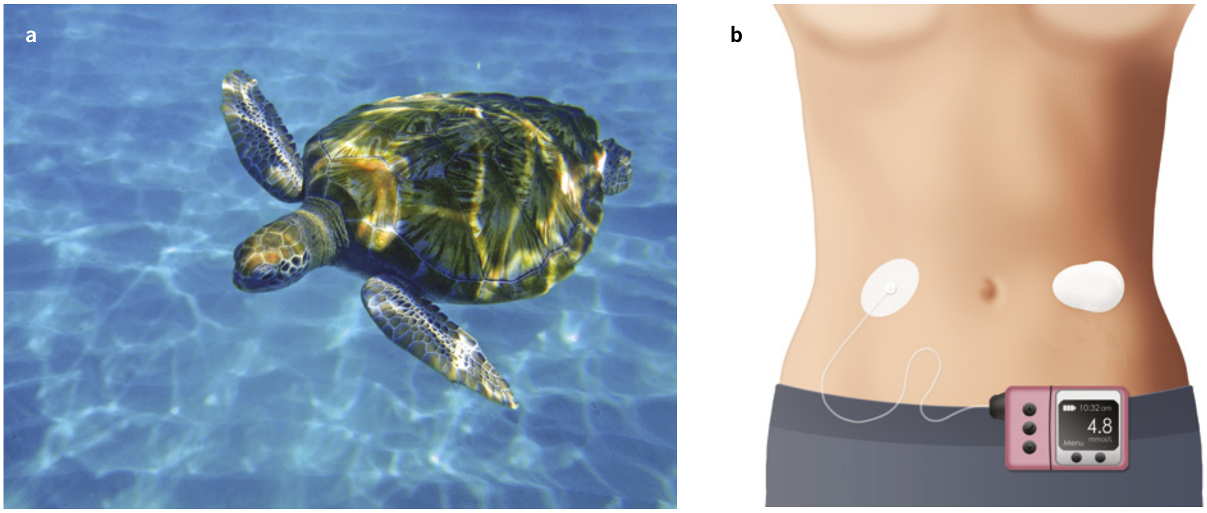


FIGURE 1.8.10 (a) When recording qualitative data, describe in detail how each variable will be defined. For example, recording the colours and different patterns on the top of a turtle's shell can help distinguish one species from another. (b) This glucose sensor samples blood glucose and sends a signal to the insulin pump, which releases insulin when needed. This is an example of a continuous variable that is quantitative data.

Controls

Controls are required in an experiment. Controlling as many variables as possible to determine what has influenced the results goes a long way to providing accurate and precise data.

It is difficult—sometimes impossible—to eliminate all **variables** that might affect the outcome of an experiment. Such variables in chemistry include concentration of solution, temperature, mass of reactants, pH and surface area of reactants. (A way to eliminate the possibility that random factors could affect results and cause uncertainty is discussed in Module 1.5.) A methodology to overcome this is to set up a second group within the experiment (called a **control group**) that is identical in every way to the first group (the **experimental group**) except for the single experimental (independent) variable that is being tested. This is a controlled experiment. This allows the examination of one variable at a time (the independent variable), which is required to validly test a hypothesis.

Reliability

Reliability refers to the notion that if the experiment is repeated many times, the results obtained should be consistent. Reliability (sometimes called repeatability) is the ability to obtain the same results if an experiment is repeated (Figure 1.8.11). The closer the results are to the true value, the more reliable (and accurate) it will be. Because a single measurement or experimental result could be affected by errors, **replication** of samples within an experiment and **repeat trials** are key components of reliability. To improve reliability you should:

- specify the materials and methods in detail (including precision, see Module 1.5 Uncertainties in measurement and error)
- include replicate (several) samples within each experiment or several observations within an investigation
- take repeat readings of each sample
- run the experiment or trial more than once.

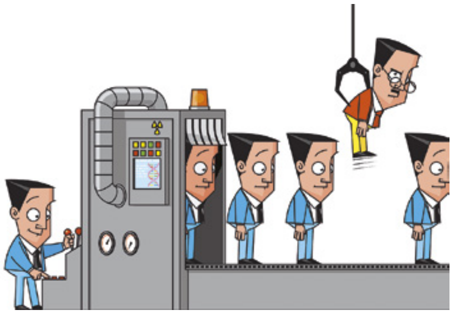


FIGURE 1.8.11 If you can reproduce your results, they are reliable.

Sample size is extremely important in scientific experiments. The sample size:

- affects the representation of the phenomenon
- affects the natural variation, errors and uncertainty
- affects the results by offering more evidence to support the experimental results
- aids with repeatability and therefore reliability.

The greater the sample size, the more reliable the data. Reasons why a measurement or observation could vary include:

- natural variation
- human error
- uncalibrated instruments or instrumental error
- influence from unforeseen variables. Accuracy and precision are also important in obtaining reliable (repeatable) data (as discussed in Module 1.5).

Sourcing appropriate equipment and materials

You will need to decide on the materials, technology and instrumentation that will be used to carry out your experiment or investigation. It is important to find the right balance between items that are easily accessible and those that will obtain accurate and precise results. When conducting your investigation, the precision of the chosen instrumentation and how this affects the accuracy and validity of the results will have to be recorded in the journal for the scientific report.

ELECTRONIC DATA ACQUISITION IN SCHOOLS

Data collection has come a long way in the past 20 years when discrete measurements were recorded from manual tools and human readings were made over short periods of time. In recent times, the ability to connect independent sensors to any device has enabled simple and highly effective electronic data acquisition over extensive periods of time, anywhere!

Data acquisition in schools is based around sensors, or probes, recording data to a standalone device, computer, tablet or even a phone (Figures 1.8.12 and 1.8.13). The software will generally graph and allow advanced calculations directly without the need for an additional program. This specialised software allows for many analysis functions not found in standard spreadsheets—and follows scientific rather than business standards.

FIGURE 1.8.12 Data acquisition software produces real-time graphs that can be downloaded or printed.

You will use data acquisition in many investigations in your Chemistry course. With the many different probes available, there are many applications.

Many smartwatches and apps in phones include digital sensors and **data loggers** that can be accessed and used for some class practicals and experiments.

FIGURE 1.8.13 Many smartwatches and phones have digital sensors and in-built data loggers.

Electronic data acquisition and data logging

Electronic data acquisition takes advantage of highly accurate sensors to collect data directly for processing by a computing device (Figure 1.8.14). There are many sensors, probes and instruments available that can measure a vast variety of phenomena in a single device.

Often one of the measurements is time; the user decides what the other measurement (or variable) is by selecting the type of sensor, or probe, to use. Probes can measure temperature, pH, gas pressure, conductivity, concentration of gases in air, or the amount of pollutants in the environment along with many more phenomena.

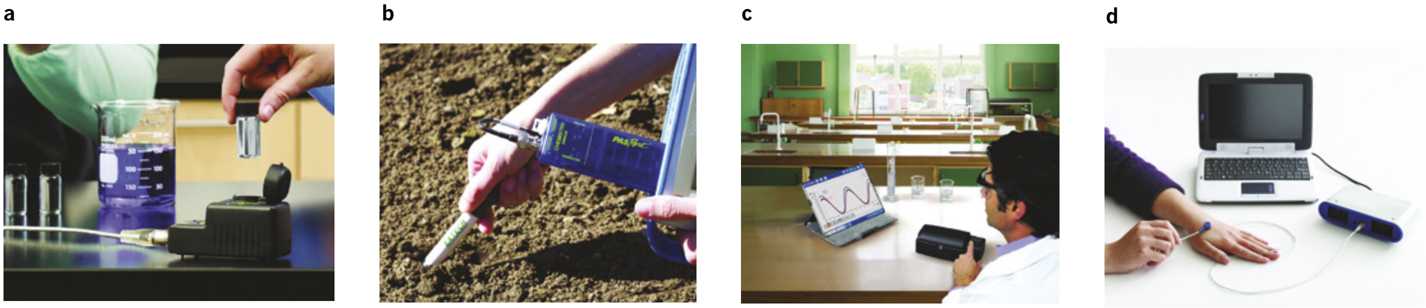


FIGURE 1.8.14 (a) Sensors connected to computing devices such as phones and tablets are readily used in the field or lab to measure the concentration of solutions and other key physical measurements. (b) Many different probes and sensors can attach to a single device, such as those being used for measurement of soil pH here. (c) They can be very precise and measure many times per second, such as the polarimeter being used here to investigate the concentration of sugar solutions. (d) They are usually easy to use and produce results immediately, for example, providing the temperature of skin upon contact.

Recording is done accurately and can be continuous or manual, and the measurements are saved electronically. The saved measurements can then be accessed via a computer or directly from the tablet screen.

The rate at which measurements are recorded is called the sample rate. Data capture rate can be varied from 100 000 times per second through to once an hour. The data capture rate selected will depend on what the operator needs. Sometimes, data may need to be captured over many days, weeks, months or even years. In these cases, time and date information are automatically recorded.

Uses of electronic data acquisition

Given the number of different probes available, electronic data acquisition has many uses (Figure 1.8.15). For example, probes can monitor the temperature of ice cream being transported from a factory to the supermarket. They are used to record the speed, engine conditions and brake force of trains, buses and trucks. An engine control unit in a car can record data from parts of the engine, and the mechanic can access this data when servicing the car. The control unit records any problems with the car's engine and computer system.

FIGURE 1.8.15 Data acquisition is used throughout many industries; for example, the automobile industry. This image shows a data logger connected to a steering wheel. The collected data provides informative feedback to manufacturers.

RISK ASSESSMENT

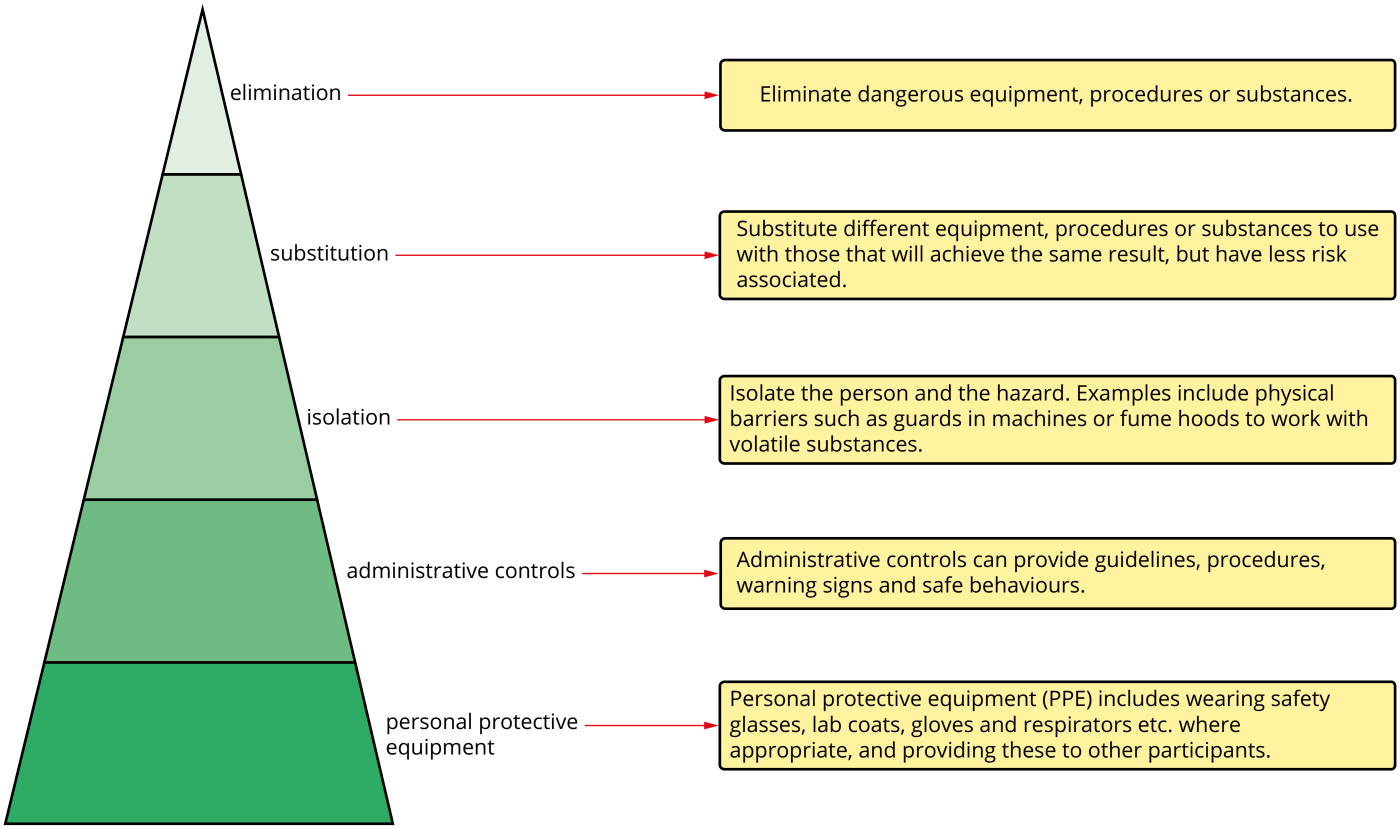
While planning for an experiment, or investigation in the laboratory or outside in the field, it is important for your safety and the safety of others that you consider the potential risks.

Everything you do has some risk involved. **Risk assessments** are performed to identify, assess and control hazards. A risk assessment should be performed for any situation, whether in the laboratory or out in the field, which could cause harm to people, property or animals. Always identify the risks and control them to keep everyone safe.

To identify risks, think about the following:

- the activity that you will be carrying out
- where in the environment you will be working, e.g. in a laboratory, school grounds, or a natural environment
- how you will use any equipment or chemicals that you will be handling, including sources of heat and electricity
- what clothing you should wear, e.g. lab coat and goggles.

The hierarchy of risk controls is shown in Figure 1.8.16 on page e71. It is organised from the most effective to least effective. The most commonly used risk control measure that addresses the majority of risks is personal protective equipment. The least common but most protective control measure is eliminating the risk from the scientific investigation.



Eliminate dangerous equipment, procedures or substances.

Substitute different equipment, procedures or substances to use with those that will achieve the same result, but have less risk associated.

Isolate the person and the hazard. Examples include physical barriers such as guards in machines or fume hoods to work with volatile substances.

Administrative controls can provide guidelines, procedures, warning signs and safe behaviours.

Personal protective equipment (PPE) includes wearing safety glasses, lab coats, gloves and respirators etc. where appropriate, and providing these to other participants.

FIGURE 1.8.16 The hierarchy for risk control is shown in this pyramid, marked from bottom to top in order of increasing importance.

Personal protective equipment

Everyone who works in a laboratory wears items that help keep them safe, such as those shown in Figure 1.8.17. These items are called **personal protective equipment (PPE)** and include:

- safety glasses
- shoes with covered tops
- disposable gloves when handling certain chemicals
- a disposable apron or a lab coat if there is risk of damage to clothing.

FIGURE 1.8.17 Examples of PPE shown are protective eye wear, lab coats and gloves.

Chemical codes

The chemicals at school or in a hardware shop have a warning symbol on the label. These warnings are called **chemical (HAZCHEM) codes**. In 2017 the Globally Harmonised System of Classification and Labelling of Chemicals (GHS) was introduced in Australia for use in workplaces, including school laboratories. Some common pictogram codes and their meanings are shown in Table 1.8.3.

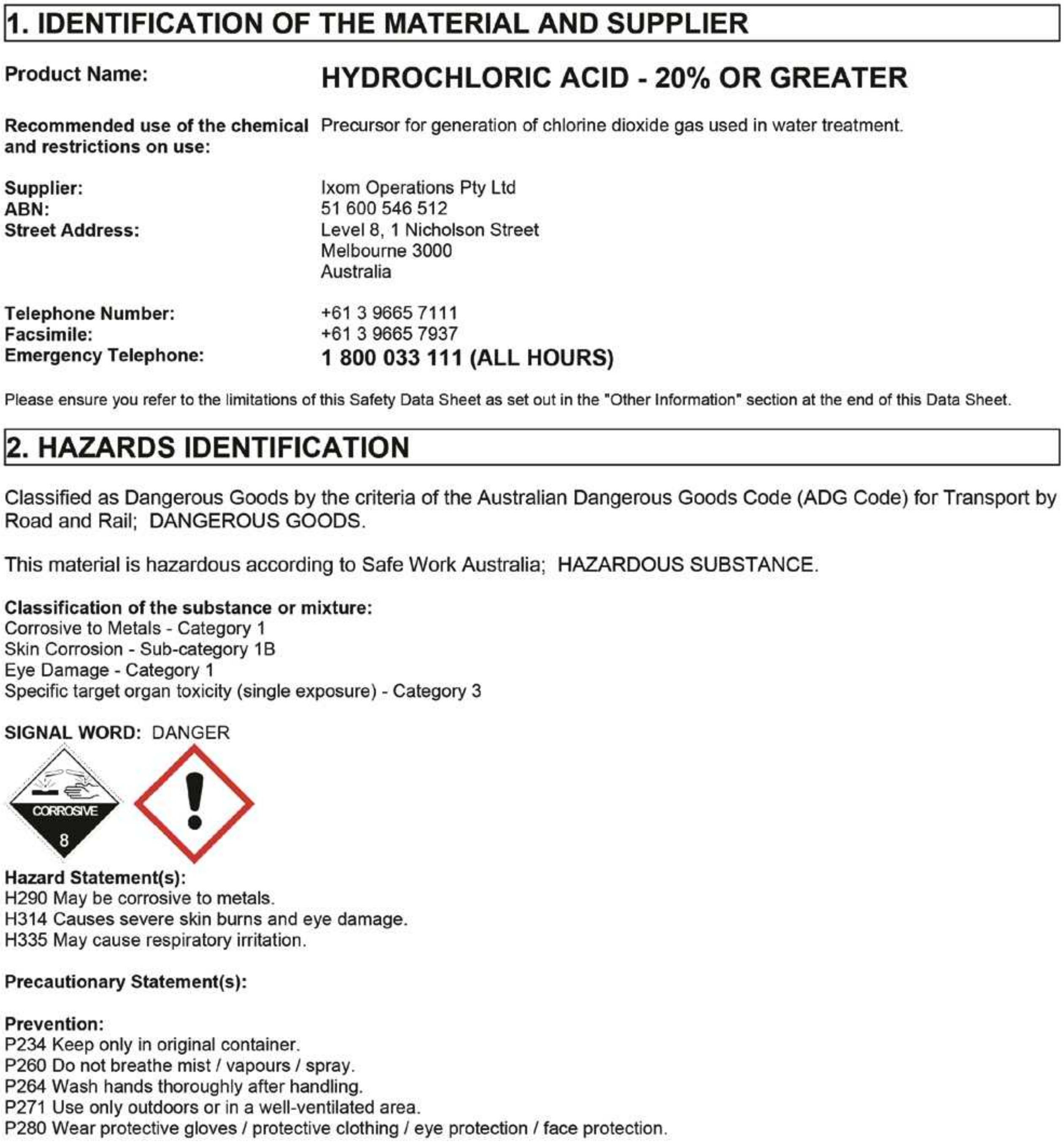
TABLE 1.8.3 Common HAZCHEM codes and their meanings

GHS Symbol	Meaning
	Corrosive: can dissolve or eat away at substances, including tissues such as your skin or airways
	Poison: can cause injury or death if ingested, inhaled or absorbed
	Irritant: causes discomfort, pain or itchiness
	Flammable
	Danger: the danger can include any biological harm, for example, cancer, allergy, breathing difficulties etc.

Safety Data Sheets

Each chemical substance has an accompanying document called a **Safety Data Sheet (SDS)** (Figure 1.8.18). An SDS contains important safety and first aid information about each chemical you commonly use in the laboratory. If the products of a reaction are toxic to the environment, you must pour your waste into a special container—*not* down the sink.

The SDS provides employers, workers and health and safety representatives with the necessary information to safely manage the risk of hazardous substance exposure.



1. IDENTIFICATION OF THE MATERIAL AND SUPPLIER

Product Name: **HYDROCHLORIC ACID - 20% OR GREATER**

Recommended use of the chemical and restrictions on use: Precursor for generation of chlorine dioxide gas used in water treatment.

Supplier: Ixom Operations Pty Ltd
ABN: 51 600 546 512
Street Address: Level 8, 1 Nicholson Street
Melbourne 3000
Australia

Telephone Number: +61 3 9665 7111
Facsimile: +61 3 9665 7937
Emergency Telephone: **1 800 033 111 (ALL HOURS)**

Please ensure you refer to the limitations of this Safety Data Sheet as set out in the "Other Information" section at the end of this Data Sheet.

2. HAZARDS IDENTIFICATION

Classified as Dangerous Goods by the criteria of the Australian Dangerous Goods Code (ADG Code) for Transport by Road and Rail; DANGEROUS GOODS.

This material is hazardous according to Safe Work Australia; HAZARDOUS SUBSTANCE.

Classification of the substance or mixture:
Corrosive to Metals - Category 1
Skin Corrosion - Sub-category 1B
Eye Damage - Category 1
Specific target organ toxicity (single exposure) - Category 3

SIGNAL WORD: DANGER

Hazard Statement(s):
H290 May be corrosive to metals.
H314 Causes severe skin burns and eye damage.
H335 May cause respiratory irritation.

Precautionary Statement(s):

Prevention:
P234 Keep only in original container.
P260 Do not breathe mist / vapours / spray.
P264 Wash hands thoroughly after handling.
P271 Use only outdoors or in a well-ventilated area.
P280 Wear protective gloves / protective clothing / eye protection / face protection.

FIGURE 1.8.18 Extracts of a Safety Data Sheet (SDS) for concentrated hydrochloric acid. The SDS alerts the reader to any potential hazards when using a substance, including appropriate measures to reduce risk of harm.

Ethical considerations

When planning an investigation, researchers should always identify all possible ethical considerations and evaluate their necessity or ways that can reduce or mitigate them. **Ethics** is a set of moral principles by which your actions can be judged as right or wrong. Every society or group of people has its own principles or rules of conduct. Scientists have to obtain approval from an ethics committee and follow ethical guidelines when conducting research that involves animals or, especially, humans.

These are some of the ethical issues that must be considered.

- How can this affect wider society?
- Does one party benefit over another; for example, one individual, a group of individuals or a community?
- Is there a risk of harm (physical or mental) to people involved in the research?
- Does it prevent anyone from gaining their basic needs?
- Will this have consequences for future ethical decisions or issues?
- Does the research cause damage to the environment or other living things?

In reality, school chemistry investigations generally will have minor ethical issues, if any, but these should be considered in your planning.

Refining the methodology

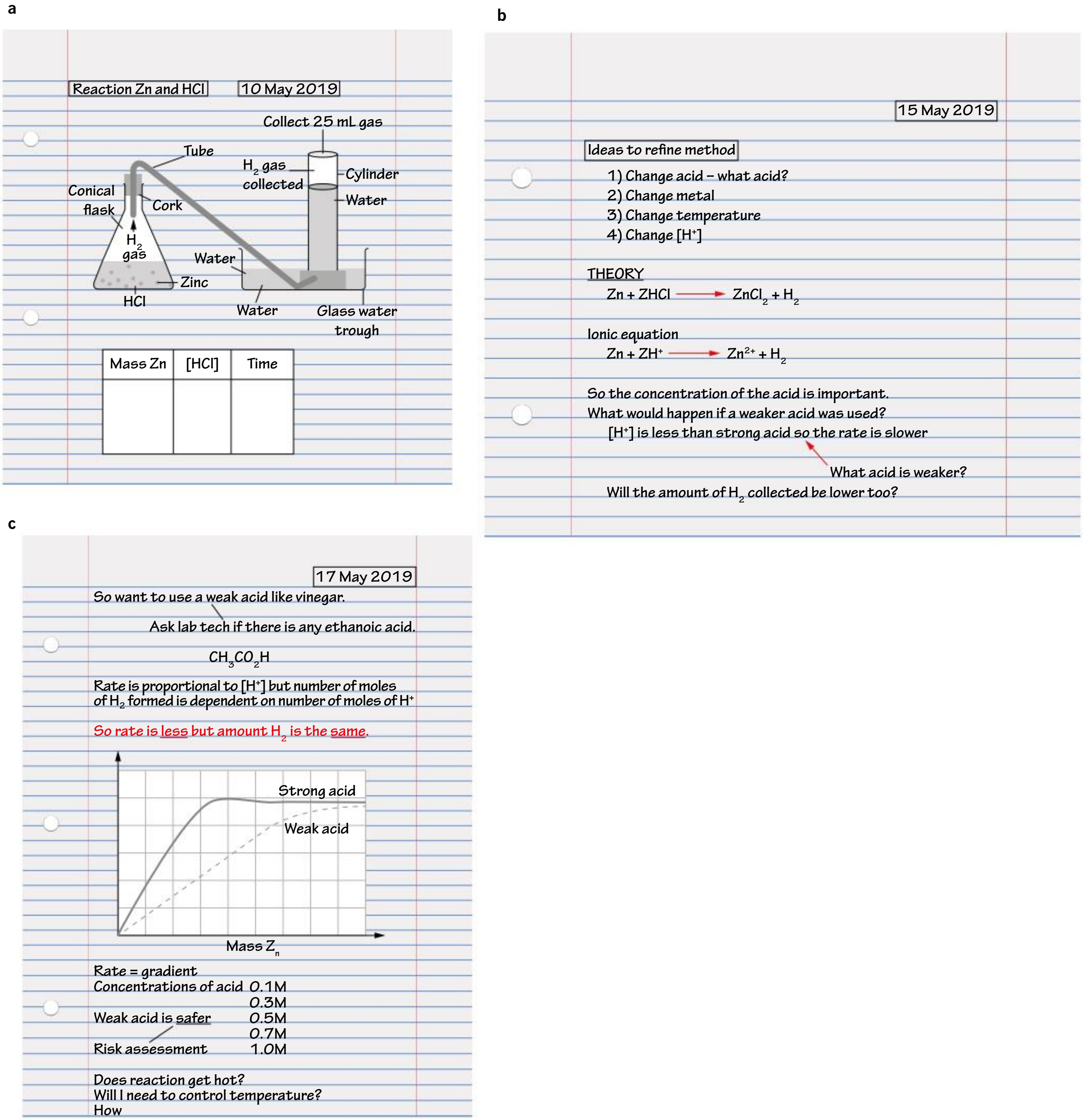
As the planning of the methodology is not linear, refinement will occur several times (due to further background research, refining the research question or as variables become understood). Scientists will employ an experimental methodology that has been refined numerous times over several years.

ℹ It is common for experimentation and testing not to go according to plan. It is vital that comprehensive background research has been undertaken. Refinements are often made during experimentation to improve validity and reliability. This may be due to time constraints, instrumental limitations or resource limitation etc. Refinements also reduce error and uncertainty as the experimenter becomes aware of these problems.

Record all refinements in the journal as shown in Figure 1.8.19 on page e75. The following may help with refining the methodology.

- Record everything.
- Be prepared to make changes and refinements to the plan and methodology.
- Note any difficulties encountered and the ways they were overcome. What were the failures and successes?

Every test carried out can contribute to the understanding of the investigation as a whole, no matter how much of a disaster it may first appear.



Reaction Zn and HCl
10 May 2019

Mass Zn	[HCl]	Time

15 May 2019

Ideas to refine method

1) Change acid – what acid?
2) Change metal
3) Change temperature
4) Change $[H^+]$

THEORY
$Zn + ZHCl \longrightarrow ZnCl_2 + H_2$

Ionic equation
$Zn + ZH^+ \longrightarrow Zn^{2+} + H_2$

So the concentration of the acid is important.
What would happen if a weaker acid was used?
$[H^+]$ is less than strong acid so the rate is slower
What acid is weaker?
Will the amount of H_2 collected be lower too?

17 May 2019
So want to use a weak acid like vinegar.
Ask lab tech if there is any ethanoic acid.
CH_3CO_2H
Rate is proportional to $[H^+]$ but number of moles of H_2 formed is dependent on number of moles of H^+
So rate is less but amount H_2 is the same.

Rate = gradient
Concentrations of acid 0.1M
0.3M
Weak acid is safer 0.5M
0.7M
Risk assessment 1.0M

Does reaction get hot?
Will I need to control temperature?
How

FIGURE 1.8.19 Student journal demonstrating basic developments and refinements in their methodology over several days

If the expected data is not obtained, do not worry. As long as it can be critically and objectively evaluated, the limitations of the investigation can be identified and further investigations proposed, the work is worthwhile. Sometimes a different point of view, such as from a fellow student or a teacher, is all that is needed to find a solution. An evaluation and suggested improvement to the methodology or experiment is required in the scientific report. This is discussed further in Module 1.10.

1.8 Review

SUMMARY

- An independent variable is a variable that is controlled by the researcher (the variable that is selected and changed).
- A dependent variable is a variable that may change in response to a change in the independent variable, and is measured or observed.
- Controlled variables are the variables that are kept constant during the investigation.
- Measured variables are the variables in an open system that cannot be controlled and therefore must be measured.
- Research questions should:
  - include measurable variables (the independent and dependent variables)
  - have a guiding word, such as *who*, *what*, *why* or *will*
  - be phrased so that a definitive answer can be developed
  - be able to link the guiding word to command verbs (such as *identify*, *describe*, *compare*, *contrast*, *distinguish*, *analyse*, *evaluate* or *create*) so that a task can be determined.
- A simple way to formulate a hypothesis is to link the independent and dependent variables using the following sentence structure:

  If [independent variable] happens, then [dependent variable] will happen.
- A scientific journal is what a scientist uses to record all their ideas, questions, background research and literature reviews, methodology drafts and revisions, results and refinements related to an experiment.
- Validity refers to whether an experiment or investigation is in fact testing the set research question or hypothesis.
- Data can either be qualitative or quantitative.
- Qualitative data is descriptive and unmeasurable and uses descriptions or adjectives to record observations.
- Qualitative data can be characterised as either:
  - nominal, when the order of data is not important
  - ordinal, when the ordering of data is important.
- Quantitative data is empirically measurable and uses instruments to record observations.
- Quantitative data can be characterised as either:
  - discrete, when data can only be recorded as particular numerical values
  - continuous, when data is not restricted to particular numerical values, but occurs within a given range.
- Reliability refers to the notion that if the experiment is repeated many times, the results obtained should be consistent.
- Reliability is improved by:
  - replication, having multiple samples within an experiment
  - repeat trials, repeating the experimental test.
- Risk assessments identify, assess and control hazards.
- HAZCHEM GHS pictograms are warning images used to identify hazardous substances.

KEY QUESTIONS

Retrieval

1 **a** State the meaning of the term ‘variable’.

b Define the types of variables in the table below.

Independent variable	Controlled variables	Dependent variable

Comprehension

2 Explain the difference between accuracy and precision.

3 Explain the difference between quantitative and qualitative data.

Analysis

4 A student wanted to find out the pH at which mussel shells begin to dissolve.

  a Identify the independent variable for the experiment.

  b Identify the dependent variable.

  c Deduce an appropriate way of measuring quantitatively the dissolution of mussel shells.

  d Identify the variables that would need to be controlled in the student's experiment.

5 A student is investigating the effect of changing the concentration of potassium hydroxide on the time taken to produce 25 mL of hydrogen gas by electrolysis.

Identify a methodology that will enable the collection of sufficient data, including identifying the dependent and independent variables.

1.9 Conducting and experimenting

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- determine relevant data that is needed to test a hypothesis
- determine what is considered to be sufficient data to test an hypothesis
- select appropriate equipment to collect relevant and sufficient data.

Now that much is known about the variables involved and the experiment methodology is planned, the experiment can be conducted. Since experiments rarely run according to the precise plan, understanding the variables involved enables valid changes or refinements to the experiment to be made.

CONDUCTING THE EXPERIMENT

When conducting the experiment, variables must be controlled and conditions maintained as well as possible to ensure the measured or recorded raw data is valid and reliable. Other scientists or investigators must be able to repeat your results by collecting and measuring raw data under the same conditions.

The precision of the instruments used to measure and record data is important. It determines the significant figures when analysing the data, it may affect the accuracy of the experiment and it could provide assistance in producing reliable results. The more precise the instrument, the more accurate the measurement and the more reliable the data is.

Considerations when conducting an experiment

Depending on what the student experiment is testing, there are several aspects of the experiment that should be included in the planning, such as:

- equipment
- instruments
- safety precautions
- time (preparation, testing)
- complexity of testing
- chemical concentrations and volumes
- sequential order of activities to complete testing.

Equipment

Choice of equipment will influence the reliability of the experiment. Therefore, it is important to choose equipment wisely. It needs to provide the required precision but also be something that you are capable of using correctly.

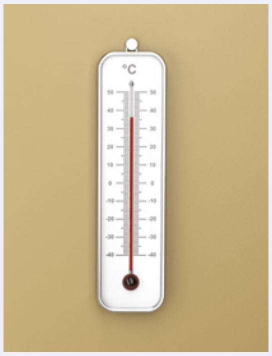
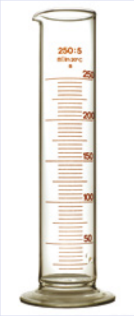
Use the equipment that is best suited to the data that needs to be collected to validate the hypothesis. Determining the units of the data being collected and at what scale will help to select the correct equipment (Table 1.9.1). Using the correct unit and scale will ensure that measurements are more accurate and precise (with smaller systematic errors).

TABLE 1.9.1 Outlining examples of rudimentary versus improved experimental testing

Rudimentary set-up	Improved set-up
estimating the temperature of a liquid by hand 	measuring the temperature using a digital thermometer (sensor) with known precision and uncertainty 
measuring time using a 5-minute sand timer 	measuring time using a digital timer with known precision and uncertainty 

When conducting the experiment, it is recommended to use the most precise instruments available. With higher-precision instruments, there is less chance of error and lower uncertainty in the measurement. Table 1.9.2 on page e80 demonstrates the benefits of greater-precision instruments through lower uncertainty.

TABLE 1.9.2 The difference in precision between instruments: (a) a glass thermometer with a precision of ±2.5°C; (b) a measuring beaker with a precision of ±5 mL; (c) a digital laser thermometer with a precision of ±0.2°C; and (d) a measuring cylinder with a precision of ±0.25 mL

Lower-precision instruments	Higher-precision instruments
a 	**c** 
b 	**d** 

Safety precautions

Always employ safe procedures. The use of common sense during activities is essential. For example all glass equipment and instruments should be used at the back of the bench so students walking by do not cause an accident. Place a sign on the lab bench warning other students and staff not to touch the equipment.

The school's and teacher's safety and risk assessment guidelines need to be adhered to. Completing the risk assessment may require completing a form or completing an online process.

COLLECT SUFFICIENT AND RELEVANT DATA

It is important that the instruments and methodology used will enable you to measure and collect the data. The data must be relevant to the variables in the research question or hypothesis. The collected and measured data must be relevant to the proposed relationship in the research question or hypothesis. Also, enough data, enough sampling and appropriate accuracy and precision are required when collecting or measuring data. If this is not achieved, the analysis and interpretation of the data will not be reliable or valid in relation to the research question or hypothesis.

Collection of sufficient data

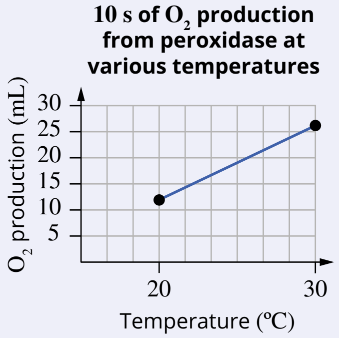
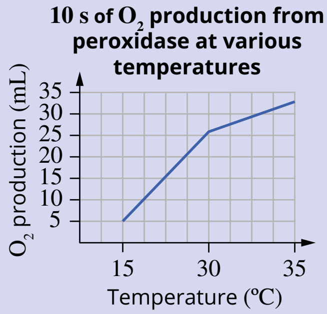
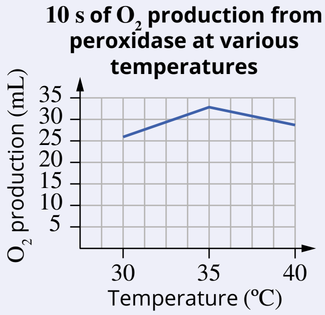
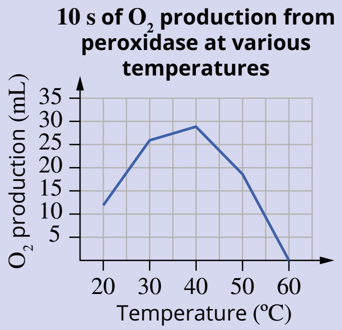
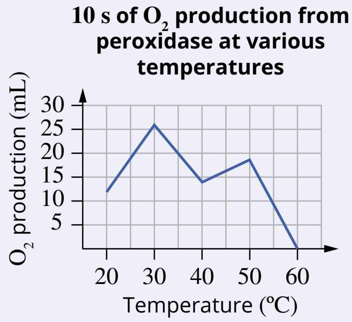
The term 'sufficient' is defined by the QCAA Chemistry General Senior Syllabus as *enough or adequate for the purpose.*

Enough data needs to be collected to substantiate whether or not a relationship exists between the variables. This includes collecting an appropriate number of replicates and also an appropriate number of individual samples (also known as observational or collection points).

This will impact on the accuracy of the data and the uncertainty within the data collected. This is vital to determine a valid interpretation of the data. In general, a minimum of five data points and three measurements per data point are needed to improve the accuracy of the individual data points.

Table 1.9.3 below outlines examples of sample size and their effect on the results of an experiment.

TABLE 1.9.3 Examples of sample size and their effect on an experiment

Example of analysed data	Effect of sample size on analysis
	A line graph with two sample points: With only two individual sample points, these results suggest an inappropriate relationship between temperature and enzyme activity; that enzyme activity increases with temperature.
	Line graph with three sample points using an uneven scale: The spread between the chosen individual sample points is inappropriate for this experiment as they should be evenly spread. This example produces results with an incorrect relationship between temperature and enzyme activity; that enzyme activity increases then begins to level off as temperature increases.
	Line graph showing limited sample points: The chosen number of individual sample points is appropriate. This may be sufficient; however, the next example provides more precision and validity for the relationship.
	Scatterplot with appropriate number and distribution of sample points: The number of individual sample points here best demonstrates the relationship between the variables.
	Line graph based on data that had no replicates: These results are due to only a single trial for each individual sample point (distance) without any replicates. In this example, the measurement at 40°C was a mistake, and without replicates these mistakes will influence the results.

Collection of relevant data

The variables to be measured or collected must be directly related to the proposed independent–dependent variable relationship. Additional variables can be measured or collected that are indirectly related to the hypothesised relationship if the background research shows it could be beneficial in the analysis or interpretation of the relationship. If you do not have any background research relating a variable to the research question or hypothesis, then it is not relevant, and therefore is not to be measured or collected.

1.9 Review

SUMMARY

- The choice of equipment and instruments will influence the reliability of the experiment.
- The precision of equipment and instruments is important for accuracy and reliability.
- The collected and measured data must be relevant to the proposed relationship in the research question or hypothesis.
- Equipment used during an experiment should enable you to collect and measure relevant data.

KEY QUESTIONS

Retrieval

1 Recall why it is important to choose appropriate equipment and instruments to conduct experiments.

Comprehension

2 Explain how the precision of equipment can affect scientific conclusions.

Analysis

3 In an experiment to determine the how the rate of a reaction varied with temperature, a student recorded the following data. Determine whether the data collected is sufficient and relevant.

Experiment no.	Temperature (°C)	Starting mass of Zn (g)	Concentration of HCl (M)	Time to collect 25 mL of hydrogen (s)
1	20	0.70	0.1	80
2	25	0.83	0.1	76
3	25	1.04	0.2	71
4	30	0.59	0.2	68
5	40	1.2	0.1	56

1.10 Results

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- analyse raw data to produce processed data
- interpret data to draw valid conclusions.

All measurements and observations made during the experiment must be recorded (in the journal). This is the **raw data**. Choosing not to record certain measurements or observations (raw data) is invalid, shows bias and is scientifically fraudulent. Unusual and unexpected measurements and observations may be due to valid relationships between variables that are unknown to the scientist. This cannot be determined until the raw data is processed, analysed and interpreted.

The results, after analysis, need to show whether or not a relationship exists between the variables in the research question. To achieve this, variables need to be presented appropriately. Being able to present results appropriately is dependent on appropriate measurement, observation and recordings (e.g. quantitative or qualitative). Make sure this is planned before the experiment is conducted.

The raw data needs to be analysed, then using mathematical and scientific conventions it is represented using tables, graphs, schematics or diagrams. Refer to Module 1.6 Tables and graphing for specific guidance on producing quality and appropriate graphs, and Module 1.11 Communicating and writing a scientific report for details regarding representing results.

IDENTIFYING ERRORS

Most practical investigations have errors associated with them. As shown in Figure 1.10.1 on page e84, there are many different types of errors that can occur.

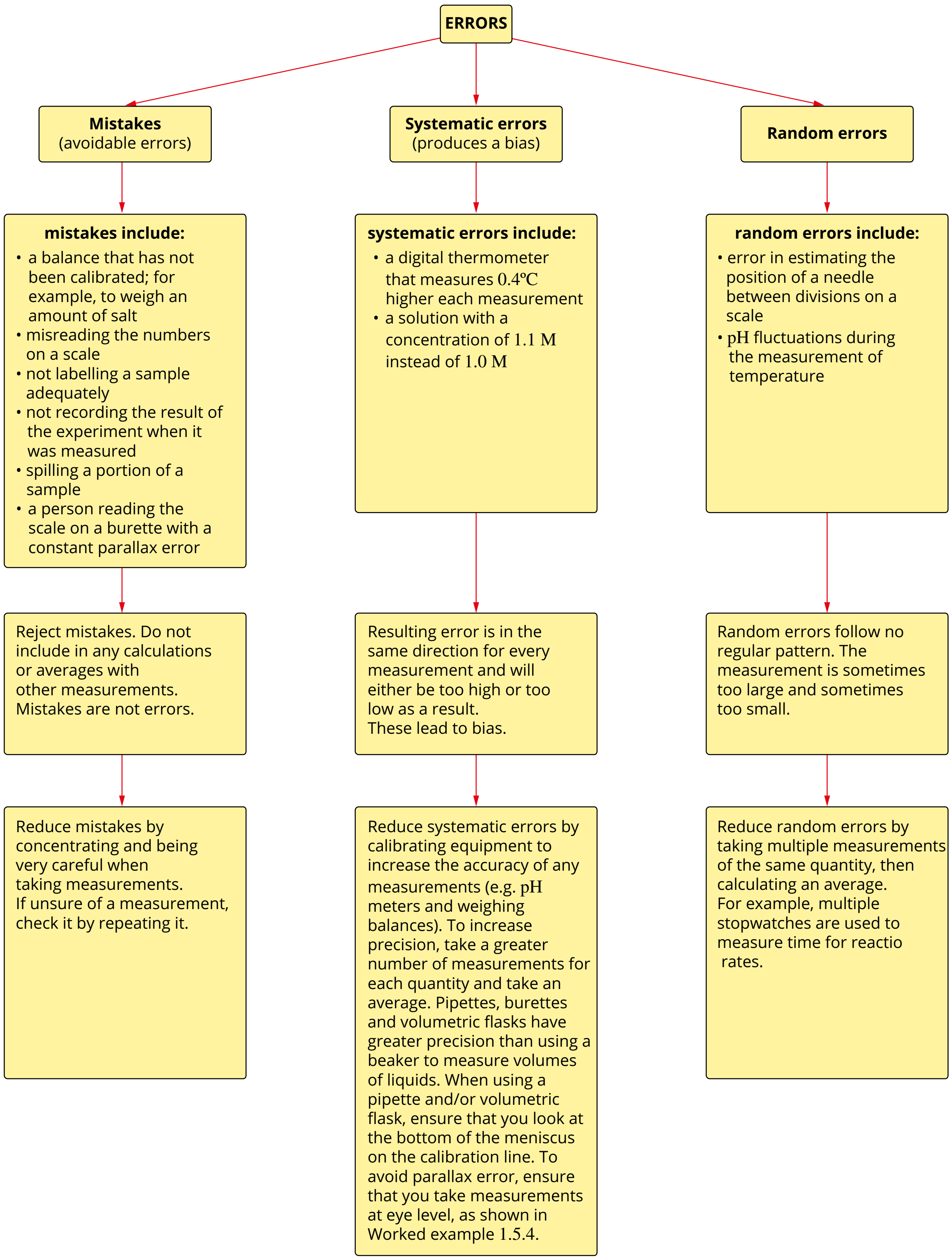
ANALYSING

Analysing the raw data enables processing in numerous ways to search for relationships between variables, trends or patterns in the data, uncertainties and errors, outliers and results of significance. This will produce **processed data**.

There are a number of ways of presenting data, including tables, graphs, flow charts and diagrams. The best way to visualise the data depends on its nature. More information on these different formats is provided in Module 1.10. In this module, you will learn how to discuss your investigation and draw evidence-based conclusions in relation to your research question.

It is important that when analysing data, it is not selectively processed to demonstrate what the scientist wants to see. Such bias will result in using analysis tools (for example, statistics) inappropriately and erroneously and will result in invalid conclusions and academic fraud. Quality scientific analysis processes the raw data as it is and is open to any result.

The QCAA criterion 'Analysis of evidence' stipulates the characteristics of the top performance level. The interpretation of these characteristics is shown in Figure 1.10.2 on page e85. To be awarded a mark of 5 or 6 for this criterion, you need to show a much more thorough, thoughtful and comprehensive engagement with the task. For example, appropriately applying algorithms and representations of data through correct and relevant data processing is working at the highest performance level. However, adequate application and basic data processing is working in the next lower performance level. If the application is rudimentary and data is incorrect and irrelevant, the performance level is lower again.



mistakes include:

- a balance that has not been calibrated; for example, to weigh an amount of salt
- misreading the numbers on a scale
- not labelling a sample adequately
- not recording the result of the experiment when it was measured
- spilling a portion of a sample
- a person reading the scale on a burette with a constant parallax error

Reject mistakes. Do not include in any calculations or averages with other measurements. Mistakes are not errors.

Reduce mistakes by concentrating and being very careful when taking measurements. If unsure of a measurement, check it by repeating it.

systematic errors include:

- a digital thermometer that measures 0.4°C higher each measurement
- a solution with a concentration of 1.1 M instead of 1.0 M

Resulting error is in the same direction for every measurement and will either be too high or too low as a result. These lead to bias.

Reduce systematic errors by calibrating equipment to increase the accuracy of any measurements (e.g. pH meters and weighing balances). To increase precision, take a greater number of measurements for each quantity and take an average. Pipettes, burettes and volumetric flasks have greater precision than using a beaker to measure volumes of liquids. When using a pipette and/or volumetric flask, ensure that you look at the bottom of the meniscus on the calibration line. To avoid parallax error, ensure that you take measurements at eye level, as shown in Worked example 1.5.4.

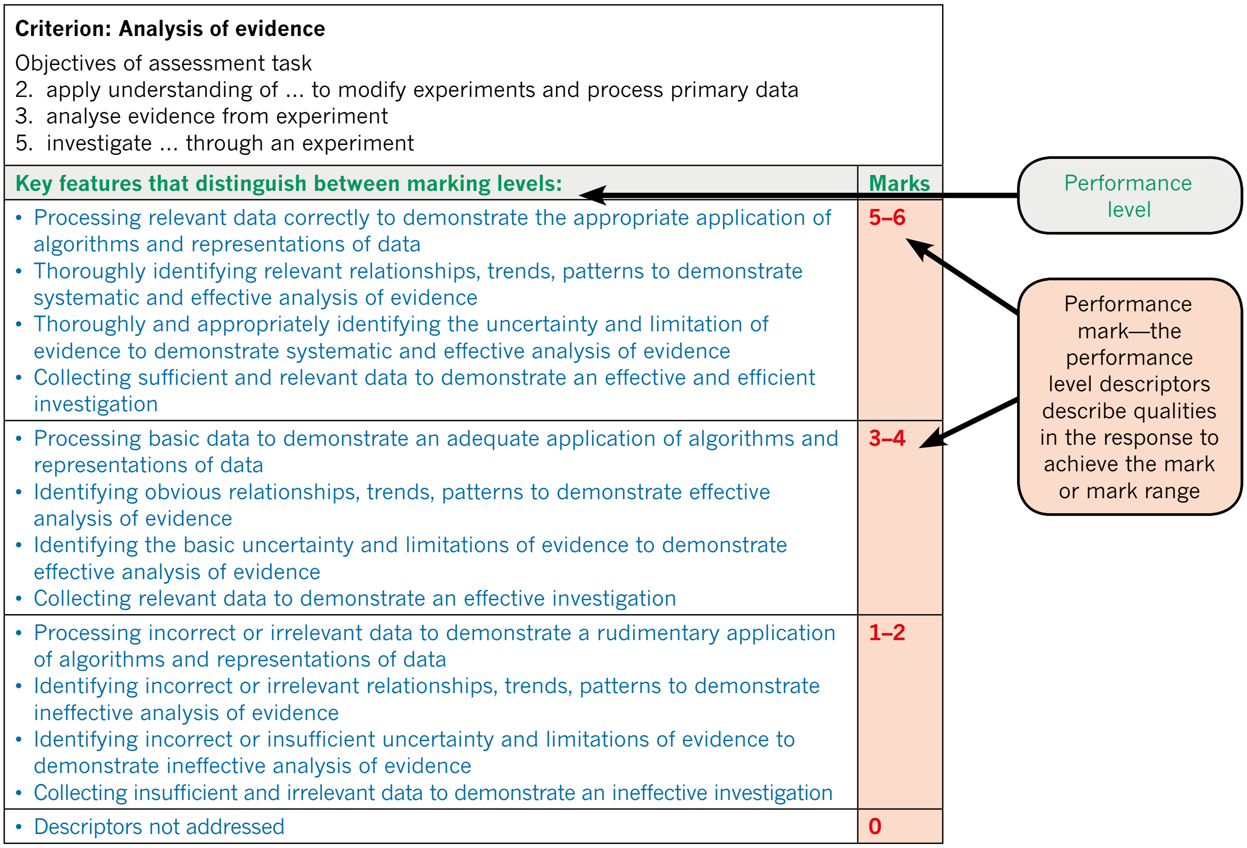
random errors include:

- error in estimating the position of a needle between divisions on a scale
- pH fluctuations during the measurement of temperature

Random errors follow no regular pattern. The measurement is sometimes too large and sometimes too small.

Reduce random errors by taking multiple measurements of the same quantity, then calculating an average. For example, multiple stopwatches are used to measure time for reactio rates.

FIGURE 1.10.1 Types of errors that can be made in an experiment



Criterion: Analysis of evidence

Objectives of assessment task
2. apply understanding of … to modify experiments and process primary data
3. analyse evidence from experiment
5. investigate … through an experiment

Key features that distinguish between marking levels:	Marks
• Processing relevant data correctly to demonstrate the appropriate application of algorithms and representations of data • Thoroughly identifying relevant relationships, trends, patterns to demonstrate systematic and effective analysis of evidence • Thoroughly and appropriately identifying the uncertainty and limitation of evidence to demonstrate systematic and effective analysis of evidence • Collecting sufficient and relevant data to demonstrate an effective and efficient investigation	5–6
• Processing basic data to demonstrate an adequate application of algorithms and representations of data • Identifying obvious relationships, trends, patterns to demonstrate effective analysis of evidence • Identifying the basic uncertainty and limitations of evidence to demonstrate effective analysis of evidence • Collecting relevant data to demonstrate an effective investigation	3–4
• Processing incorrect or irrelevant data to demonstrate a rudimentary application of algorithms and representations of data • Identifying incorrect or irrelevant relationships, trends, patterns to demonstrate ineffective analysis of evidence • Identifying incorrect or insufficient uncertainty and limitations of evidence to demonstrate ineffective analysis of evidence • Collecting insufficient and irrelevant data to demonstrate an ineffective investigation	1–2
• Descriptors not addressed	0

Performance level

Performance mark—the performance level descriptors describe qualities in the response to achieve the mark or mark range

FIGURE 1.10.2 Analysis of evidence characteristics

To be awarded a mark of 5 or 6 for this criterion, you need to show the steps and calculations that you have performed to accurately process the data, as identified in the statement ‘demonstrated by correct and relevant processing of data’.

Analysing precision

Understanding the uncertainty and precision is vital to any analysis of data. In chemistry, there is always variation in measurements, so if the variation is caused by the systematic or random error this needs to be determined. Or, in other words, how much of the variation is due to nature and how much is due to instrumentation or recording? (Precision and uncertainty were explained in Module 1.5.)

The precision and uncertainty of the instruments must always be displayed as a range of data next to the results (e.g. measurement $\pm \Delta$ measurement). If calculations are performed with the results then the appropriate calculations must be done with the uncertainties. When the total uncertainty is known, then it can be established if variation in the data is due to the instrument or the variables being tested.

If the measurements between trials fall within the uncertainty range of the instrument (Table 1.10.1), then the variation in results could simply be due to the instrument. If the differences between the measured results are greater than the uncertainty range, then the variation in the results is not due to the instrument and is due to other variables.

Table 1.10.1 shows the volume of 1.0 M hydrochloric acid required to neutralise 25 mL of an unknown concentration of sodium hydroxide solution, the mean and the absolute uncertainty in the mean. The uncertainty in the mean is calculated using the methods outlined in Module 1.5. There is a notable variation between the measured results. The difference between the volume of hydrochloric acid used between titrations 1 and 2 is just 0.3 mL; however, the difference between titrations 2 and 3 is 0.8 mL. The uncertainty due to the precision in the measurement due to the precision of the instrument is half a division of the scale, so 0.05 mL. Since the difference between the measurements is greater than the uncertainty, these results represent an actual difference in the volumes used.

TABLE 1.10.1 Volume of 1.0 M hydrochloric acid required to neutralise 25 mL of an unknown concentration of sodium hydroxide solution

Trial	Volume of hydrochloric acid used (mL)
titration 1	20.6 ± 0.05
titration 2	20.9 ± 0.05
titration 3	20.1 ± 0.05
titration 4	20.4 ± 0.05
mean	20.5 ± 0.4

If the difference in the values is less than the uncertainty in the measurement of the instrument then it cannot be ruled out that the difference is due to the instrument and not some physical reason. Therefore, the result would be interpreted as being the same.

It is important to understand the accuracy and precision of the instruments as it effects the interpretation of the results. There are a few ways to analyse precision—here are a few examples:

- instrumental uncertainty—displays the precision of the instrument and explains instrumental variation in the measured results.
- range—outlining the difference between the smallest measurement and the greatest.
- central tendency (e.g. mean) with instrumental uncertainty—outlining the potential variation in instrumental measurements due to the instrument's design or increments.

Analysing validity and theoretical relationships

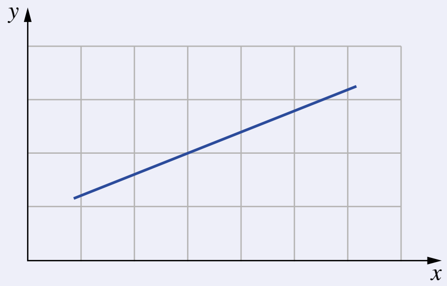
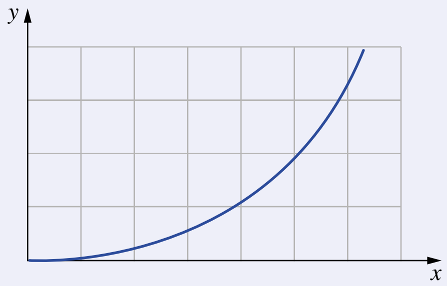
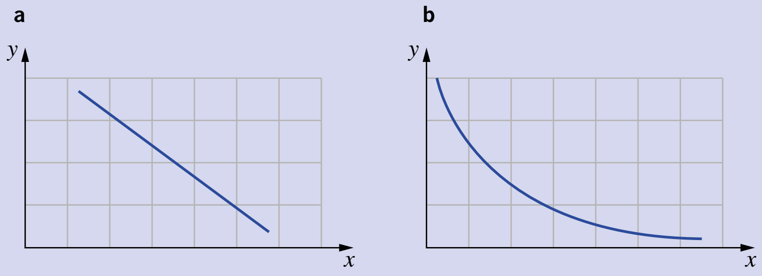
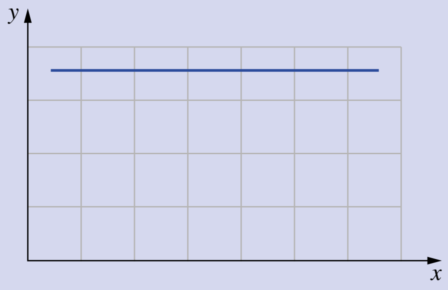
The results and data also need to be processed to look for trends, patterns or differences. As you learnt in the previous section, some of the first processes in analysing data are to use statistical calculations to determine the true values, uncertainties, errors and significance of the measurements. Once the quality of the data is understood then the validity in relation to established theoretical concepts can be analysed.

Analysis can also find **anomalies** and outliers in data, which are not valid measurements. During the experiment, your record of observations may provide a reason for an outlier in your data if one is found. This can then be used to suggest improvements in the methodology to remove such measurements.

Trends in line graphs

Graphs are drawn to show the relationship, or trend, between two variables (Table 1.10.2). Remember that the results may be unexpected. This does not make the investigation a failure. However, the findings must be related to the hypothesis, purpose and method.

TABLE 1.10.2 Graphs showing different trends and relationships between variables

Variables that change in linear or direct proportion to each other produce a straight, sloping trend line.  	Variables that change exponentially in proportion to each other produce a curved trend line. 
When there is an inverse relationship, one variable increases as the other variable decreases. It may be (a) direct or (b) non-linear.  	When there is no relationship between two variables, one variable will not change even if the other changes. 

INTERPRETING

The final part of your investigation involves summarising your findings in an objective, clear and concise manner for your audience. Consider the message you want to convey to the audience when writing your discussion. Statements need to be clear and concise. At the conclusion of your discussion, the audience must understand the context, results and implications of your investigation.

The discussion is the part of your investigation in which you evaluate and explain your methods and results. You should interpret what the results mean. The key sections of the discussion are:

- interpreting and evaluating data
- evaluating the investigative method
- explaining the link between the investigation findings and relevant chemical concepts.

When the results have been processed and analysed, you must apply your scientific understanding and conceptual knowledge to offer explanations of what occurred and why. The explanation of the results is an interpretation. Interpretations are limited by the methodology employed, the instruments used, the variables controlled and the raw data measured. Interpretations should never provide an explanation beyond the constraints of the experimentation and methodology. Interpretations are not meant to provide answers and comprehensive explanations, only a valid and reliable plausibility based only on what was measured. To meet the student experiment word limit, the interpretation of results in the scientific report will need to be concise.

DISCUSSING RELEVANT CHEMICAL CONCEPTS

To make the investigation more meaningful, it should be explained within the right context, meaning the related chemical ideas, concepts, theories and models. Within this context, explain the basis for the hypothesis. For example, if you were studying the impact of dissolved carbon dioxide on the pH of sparkling mineral water you could include the information in Table 1.10.3 in your discussion.

TABLE 1.10.3 Examples of how to include chemical concepts in your discussion

Key ideas	Example
definitions of key terms	'pH', 'dissolved carbon dioxide' and 'sparkling' mineral water
the function of added carbon dioxide	to create 'sparkling' water
relationship between variables	dissolved carbon dioxide and pH of water, temperature was controlled in the experiment
chemical principles	dissolved carbon dioxide and formation of carbonic acid (H_2CO_3), including relevant equations
sources of error	reducing random error, by repeating measurements and calculating average

Relating your findings to a chemical concept

During the analysis stage of your investigation, you identified trends, patterns and mathematical models of your results. This is the framework needed in which to discuss whether the data supported or refuted the hypothesis. Ask questions such as these.

- Was the hypothesis supported?
- Has the hypothesis been fully answered? (If not, give an explanation of why this is so and suggest what could be done to either improve or complement the investigation.)
- Do the results contradict the hypothesis? If so, why? (The explanation must be plausible (reasonable) and must be based on the results and previous evidence.)

Providing a theoretical context also enables comparison of the results with existing relevant research and knowledge. After identifying the major findings of the investigation, ask questions such as these.

- How does the data fit with the literature?
- Does the data contradict the literature?
- Do the findings fill a gap in the literature?
- Do the findings lead to further questions?
- Can the findings be extended to another situation?

Be sure to discuss the broader implications of the findings. Implications are the bigger picture. Outlining them for the audience is an important part of the investigation. Ask questions such as these.

- Do the findings contribute to or impact on the existing literature and knowledge of the topic?
- Are there any practical applications for the findings?

EVALUATING THE DATA

In the discussion, you need to be able to:

- explain your results in relation to chemical ideas, concepts and theories. This will allow you to explain your research question.
- discuss whether your data supports or refutes your research question
- discuss any **secondary data** you have found from journal articles or other reliable sources that is directly connected to the experiment in the interpretation and evaluation of your results and methodology. Explain if the results support your observations. You should report that your results are consistent with the findings. Otherwise, it is essential that you reflect on the contrasting results and identify or suggest possible explanations as to why the results differ. This could be due to differences in the measurement technique or other factors such as the control or measured variables. It might also help you identify possible sources of error within your investigation. At this stage, it is important to check your calculations to make sure that you have not made an error.

Ask yourself questions relating to your research question and the literature justification regarding chemical concepts, as shown in Figure 1.10.3.

DRAWING EVIDENCE-BASED CONCLUSIONS

To match the characteristic in the top performance level in the ISMG (IA2), you need to insightfully interpret the experimental evidence.

This must be demonstrated in your report by including a justified conclusion linked to the research question.

So, the report must analyse your data to link the findings from the experiment back to the research question that has informed your investigation. A conclusion is usually a paragraph that provides a justified and relevant response to your research question. Read the examples of strong and weak conclusions in Table 1.10.4 for the hypothesis and research question shown.

TABLE 1.10.4 Strong and weak conclusions to a research question and hypothesis

	Strong conclusion	Weak conclusion
Research question: **How does the temperature affect the pH of water?**	Analysis of the results on the effect of an increase in temperature of water from 5°C to 40°C showed an inverse relationship in which the pH of water decreased from 7.4 to 6.7. These results support the current knowledge that an increase in water temperature results in a decrease in its pH.	The results show that temperature does affect the pH of water.
Hypothesis: **An increase in the temperature of pond water will result in a decrease in the measured pH of the water sample.**	An increase in temperature from 5°C to 40°C resulted in a decrease in the pH of the water from 7.4 to 6.8.	The pH of water decreased as temperature increased.

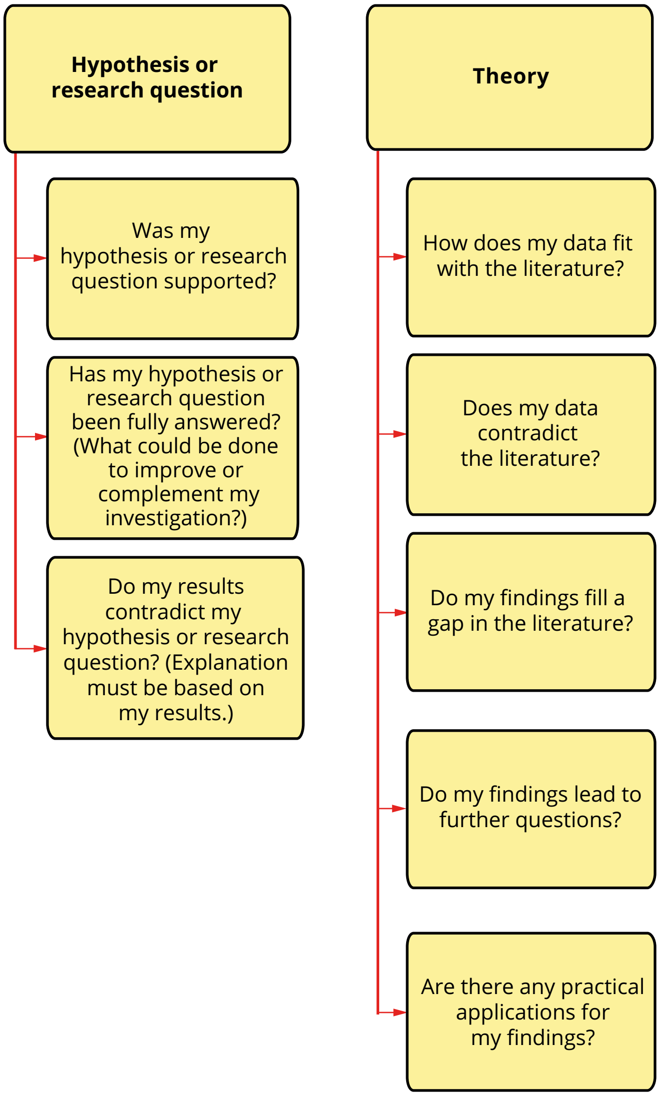


FIGURE 1.10.3 Use these questions to help relate your data to your research question or hypothesis.

Evaluating the method

In addition, the interpretation and evaluation criterion requires:

- critical evaluation of experimental processes
- suggested improvements and extensions to the experiment, which are logically derived from the analysis of the evidence.

These points can help your evaluation.

- Evaluate the method. When relevant, compare the chosen method with any other methods that might have been selected, evaluating the advantages and disadvantages of the selected method and the effect on the results.
- Discuss the results. Indicate the range of the data obtained from replicates.
- Explain how the sample size was selected. Larger samples are usually more reliable, but time and resources might have been scarce. Discuss whether the results of the investigation have been limited by the sample size.
- Discuss reliability in terms of the control group. A control group helps determine if a variable that should have been controlled has been overlooked and may explain any unexpected results.

- Discuss any source of systematic or random errors and suggest ways of improving the investigation. For instance, the more times an experiment is repeated, the more reliable the results are.
- Identify issues that could affect validity, accuracy, precision and reliability of data.
- State sources of systematic and random errors.
- Identify any **bias** in any part of the investigation method, including sampling and measurements. Bias is a form of systematic error resulting from the researcher's personal preferences and can include poor definitions, incorrect assumptions, errors in design and methodology. Some biases cannot be eliminated, but should at least be addressed in the discussion.
- Evaluate the degree of accuracy and precision of the measurements for each variable of the hypothesis. Comment on the uncertainties obtained.
- Recommend improvements to the investigation if it is to be repeated.

Once any limitations or problems in the methodology have been identified, recommend improvements for how the investigation could be conducted if repeated. In the evaluation of the experiment, ensure you discuss the broader implications of your findings.

1.10 Review

SUMMARY

- Raw data includes the measurements and observations made during an experiment.
- Processed data is derived from processing and manipulating raw data.
- Processed data enables trends, patterns and differences to be identified.
- If the measurements between samples or tests fall within the uncertainty range of the instrument, then the variation in results could simply be due to the instrument.
- Interpretations of data attempt to explain the observed results.

KEY QUESTIONS

Retrieval

1 Name two types of data.

2 State what a systematic error is.

3 State what a random error is.

4 State the evidence that is required for the top performance level in the analysis criterion.

Comprehension

5 Explain the difference between processing data and interpreting data.

6 Explain whether the following are mistakes, systematic errors or random errors.

- **a** a student spills some solution during a titration
- **b** the reported measurements are above and below the true value
- **c** a weighing balance that has not been calibrated

Analysis

7 A group of students wanted to test whether the amount of hydrogen gas produced from 0.5 g of zinc was different with 1 M hydrochloric acid or 1 M ethanoic acid.

Their hypothesis was that the ethanoic acid would not produce the same amount of hydrogen gas as it was a weak acid.

Their results are given in the table below.

Trial number	Volume of hydrogen gas from hydrochloric acid (mL)	Volume of hydrogen gas from ethanoic acid (mL)
1	90 ± 2	86 ± 2
2	85 ± 2	86 ± 2
3	87 ± 2	87 ± 2
4	86 ± 2	88 ± 2
5	84 ± 2	85 ± 2
6	89 ± 2	84 ± 2
Mean	87 ± 3	86 ± 2

Determine if the student's hypothesis is correct, based on the student's results.

1.11 Communicating and writing a scientific report

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- identify and explain the sections of a scientific report
- write a scientific report.

To write a scientific report some general conventions need to be followed. Even though there are many ways to present a report, the report must follow a scientific genre and meet the requirements of the Chemistry General Senior Syllabus. This module will provide a guide to writing an appropriate scientific report.

Your journal should contain all the information required to complete the scientific report. After all, it is a report, and reports summarise research and information gathered on a topic. The scientific report does not require much time to complete when reporting from a comprehensive journal.

SCIENTIFIC WRITING AND LITERACY

The words used by scientists (the agreed scientific language and terms) have specific meanings and may differ from the understanding of the term in everyday language (Table 1.11.1). This allows a common understanding of words across all languages, to convey scientific meaning.

TABLE 1.11.1 A comparison of common and scientific language

Term	Common language meaning	Scientific language meaning	Difference in meaning
sample	a representative part or a single item from a larger whole or group especially when presented for inspection or shown as evidence of quality	the specifically chosen physical representatives of a phenomenon that were tested, measured or observed during experimentation or an investigation	Very similar. The scientific term is specific to a controlled or measured phenomenon. The quality is not known until after analysis, though inspection is carried out through testing. In common language it can be used generically, while in scientific language it is specific to the independent variable.
results	to proceed or arise as a consequence, effect or conclusion; to have an issue or result	the recorded evidence of the sample during experimentation or observation	Very similar. The scientific term requires measurement or explicit recording of the observed consequence or effect arising from the independent variable.
significance	something that is conveyed as a meaning often obscurely or indirectly; the quality of conveying or implying; the quality of being important	important; of consequence; expressing a meaning; indicative; includes all that is important; sufficiently great or important to be worthy of attention; noteworthy; having a particular meaning; indicative of something	Similar. The scientific term refers to the quality of being important. To establish importance, statistical analysis has to have been conducted to achieve the meaning, or indicate sufficient difference or worth, and to indicate a particular meaning.
correlation	the state or relation of being correlated; specifically a relation existing between phenomena or things or between mathematical or statistical variables which tend to vary, be associated, or occur together in a way not expected on the basis of chance alone	a relationship existing between phenomena on the foundation of statistical or processed analysis	Almost identical. The scientific term refers to when data processing or statistics has been applied to the common term. As such the strength of the correlation can be predicted or assumed.

The page e93 report must be written using scientific language and conventions. The report should be fluent and concise. Carefully plan all parts of the report: the introduction, the summary of the altered methodology, the results, the analysis (including the interpretation) and conclusion (including evaluation and suggested improvements). Having a plan helps to achieve fluent and concise writing.

Writing fluently and concisely

The report should be brief and comprehensive, so only use the words required to communicate your information. For the report to be concise avoid repetition and remain within the required word count or length. Use scientific language as it allows you to communicate details and knowledge in fewer words. Table 1.11.2 demonstrates communication of the same concept in a few different ways. Being fluent and concise will significantly influence the word count of the report and the quality of the writing.

TABLE 1.11.2 Communicating concepts in various ways

Concept	Communication	Commentary of communication
the basic structure of an atom	An atom is made up of protons, neutrons and electrons.	This is fluent and concise, but lacks scientific terms or language to convey in-depth understanding.
	Atoms are tiny little objects made up of numerous parts that form the basis of matter. Atoms are made up of electrons, neutrons and protons. The electrons are found away from the centre of the atom. The centre of the atom contains the protons and neutrons.	This is fluent, but not concise. There is repetition (e.g. 'tiny' and 'little'—these terms have the same meaning; also 'made up' is stated twice). It also uses many words without using scientific terms to provide detail and understanding. This may be useful language in the journal when learning about a topic, but not in the scientific report.
	Atoms are particles composed of three subatomic particles: the proton, electron and neutron. The centre of an atom is the nucleus and contains the protons and neutrons, with the electrons occupying space external to the nucleus.	This is fluent and concise. Much in-depth understanding is communicated through scientific terms, which enable much information to be conveyed in fewer words.

Writing in a scientific genre

There are a number of genres or styles of scientific communication: literature review, empirical essay and poster presentation etc. The focus here will be on scientific reports. A report is a document that communicates a summary of information, focusing on the main points of interest. It uses headings, sections, tables and graphs to present information. A **scientific report** is written to scientific conventions including format and language.

Reporting

Using headings in scientific reports is essential. There are international conventions for scientific report writing; however, they are specific to the journal publication. There is no single convention for scientific report writing. Table 1.11.3 lists headings that are commonly used sections in scientific reports and describes the information that would be provided under each heading. Sections can be broken down further into subsections. For example, the method can be broken into subsections like sampling techniques, preparation for experiment, equipment used and methodology. As it can be seen, some subheadings are suitable for more than one section; however each scientific report will only use each heading and subheading once. It is best to ask your teacher which headings they prefer and ask how to align your scientific report and its headings to the syllabus.

TABLE 1.11.3 Scientific report sections, headings and information

Scientific report section title	Common title alternatives	Expected information within the section
title	n/a	• a specific statement that outlines the expected relationship between independent and dependent variables or a question asking about the relationship between the independent and dependent variables
abstract	n/a	• summarises the entire experiment/investigation into a single paragraph, outlining the main information for each section of the background information, method, results, analysis and conclusion • summary for each section is usually 1–3 sentences per section, usually less rather than more
introduction	background information background research literature review	• provides the information already known or inferred from previous experimentation and scientific literature specific to the hypothesis • explains the current scientific knowledge about the relationship between the independent and dependent variables and any other variables that may alter the relationship • must also make reference to the original experiment and justify the modification that was made
method - sampling technique - preparation - experiment	procedure	• outlines the exact details for other scientists to repeat the experiment including specific details about instruments, equipment models and precision, techniques employed and all information required for others to repeat the same results • it is not common to differentiate the materials (equipment and instruments used) from the procedure that uses them
results - raw data - analysis - statistical analysis - interpretation	processed data analysis	• the type of data presented is unique to specific journal publications; however, in general, all the data, observations and results need to be presented to explain the interpretation and conclusion • must show all the required information to answer the research question or hypothesis
conclusion - analysis - interpretation - discussion - evaluation	interpretation discussion evaluation	• provides an explanation of the results including quality of experimentation (accuracy, precision, validity and reliability of the methodology) and relates the results to current scientific understanding (theory) • the strength of the relationship between the experimental evidence (data, observations or results) is to be stated • no new data is to be included in this section
references	n/a	• a list of all sources used in the scientific report

Scientific writing

Since the report is written for a scientific audience, it is important to ensure that the report uses appropriate scientific language and conventions. This contrasts with English writing used in everyday situations.

The experiment report is to be written using paragraphs, with each paragraph explaining only one idea. The first sentence (topic sentence) of the paragraph introduces the topic. Following sentences provide the details of the idea. The final sentence concludes the idea. Each sentence should flow onto the next, slowly building the details of the explanation. The report should be brief and comprehensive, so you should only use the words required to communicate and use language that conveys detailed understanding.

Reports should be written using:

- past tense—as the experiment was conducted in the past, so the report should be in the past tense
- third person, passive voice and impersonal verbs—science uses this language convention (Table 1.11.4)
- scientific language—the terms used are specific to concepts, models and theories
- objective, unbiased language—avoid subjective and emotional or persuasive writing (Table 1.11.5)
- concise language—avoid unnecessary repetition and express ideas. Scientific language allows more details, knowledge and understanding to be communicated in fewer words. Use shorter sentences that are less wordy. Table 1.11.6 shows some examples of more concise wording.

TABLE 1.11.4 Examples of first-person and third-person narrative

First person	Third person
I put 50 g of marble chips in a conical flask and then added 10 mL of 2 M hydrochloric acid.	*First, 50 g of marble chips was weighed into the conical flask and then 10 mL of 2 M hydrochloric acid was added.*
After I observed the reaction, I found that …	*After the reaction was completed, the results showed …*
My colleagues and I found …	*Researchers found …*

TABLE 1.11.5 Persuasive writing versus scientific writing styles

Persuasive writing examples	Scientific writing equivalent examples
Biased and subjective language: *The results are extremely bad/atrocious/ wonderful etc.* *This is terrible because …* *This produced a disgusting odour.* *Health crisis!*	Unbiased and objective language: *The results showed …* *The implications of these results suggest …* *The results imply …* *This produced a pungent odour.* *Health issue*
Exaggeration: *The object weighed a colossal amount, like an elephant.*	Non-emotive language: *The object weighed 256 kg.*
Use of everyday or colloquial language: *The bacteria passed away.* *The results don't …* *The researchers had a sneaking suspicion …*	Use of formal language: *The bacteria died.* *The results do not …* *The researchers predicted/hypothesised/ theorised …*

TABLE 1.11.6 Examples of verbose and concise writing

Verbose writing	Concise writing
Due to the fact that …	*Because …*
Smith undertook an investigation into …	*Smith investigated …*
It is possible that the cause could be …	*The cause may be …*
end result …	*result …*
In the event that …	*If …*
shorter in length …	*shorter …*

To produce a fluent report, the scientific language must be used without error so that the reader understands the meaning of the information easily. For the report to be concise, there should not be any repetition. Being fluent and concise will significantly influence the word count and the quality of writing.

Language constructs

Developing paragraphs is essential to writing fluently and concisely. Each paragraph should explain only one topic. The first sentence of the paragraph introduces the topic, the following sentences provide the details of the topic, and the final sentence concludes it. Each sentence within a paragraph should contain only one subject (perhaps two if necessary), and each sentence should flow onto the next, slowly building the details of the explanation.

Presenting scientific ideas

An efficient way of presenting complex data and explaining scientific concepts is through photographs, graphs, tables and scientific models such as flow charts and diagrams. Ensure you include:

- a descriptive title
- labels, captions or descriptions
- numbering; for example, Figure 1, Figure 2 etc. or Table 1, Table 2 etc.
- a source if the work is not your own or is adapted from work that is not your own.

Using tables and graphs

In general, tables provide more detailed data than graphs. However, it is easier to observe trends and patterns in graphs, making them a very useful tool for presenting evidence. Pie charts illustrate percentages well, while bar charts and line graphs illustrate relationships between variables well. More information on tables and graphs can be found in Module 1.6.

Editing your report

Editing your report is an important part of the process. After editing your report, save new drafts with a different file name and always back up your files in another location. Pretend you are reading your report for the first time when editing. Once you have completed a draft, it is always good practice to read your work, hours or a day or two after you have completed it. When reading your own work, do not read it as you intended. Instead, carefully read your work, following the punctuation, grammar and spelling as it appears on the page. This is more easily achieved if you read the report aloud.

When editing, look for content that is:

- ambiguous or unclear
- repetitive
- awkwardly phrased
- too lengthy
- not relevant to your research question
- poorly structured
- lacking evidence
- lacking a reference (if it is another researcher's work)
- spelt incorrectly.

Acknowledging sources

All the quotations, documents, publications and ideas used in your investigation need to be listed in the references and acknowledgments. In order to avoid plagiarism and to ensure creators are properly credited for their work, this must be completed accurately. References and acknowledgements also give credibility to your study and allow the audience to locate information sources should they wish to study it further.

Plagiarism is using other people's work without acknowledging them as the author or creator. To avoid plagiarism, include a reference every time you report the work of others; for example, at the end of a sentence or following a diagram. If you use a direct quotation from a source, enclose it in quotation marks. This will ensure you give credit to the original author and it will enable the reader to find the original source.

Referencing

There are a number of different referencing styles, such as Harvard or APA. The sources of information must be acknowledged using a referencing style that is suitable for the purpose of the investigation. Check the preferred referencing style or convention of your teacher for in-text referencing and a complete **reference list**. APA (American Psychological Association) style is the most commonly used referencing style.

The student experiment does not require a bibliography; a reference list is sufficient. A bibliography is a list of all the sources used during the research to develop understanding (including information in the journal) even if the information was not used directly or explicitly in the scientific report. A reference list only lists the sources used and cited (or in-text referenced) in the scientific report.

The sources must be listed at the end of the report in alphabetical order (by author last name or organisation name).

Compile your references in a separate document as you conduct the student experiment. This will save you time later.

In-text citations

Each time you write about the findings of other people or organisations, you need to provide an in-text citation and provide full details of the source in a reference list. In the APA style, in-text citations include the first author's last name and date in brackets (author, date).

The following example shows two types of in-text citation and the reference list entry.

In-text citations
It was reported that in testing of five pro-oxidant additives added to commonly manufactured polymers, none resulted in significant biodegradation after three years (Selke et al., 2015).
Selke et al. (2015) reported that in testing of five pro-oxidant additives added to commonly manufactured polymers, none resulted in significant biodegradation after three years.
Reference list entry for the example above would be:
Selke, S., Auras, R., Nguyen, T.A., Aguirre, E.C., Cheruvathur, R., & Liu, Y. (2015), Evaluation of biodegradation—promoting additives for plastics. *Environmental Science & Technology*, 49(6), 3769–3777.

ADDRESSING THE SYLLABUS INSTRUMENT-SPECIFIC MARKING GUIDE

It is imperative that your final scientific report addresses all the descriptors of the Student experiment instrument-specific marking guide (ISMG). Before you begin the scientific report it would be best to plan the sections and titles and then assign the ISMG descriptors to each section. It would be best to discuss this with your teacher and develop a plan together.

Word limits or word count guides can be assigned to each section. Apply a larger word count to the sections requiring explanations such as introduction, discussion/conclusion and maybe the analysis (depending on the convention chosen). This word allocation will act as a guide, but is flexible. Make sure your response does not exceed the word limit prescribed by the assessment specifications.

When you edit your completed report assess whether it includes all the evidence in the four criteria (Research and planning, Analysis of evidence, Interpretation and evaluation, Communication) that make up the ISMG (IA2). These are the words that follow 'demonstrated by …'

1.11 Review

SUMMARY

- A scientific report has these features.
  - Title: a specific statement that outlines the expected relationship between independent and dependent variables
  - Abstract: a single-paragraph summary of the entire experiment/investigation
  - Introduction: an explanation of the current scientific knowledge about the relationship between the independent and dependent variables and any other variables that may alter the relationship
  - Method: an outline of the exact details of the experiment, including specific details about instruments, equipment models and precision, techniques employed and all information required for others to repeat the same results
  - Results: presentation of the data, observations and results in the most appropriate way to explain the interpretation and conclusion. The results must show all the required information to answer the research question or hypothesis.
  - Conclusion: an explanation of the results, including quality of experimentation. The explanation is related to current scientific understanding. The strength of the relationship between the experimental evidence is stated. Do not put any new data into this section
  - References: a list of all sources used in the scientific report.
- The scientific report must address the requirements of the syllabus.

KEY QUESTIONS

Retrieval

1 Describe what information is included in the following sections of a scientific report.

  a method

  b conclusion

2 Recall in which section of a scientific report you would find processed data.

Comprehension

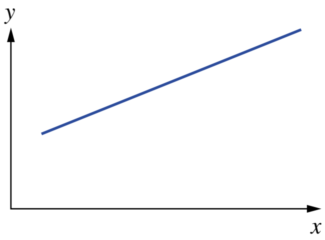
3 Explain the difference between a bibliography and a reference list.

Analysis

4 Consider the graphs below to answer the following questions.

Direct or linear proportional relationship

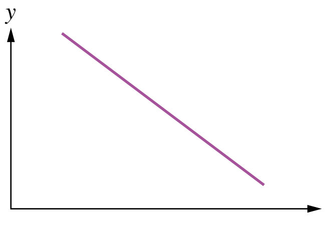
a



- Variables change at the same rate (graph line is straight, slope is constant).
- Positive relationship—as x increases, y increases.

Inverse direct or linear proportional relationship

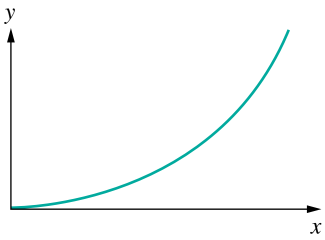
b



- Variables change at the same rate (graph line is straight, slope is constant).
- Negative relationship—as x increases, y decreases.

Exponential relationship

c



- As x increases, y increases slowly, then more rapidly.

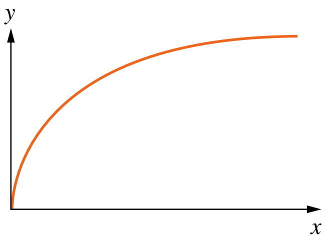
Inverse exponential relationship

d

- As x increases, y decreases rapidly, then more slowly, until a minimum y value is reached.

Exponential rise, then levels off or plateaus (stops rising)

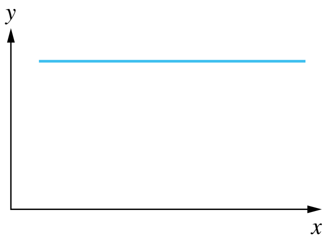
e



- As x increases, y increases rapidly at first, then slows, then finally does not increase at all—y reaches a maximum value.

No relationship between *x* and *y*

f



- As x increases, y remains the same.

a Identify the graph/s that shows the dependent variable increasing at a constant rate.

b Identify the graph/s that shows the dependent variable decreasing at a decreasing rate.

c Identify the graph/s that describes the following observation.

You are measuring the volume of air in a vertically held syringe. A number of books are placed on top of the syringe and you observe the volume of the air decrease at a decreasing rate.

d Describe the relationship between the dependent variable and the independent variable in graph **f**.

5 Determine what might cause a sample size to be limited in an investigation.

PART C RESEARCH INVESTIGATION

QCAA makes the following statement relating to the research investigation in Unit 1 Chemistry: ‘It is suggested that student performance on Unit 1 is assessed using a research investigation … modelled on the techniques used in Units 3 and 4’.

In constructing assessment instruments for Unit 1, schools should ensure that the objectives cover, or are chosen from, the unit objectives. If one assessment instrument is developed for a unit, it must assess all the unit objectives; if more than one assessment instrument is developed, the unit objectives must be covered across those instruments.

It is suggested that student performance on Unit 1 is assessed using a research investigation and a data test modelled on the techniques used in Units 3 and 4, and an examination covering Unit 1 and 2 subject matter to be held at the conclusion of Units 1 and 2.

The QCAA requires students to complete a research investigation in Unit 4 Chemistry. In preparation for Unit 3, teachers may choose to assign a similar assessment task in Units 1 or 2 as preparation for Unit 3.

The research investigation assessment task requires students to investigate a claim, by drawing on secondary evidence from scientific texts. Students use research conventions to analyse and interpret the evidence to reach a justifiable conclusion about the claim. The research requires students to locate and use information beyond the scope of their knowledge and data they have been given.

The research investigation requires you to gather secondary evidence on a research question. The QCAA Syllabus states that students must work individually to develop and investigate their own research question based on a number of possible claims the teacher provides.

Evidence must be obtained by researching scientifically credible sources, such as scientific journals, books, government websites, universities, independent research bodies or science and technology manufacturers.

The research investigation constitutes 20% of the total assessment in Unit 4 Chemistry.

A summary of the objectives, specifications and marking for the research investigation (Unit 4) is provided below.

The research investigation may be presented in:

- written form (e.g. scientific report), 1500–2000 words, or
- multimodal presentation form (e.g. poster presentation), 9–11 minutes.

Criteria	Assessment objectives	Specifications	Marks
Research and planning	• Apply understanding • Perform an investigation	• a considered rationale showing how the research question was developed from the claim • a research question that is specific and relevant • collected sources that are sufficient and relevant	6
Analysis and interpretation	• Analyse the evidence sourced during the research • Interpret the research evidence	• collection of sufficient and relevant sources • detailed and careful coverage of relevant trends, patterns and relationships • detailed and careful coverage of the limitations of the evidence • justified scientific arguments based on evidence	6
Conclusion and evaluation	• Interpret the evidence from the research • Evaluate the processes, claims and conclusions within the research	• a conclusion that is justified and addresses the research question • insightful examination of the quality of the evidence • extension of investigation findings that are credible • consideration of possible improvements and extensions to the investigation that are relevant to the claim	6
Communication	• Present the research findings, including arguments and conclusions	• scientific language and representations that are concise and fluent • suitable use of genre conventions • appropriate referencing conventions to acknowledge sources	2
Total			20

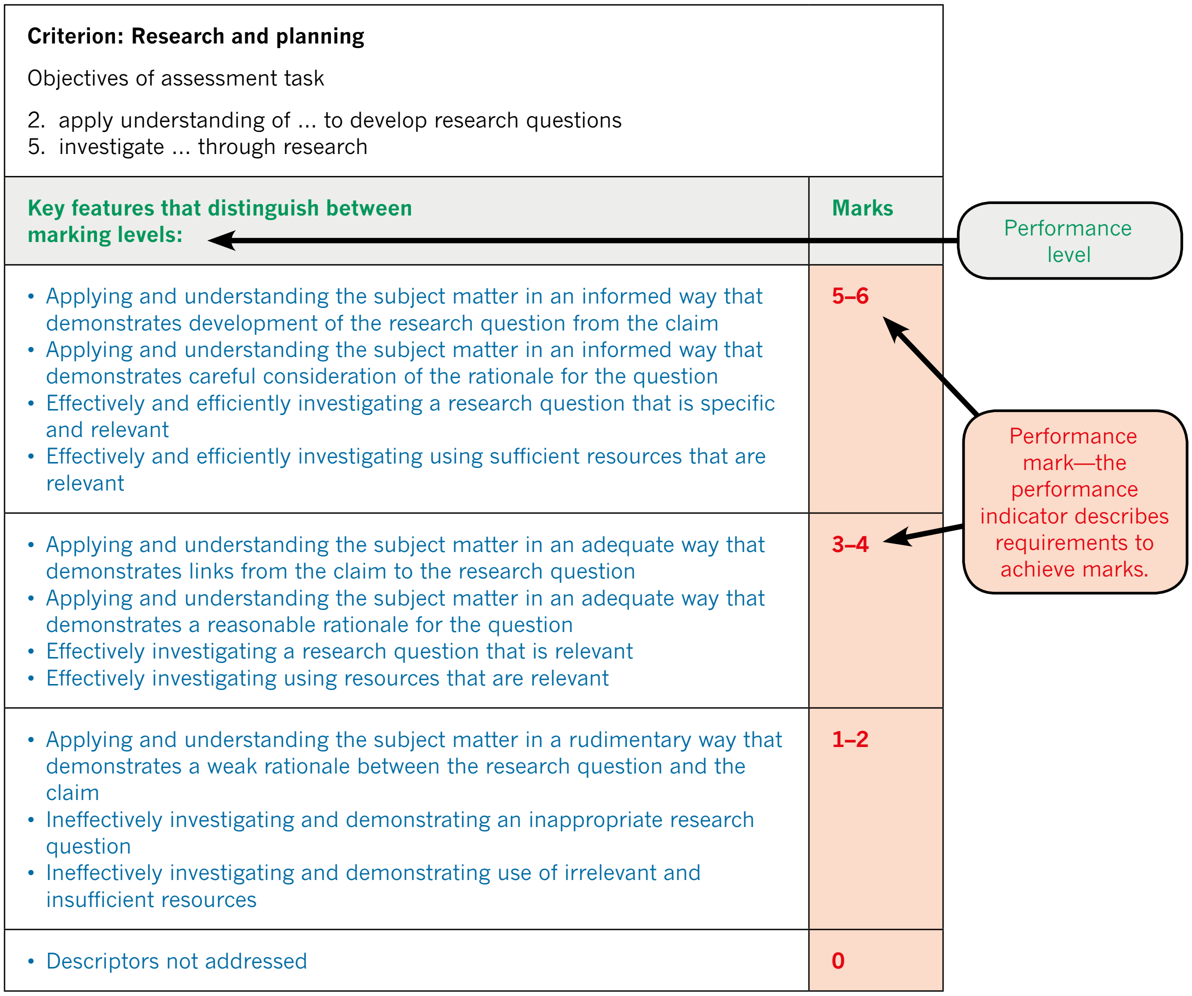
The scientific inquiry is not a linear process. Scientists will not necessarily complete these steps in the stated order and some steps may need to be repeated or altered to more accurately address the research question.

INSTRUMENT-SPECIFIC MARKING GUIDE

Student responses are assessed against an instrument-specific marking guide (ISMG). In developing your research investigation and planning your response it is important to always have in mind the assessment objectives, and in particular the characteristics that are described in the performance level descriptors.

The major features of ISMG are outlined below and shown for the Research and planning criterion. Just as with the Student experiment, the ISMG is organised into:

- four criteria, though these differ for the research investigation—research and planning, analysis and interpretation, conclusion and evaluation, and communication
- performance levels, against which the qualities of the response are assessed
- performance level mark, which may be a single mark or two-mark range
- performance level descriptor.



Criterion: Research and planning

Objectives of assessment task

2. apply understanding of … to develop research questions
5. investigate … through research

Key features that distinguish between marking levels:	**Marks**
• Applying and understanding the subject matter in an informed way that demonstrates development of the research question from the claim • Applying and understanding the subject matter in an informed way that demonstrates careful consideration of the rationale for the question • Effectively and efficiently investigating a research question that is specific and relevant • Effectively and efficiently investigating using sufficient resources that are relevant	**5–6**
• Applying and understanding the subject matter in an adequate way that demonstrates links from the claim to the research question • Applying and understanding the subject matter in an adequate way that demonstrates a reasonable rationale for the question • Effectively investigating a research question that is relevant • Effectively investigating using resources that are relevant	**3–4**
• Applying and understanding the subject matter in a rudimentary way that demonstrates a weak rationale between the research question and the claim • Ineffectively investigating and demonstrating an inappropriate research question • Ineffectively investigating and demonstrating use of irrelevant and insufficient resources	**1–2**
• Descriptors not addressed	**0**

Performance level

Performance mark—the performance indicator describes requirements to achieve marks.

The following modules outline the steps that assist the development of the research investigation report.

1.12 Developing the research question from a claim

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- analyse a claim to identify scientific concepts, variables and measurable terms within a claim
- develop a research question or hypothesis from a claim.

The research question should specifically address one of the concepts associated with the claim. The teacher will provide a number of possible claims. Students develop their own research question based on one of these claims. The research question should clearly state the relevant variables. Students collect secondary evidence to evaluate the claim. All the research and the research investigation will be directly related to the research question. Therefore, the process begins with the claim and understanding the concepts it addresses.

UNDERSTANDING THE CLAIM

The Queensland Chemistry Senior Syllabus defines a **claim** as 'an assertion made without any accompanying evidence to support it'.

The assertion or claim can be a sentence, a statement within a sentence, the title of an article, a quote or anything published in any form. The assertion has to be in isolation from any justification using data, research, evidence or reason from known information to support it.

So how is a research question formed from a claim that is not supported by evidence? The claim itself has to be analysed and understood. Within the claim, identify one or more of the following:

- known scientific concepts
- variables
- measurable terms
- ideas related to concepts
- a term that is stated to influence another.

You should record all the information you collect during the investigation in a journal, including the process of developing a research question from a claim. This will be used in your research investigation report to address the 'development of the research question from the claim' requirement of the ISMG. Write down all the elements found in the claim and try to categorise them using the above list. Each element within the claim may suit more than one category. Table 1.12.1 on page e102 outlines an example of analysing and categorising elements of a claim.

TABLE 1.12.1 Three examples of analysing and classifying elements of a claim

Example 1		
claim	Plastics are bad for the environment.	**Classifying elements of the claim**
elements	• Plastics are made from crude oil. • Plastics are disposed of and not recycled. • Plastics release toxic chemicals. • Plastics are not biodegradable.	Plastics are made from crude oil. • hydrocarbons • polymers • non-renewable resource • manufacture requires lots of resources/energy Plastics are disposed of and not recycled. • all plastics can be recycled but not all are • waste Plastics release toxic chemicals. • breakdown in environment • chemicals released from combustion • additives Plastics are not biodegradable. • polymers are inert • not easily broken down in nature by air, water, sunlight
Example 2		
claim	Natural shampoos are better than commercial shampoos.	**Classifying elements of the claim**
elements	• Define the term 'better'. • Natural shampoos do not contain harmful chemicals. • Natural shampoos are as effective cleaning products as commercial shampoos.	Define the term 'better'. • more environmentally friendly • better cleaning efficiency Natural shampoos do not contain harmful chemicals. • synthetic chemicals are toxic Natural shampoos are as effective cleaning products as commercial shampoos. • how surfactants work • cleaning effectiveness is measurable and related to surface tension, wetting time, solid content, foaming ability
Example 3		
claim	Green chemistry is clean chemistry.	**Classifying elements of the claim**
elements	• green chemistry • environment • clean chemistry	Green chemistry • green solvents • preventing pollution • renewable energy • atom economy—more effective synthetic reaction • renewable materials Environment • fewer pollutants released • fewer resources required Clean chemistry • non-toxic • less energy required • reduced reagents

Separating the claim into elements, such as related terms and concepts, enables you to formulate variables, measurable items and questions.

FORMING A QUESTION

A research question needs to be formed from the claim, as the variables stated in the claim may not be measurable or directly observable, and therefore they may not be scientific.

After **extrapolating** and expanding the claim further into possible scientific elements, each of these elements can be used to form a question (Table 1.12.2 on page e103). It is best to formulate a number of questions related to the claim and write them all down in the journal. Each question must specify a dependent variable and enable a response that will evaluate the validity or the reliability of the claim.

This necessitates forming questions and using the elements of the claim in the phrasing of the research question. To complete this task you will need to have an understanding of the dependent and independent variables related to the claim and the relationship between the variables. The claim must identify the dependent variable. The claim may also specify the independent variable.

Some guidelines to help form a research question include:

- finding the dependent variable in the claim, or refining it by rephrasing it into something measurable.
- choosing an element of the claim to become the independent variable, or identifying it in the material.
- phrasing a question to ask if the independent variable will influence, cause or correlate with the dependent variable.
- writing a few different questions. Usually the questions improve as you write more, which allows the formulation of a more developed research question.

Read through Table 1.12.2 to see examples of questions formed from the elements of the claims that are refined.

TABLE 1.12.2 Three examples of questions formulated from a claim

Example 1	
claim	Plastics are bad for the environment.
elements of a claim	Plastics are made from crude oil. About 4% of the crude oil is used to make plastic. Hydrocarbons are collected by distillation. Non-renewable nature of the resource.
formulated questions	Can the efficiency of the distillation process be increased? Can reactants in the plastics process be made from something other than crude oil?

Example 2	
claim	Natural shampoos are better than commercial shampoos.
elements of a claim	Natural shampoos are as effective cleaning products as commercial shampoos. Shampoo effectiveness can be evaluated by comparing the surfactant properties of the shampoos. Surfactants have a non-polar aliphatic chain and polar head. Properties: surface tension wetting ability solid content detergency foaming ability
formulated questions	Does a greater reduction in the surface tension mean that the shampoo is more effective? Do natural shampoos work in the same way as synthetic shampoos?

Example 3	
claim	Green chemistry is clean chemistry.
elements of a claim	green chemistry atom economy
formulated questions	Can the chemical yield of a reaction be increased? Does a higher atom economy mean a greener product? Is it possible to make the same product in different ways?

Refining the research question

It is possible that one of the questions you write in your journal, from the elements in the claim, will become your research question. As you conduct research into the concepts of the independent and dependent variables, new information will refine the question.

The process to developing a research question is often cyclical (Figure 1.12.1).

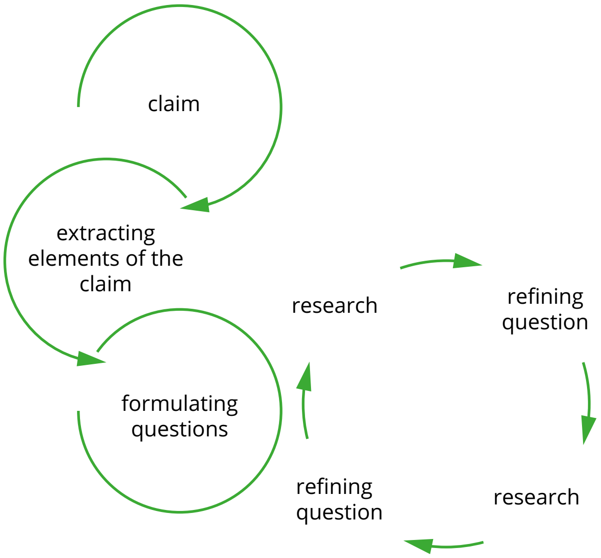


FIGURE 1.12.1 A chart of a common process for developing a research question. This cycle can be repeated as many times as necessary until a scientist is satisfied with the investigation question.

It is important to record your development of conceptual understanding and the knowledge you gain about the relationships between variables. The research investigation requires evidence of the development from the claim to the research question, as stated in the ISMG.

The research question and the development of the research question from the claim using research, scientific concepts, knowledge and understanding are related to the ISMG dot points after the term 'demonstrated by'.

During the research, continue to record your findings in the journal and make note of any ideas that may arise that are related to the questions. As your knowledge and understanding about the variables is developed, refine the question to be more specific.

The goal is to develop the research question to a point at which exact data or evidence can be found regarding the variables in the question. Hence it will develop into a research question when evidence from research can answer the question. Table 1.12.3 compares a formulated question with a developed research question.

TABLE 1.12.3 Development of the original questions formulated into research questions

Example 1		
claim	Plastics are bad for the environment.	
formulated questions	Can the efficiency of the distillation process be increased?	Can reactants in the plastic process be made from something other than crude oil?
refined research question	If the number of stages within the fractional distillation tower is increased to collect fractions with a smaller temperature range, could polymers be made more environmentally friendly?	Can the monomers that are used to generate the plastics used in polyurethanes be generated from a renewable resource?

Example 2		
claim	Natural shampoos are better than commercial shampoos.	
formulated questions	Does a greater reduction in the surface tension mean that the shampoo is more effective?	Do natural shampoos work in the same way as synthetic shampoos?
refined research question	Are natural shampoos better cleaners than synthetic shampoos based on the surface tension, wetting ability, solid content, detergency and foaming ability?	Do the chemicals in natural shampoos create micelles that have a non-polar aliphatic chain with a polar head that increase the solubility of the oily substances in water?

Example 3			
claim	Green chemistry is clean chemistry.		
formulated questions	Can the chemical yield of a reaction be increased?	Does a higher atom economy mean a greener product?	Is it possible to make the same product in different ways?
refined research question	Would increasing the temperature in the Haber process to shift the equilibrium towards the products increase the overall greenness of the ammonia production process?	Can improving the chemical yield of the synthetic reactions through higher atom economy create a greener manufacturing process?	Can the polymers be made from non-petroleum sources and so increase the environmentally friendly nature of plastics?

In your research investigation report, your recorded research in the journal will be used to write the considered rationale for the research question, displaying its clear development from the claim. This will be achieved by using the research in your journal to outline its step-by-step development, justifying the steps using scientific concepts, knowledge and understanding, as shown in Figure 1.12.1.

The QCAA Chemistry General Senior Syllabus defines 'specific' as *clearly defined or identified; precise and clear in making statements or issuing instructions; explicit* and 'relevant' as *bearing upon or connected with the matter in hand; to purpose; applicable and pertinent; having a direct bearing on.*

Therefore, a specific research question must explicitly identify the dependent and independent variables. The research question must be connected to the considered rationale and the topic of study.

1.12 Review

SUMMARY

- A claim is an assertion made without any accompanying evidence to support it.
- Research questions and hypotheses can be developed from claims by identifying the underlying scientific concepts and variables of a claim.
- The number and nature of the sources selected will depend on the specific nature of the claim but need to be sufficient to be able to address the claim within the assessment conditions.

KEY QUESTIONS

Retrieval

1 State what a research question should explicitly identify.

Comprehension

2 Explain why the research question is refined.

3 Explain the relationship between research and understanding, and the research question that is relevant to the ISMG.

Analysis

4 Consider the following claims and develop a research question for each of them.

- **a** Biofuels are a realistic alternative to traditional petroleum-derived fuel. The chemicals in cosmetics are killing you.
- **b** Nanomaterials are the way of the future.

1.13 Finding and choosing suitable resources

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- distinguish between primary and secondary sources
- locate a range of primary and secondary sources
- determine the validity and reliability of a source.

When gathering scientific evidence for the research investigation, source it from reputable publications including:

- scientific journals: research papers and scientific reviews
- scientific articles written by organisations who apply scientific research to their industry
- commercial articles such as science magazines, newspapers and also websites.

The research investigation must include a reference list of sufficient and relevant sources that were used for the investigation. The QCAA Chemistry General Senior Syllabus defines *sufficient* as ‘enough or adequate for the purpose’, and *relevant* as ‘bearing upon or connected with the matter in hand; to the purpose; pertinent; applicable and pertinent; has direct bearing on’. Figure 1.13.1 points out the features of sources deemed *sufficient* and *relevant*.

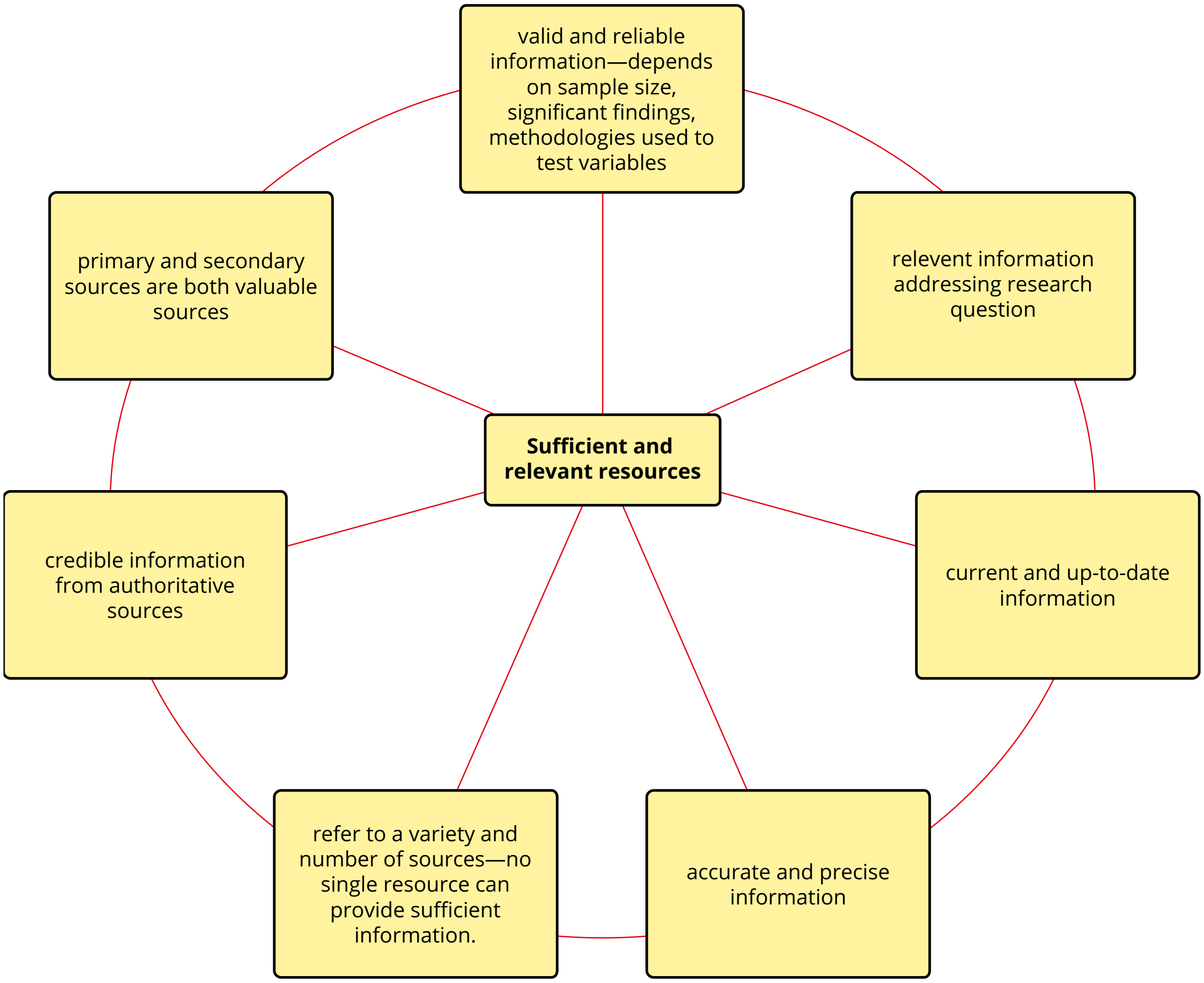


FIGURE 1.13.1 Features of sources suitable for the research investigation

SOURCING INFORMATION

Consider whether the information you use is from a primary or secondary source.

Primary and secondary sources

Table 1.13.1 summarises the characteristics of primary and secondary sources. Sometimes the same information may be contained within a primary and a secondary source, depending on when and by whom it was written. For example, a scientist's journal article on a clinical trial of treatments for teenage obesity is a **primary source**, while a general magazine article about teenage obesity, written by a journalist and referring to the scientific study, is a **secondary source**.

TABLE 1.13.1 Summary of primary and secondary sources

	Primary sources	Secondary sources
characteristics	• first-hand records of events or experiences • written at the time the event happened • original documents	• interpretations of primary sources • written by people who did not see or experience the event • use information from original documents but rework it
examples	• results of experiments • scientific journal/magazine articles • reports of scientific discoveries • photographs, specimens, maps and artefacts • interviews with experts • websites (if they meet the criteria above)	• textbooks • biographies • newspaper articles • magazine articles • radio and television documentaries • websites that interpret the scientific work of others • podcasts

A primary source is written by the observer/witness of an event or the scientist who conducted the research. The information has only been processed by the original observer, therefore it is the least biased of all available sources of information. However, even primary sources may be biased, as the observer or researcher had to make choices related to the observation, control of variables, use of instruments and choices of processing.

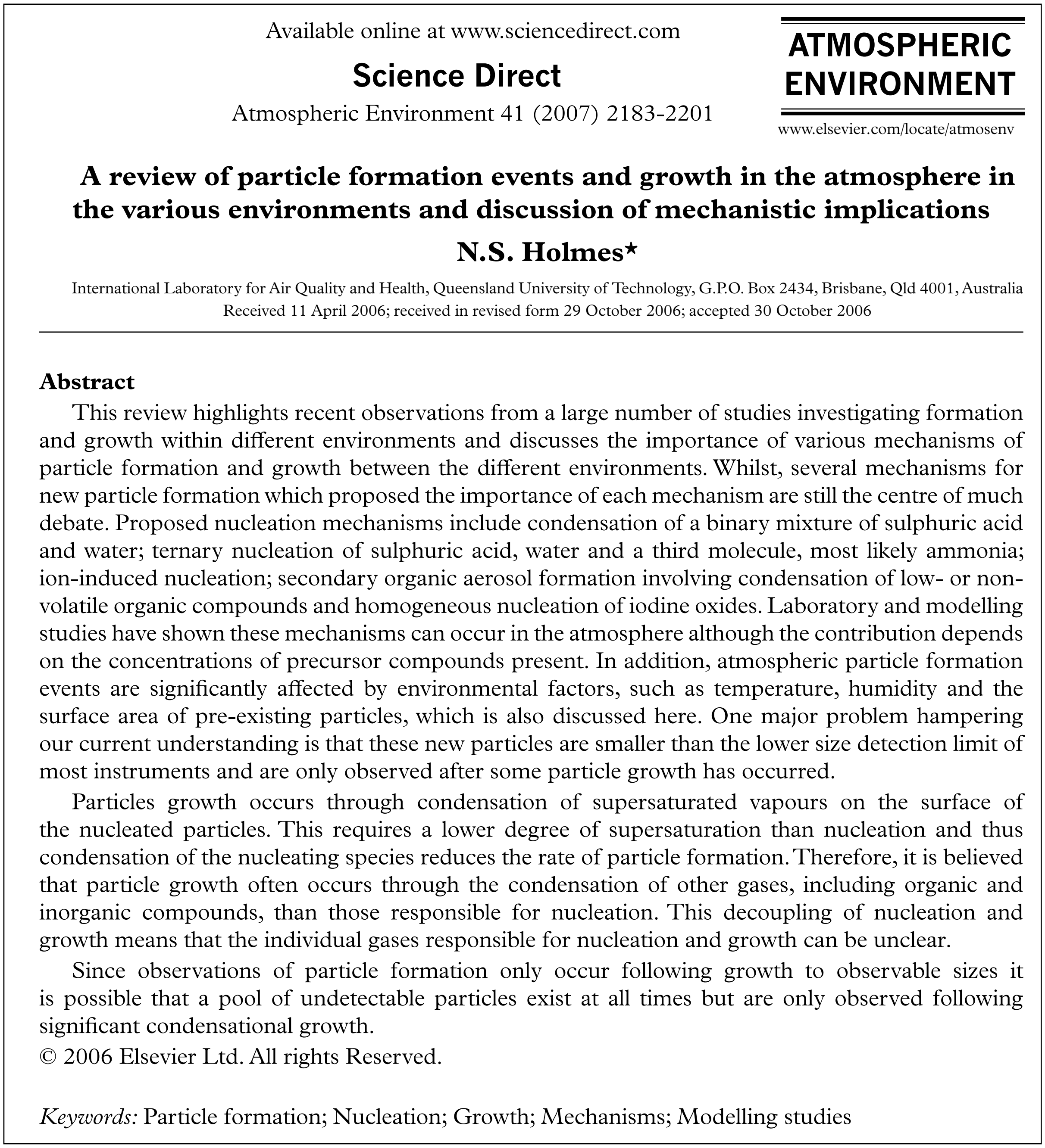
Secondary sources of information are not eyewitness accounts but interpretations of events by other people. As second-hand information, accuracy and reliability may be reduced, and events may be interpreted through the writer's perception and bias. You should aim to use a wide range of data sources when using secondary data, to cross-check for accuracy, reliability and validity of information.

When searching for information and evidence, follow these guidelines.

1. Determine if it is a primary or secondary resource.
2. Confirm that it is valid.
    - **a** Check that it contains information that is specifically related to the claim.
    - **b** Check that the evidence and information are pertinent to the variables in the research question.
3. Assess its reliability.
    - **a** Is it current/recent information?
    - **b** Is it up-to-date in its understanding of relationships?
    - **c** Is the evidence equivalent to other sources?
    - **d** Check credibility—consider who the author is, their qualifications and expertise.
    - **e** Evaluate the methodology, including what variables were controlled or measured.

Articles in scientific journals

Peer-reviewed scientific journals are excellent sources of information. Journals are a collection of scientific reports written by scientists who conducted research. The reports and articles found in scientific journals are published primary sources, meaning they are the results of the experiments (Figure 1.13.2).



Available online at www.sciencedirect.com
Science Direct
Atmospheric Environment 41 (2007) 2183-2201

ATMOSPHERIC
ENVIRONMENT
www.elsevier.com/locate/atmosenv

A review of particle formation events and growth in the atmosphere in the various environments and discussion of mechanistic implications

N.S. Holmes*

International Laboratory for Air Quality and Health, Queensland University of Technology, G.P.O. Box 2434, Brisbane, Qld 4001, Australia

Received 11 April 2006; received in revised form 29 October 2006; accepted 30 October 2006

Abstract

This review highlights recent observations from a large number of studies investigating formation and growth within different environments and discusses the importance of various mechanisms of particle formation and growth between the different environments. Whilst, several mechanisms for new particle formation which proposed the importance of each mechanism are still the centre of much debate. Proposed nucleation mechanisms include condensation of a binary mixture of sulphuric acid and water; ternary nucleation of sulphuric acid, water and a third molecule, most likely ammonia; ion-induced nucleation; secondary organic aerosol formation involving condensation of low- or non-volatile organic compounds and homogeneous nucleation of iodine oxides. Laboratory and modelling studies have shown these mechanisms can occur in the atmosphere although the contribution depends on the concentrations of precursor compounds present. In addition, atmospheric particle formation events are significantly affected by environmental factors, such as temperature, humidity and the surface area of pre-existing particles, which is also discussed here. One major problem hampering our current understanding is that these new particles are smaller than the lower size detection limit of most instruments and are only observed after some particle growth has occurred.

Particles growth occurs through condensation of supersaturated vapours on the surface of the nucleated particles. This requires a lower degree of supersaturation than nucleation and thus condensation of the nucleating species reduces the rate of particle formation. Therefore, it is believed that particle growth often occurs through the condensation of other gases, including organic and inorganic compounds, than those responsible for nucleation. This decoupling of nucleation and growth means that the individual gases responsible for nucleation and growth can be unclear.

Since observations of particle formation only occur following growth to observable sizes it is possible that a pool of undetectable particles exist at all times but are only observed following significant condensational growth.

Keywords: Particle formation; Nucleation; Growth; Mechanisms; Modelling studies

FIGURE 1.13.2 An extract of an article in a scientific journal of a research report written by scientists. It is strictly structured with a pertinent title, names of authors, an abstract and the report using the convention of: introduction, method, results, analysis, conclusion and a reference list.

Access to scientific journals can be difficult as many journals require you to purchase a subscription or membership, though some are free. Figure 1.13.3 provides ideas for accessing scientific journals.

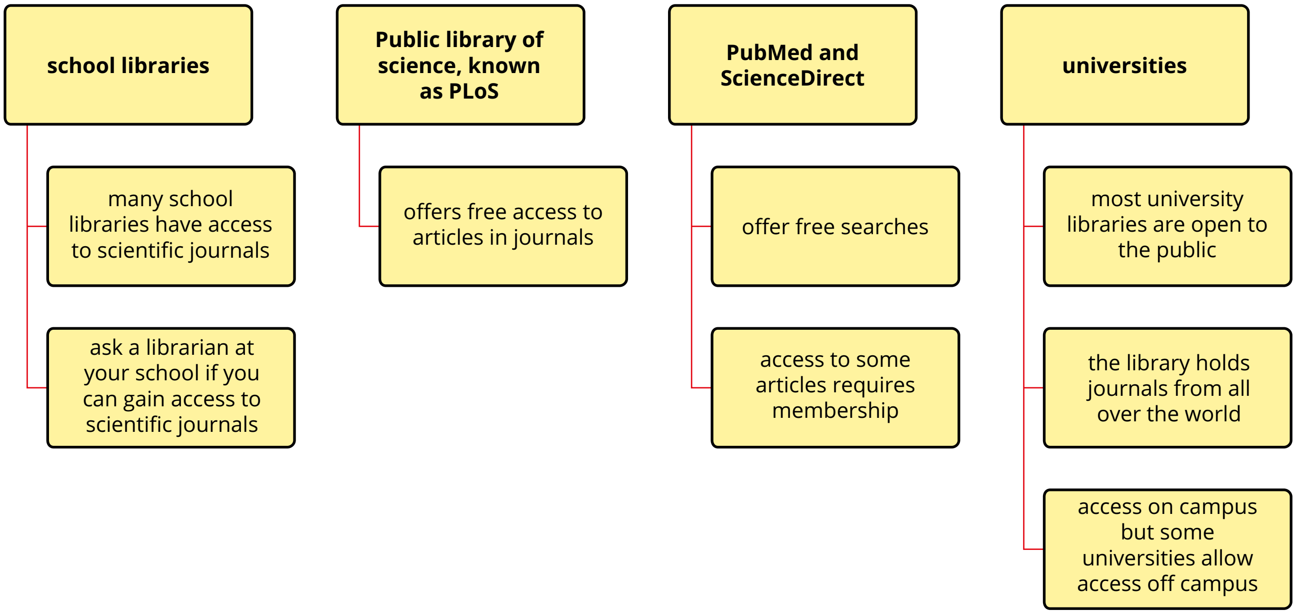


FIGURE 1.13.3 Suggestions and information on accessing scientific journal articles

Scientific articles are excellent sources of information, but also have their drawbacks as sources for the research investigation. Table 1.13.2 outlines some advantages and disadvantages of sourcing and using articles in scientific journals for the research investigation.

TABLE 1.13.2 Advantages and disadvantages of scientific articles

Advantages	Disadvantages
• written by experts • authoritative information (peer-reviewed) • most current information • logical, organised layout • content is relevant to the topic • contains an abstract that summarises all information in the article (if you do not find the information in the abstract, the article is not relevant) • primary source	• information is complex and challenging to understand • complex language and advanced processing and analysis of data • requires a high level of scientific literacy, language and numeracy to understand • may be time consuming to read and analyse • well-established concepts do not have recently published articles • may be difficult to locate

Books and physical publications

Secondary sources such as good science magazines and books are valuable sources of secondary information.

The first source you should use is your textbook. The language and concepts are presented specifically for high school students. In addition, the textbook addresses the syllabus objectives. Non-fiction books and magazines will probably be commonly used resources for the research investigation. Common commercial science magazines you might find in your school library include *The Helix*, *Cosmos*, *Scientific American* and *New Scientist* (Figure 1.13.4). Table 1.13.3 outlines some advantages and disadvantages of non-fiction books as sources for your research.

FIGURE 1.13.4 A science magazine you might find in your school library

TABLE 1.13.3 Advantages and disadvantages of book resources

Advantages	Disadvantages
• may be written by experts • potentially authoritative information • logical, organised layout • content is relevant to the topic • contain table of contents and index to help find relevant information • easily located in libraries • written in language that is understandable	• may not have been published recently • usable by only one person at a time • may have more bias than primary sources

Searching online

Online sources include online scientific organisations such as CSIRO, NCBI, AIMS, Bioscience (Oxford Academic) and other university publications and press. Many government and privately funded science organisations publish material, along with not-for-profit scientific organisations. Websites may direct you to magazines and scientific journals, such as those described above, the news, podcasts, blogs and videos (institutional, company and personal).

Information located on the internet requires very careful scrutiny. The openness and ease of publishing on the Internet means information may not be valid and reliable. Use the earlier guidelines above to help you evaluate sources. Table 1.13.4 outlines advantages and disadvantages of locating and using information online.

When searching for relevant information, you need appropriate search terms to enter into a search engine. These tips can help your search.

- Break your search statement into concepts and key words.
- Find synonyms, other related terms and concepts that apply to the topic.
- Create concepts of 1–3 words to enter into the search engine.
- Try different combinations of terms.
- Do not settle for the first sites on the list or your first attempt—look beyond the first page of results.
- Look through the results for sites from science organisations and research institutions (e.g. CSIRO, WEHI, NIH; .gov, .org), universities (.edu) and science journals and magazines.

TABLE 1.13.4 Advantages and disadvantages of internet resources

Advantages	Disadvantages
• quick and easy to access • allows access to hard-to-find information • access to the whole world; millions of websites • information is potentially more up-to-date • may be interactive and use animations to enhance understanding	• can easily sidetrack with non-relevant information • a lot of 'junk' sites and potentially more biased material • need to discern search engine results to find most useful sites • cannot always tell how up-to-date information is • can be hard to tell who has responsibility for authorship • information may not be well ordered • may not be reliable, valid or credible

Overview of resources

Your textbook should be your first source of reliable information. Other information should be consistent with this. Articles published in journals and magazines often present findings of new research, which may or may not be confirmed later, so be careful not to treat such sources of information as established fact. Scientific journals are peer-reviewed (critically reviewed by other specialist scientists), which gives them more credibility than other sources.

SKILLBUILDER 1.13.1

Evaluating sources for validity and reliability

Determining the validity and reliability of a source can be a challenging task. For some sources, it is easy to find details about the author, evidence and concurrency, while others only contain content and do not offer any other details. Another difficulty is when learning about a new topic or concept, we are all novice learners so it can be challenging to tell if a source is valid or not.

Table 1.13.5 outlines an example of evaluating a resource step-by-step, for a claim comparing herbal shampoos with commercial shampoos.

TABLE 1.13.5 Evaluation of a resource

Source evaluated: Formulation, evaluation and comparison of herbal shampoo with commercial shampoos			
Criteria		**Decision**	**Support/justification**
primary or secondary	Is this an eyewitness account or a second-hand source?	primary	• Research article published in www.sciencedirect.com • Research results published in article
validity	Does it contain information that is specifically related to the claim?	yes	• Outlines information directly related to the performance of herbal shampoos
	Is the evidence and information pertinent to the variables in the research question?	yes	• Information, experiment and results are specific to the performance and physiochemical properties of three shampoos
reliability	Is it current/recent information?	yes	• Published Dec 2014
	Is it up-to-date in its understanding of relationships?	yes	• Information outlined about the cleansing action of synthetic surfactants, physiochemical properties of shampoos (up to 2014)
	Is the evidence equivalent to other sources?	mostly	• Results are new and previously untested, therefore provides new knowledge never before seen
	Check the credibility—consider who the author is, their qualifications and expertise.		• Authors Khaloud Al Badi and Shah A. Khan are scientists from the Department of Pharmacy, Oman Medical College, Oman • *Journal* of *Basic and Applied Sciences* is a reputable peer-reviewed journal
	Try to find the sample size.	known	• 20 volunteer researchers
	Try to establish what variables were controlled or measured.	known	• Methodology of the testing for each of the physiochemical properties was controlled • Methodology of the testing on the volunteers was consistent

A judgement could be made about the credibility of this source, such as:

The information and evidence was published by experts in a peer-reviewed journal article that is concurrent and with variables of experimentation known and directly related to the claim and research question. Though, as the results are new, it is yet to be seen if they are substantiated, therefore affecting the reliability of the evidence. This resource is both valid and reliable.

1.13 Review

SUMMARY

- Scientific evidence for the research investigation can be sourced from numerous publications including scientific journals, scientific articles and commercial articles.
- A primary source is written by the observer/witness of an event, or the scientist who conducted the research.
- A secondary source is a document that refers to or analyses a primary source.
- It is important that all sources are evaluated for their credibility, validity and reliability.

KEY QUESTIONS

Retrieval

1 Define the following.

a a primary source

b a secondary resource

Comprehension

2 Discuss whether each of the following is a primary or a secondary source.

a a newspaper article about climate change

b an experiment to investigate chemical changes when mixing combinations of chemicals

c an interview with a forensic scientist about using science in tracking criminals

d a website with information about the concentration of atmospheric pollutants emitted from a vehicle

Analysis

3 You are learning about the use of fuel cells and are searching for information about fuel cells in space craft. Assess from the options below which is the best resource to use. Explain your answer.

A the book *Fuel Cells: and their applications,* published in 2006

B the article ‘Batteries and Fuel Cells in Space’ by G. Halpert, published by the Electrochemical Society in 1999

C the website www.nasa.gov/topics/technology/hydrogen/hydrogen_2009.html, accessed on 30 October 2017

1.14 Research: taking and organising notes

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- use your scientific journal to take and organise notes, and to refine experimental procedures
- paraphrase information found in primary and secondary sources.

As previously mentioned, a journal should be maintained, recording all the ideas, research and developments of the research investigation. Once all the work is complete, the information in the journal should form the basis of your response. There is no need to produce new work as it is already completed in the journal.

Scientists organise their journal notes with the following general features:

- recording date of journal entries
- recording journal entries for:
  - ideas, observations, proposals and questions
  - research background information for ideas, observations, proposals and questions
  - refinements
  - personal explanations of information, concepts, ideas, observations, proposals and questions (this often includes diagrams)
  - results, data and evidence
- recording the origin of information (recording sources).

RECORDING DATE OF JOURNAL ENTRIES

Journals usually record the date of entries and a title. These titles are often related to their own work and objectives (categories) rather than the title of the source. The title of the source is recorded when the origin is recorded.

Always place the date at the beginning of the work you are doing. This is done by professionals worldwide to catalogue and file information. When trying to find previous work completed, most people search their memory as to ‘when’ they completed the work. They think ‘I’m sure I did that research of “Y” after I found the information on “X” last week’. Always date your work and research; it will become a simple yet effective filing system.

RECORDING JOURNAL ENTRIES

Each journal entry should follow a cataloguing system; most typically this is the date and a title. A reference point or cataloguing system helps you to find information when you are searching for information later.

Ideas, observations, proposals and questions

You may be surprised how often simple ideas, observations, proposals and questions influence, direct and help your research days or weeks later.

Your ideas, observations, proposals and questions could be related to:

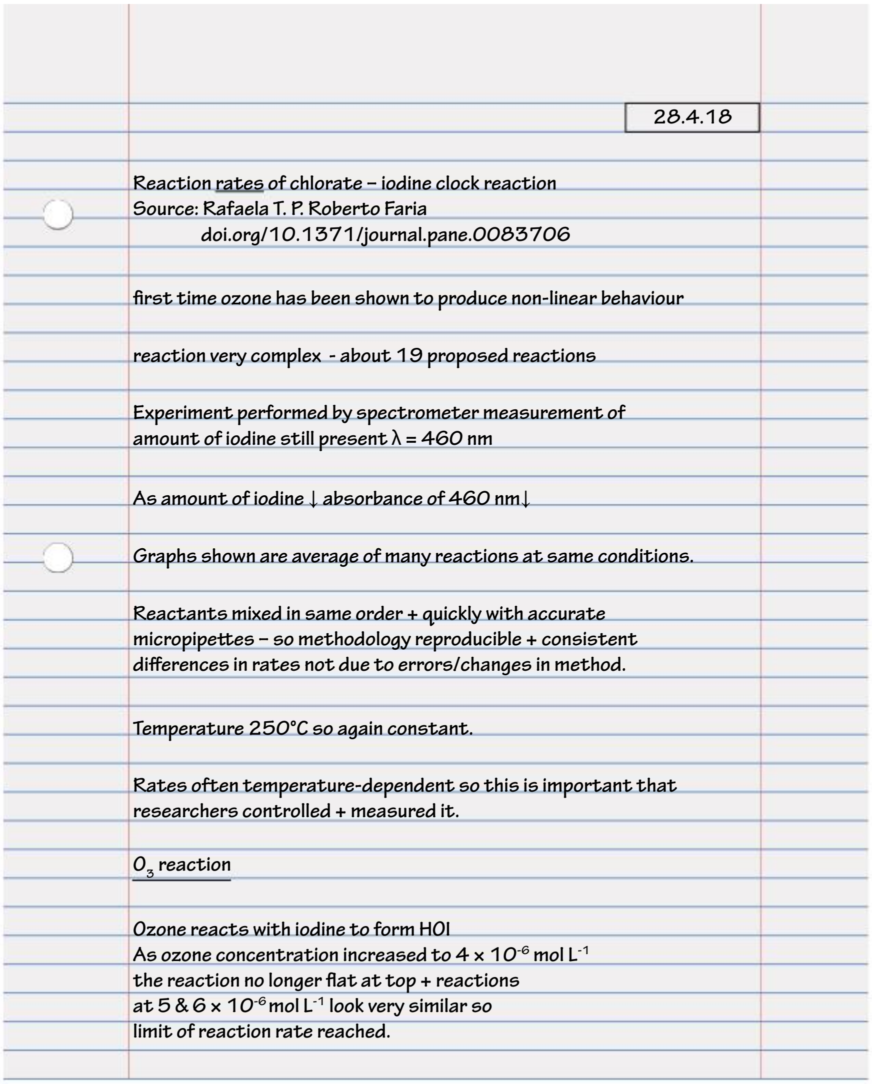
- the variables or concepts involved in your investigation
- new terms you were exposed to and do not understand
- data or evidence you do not currently comprehend but are important to a part of the investigation
- statistics about significance in the analysis of some information
- ideas and proposals about possible future research and questions.

Record all of these in your journal. The value and benefit they may offer later could include:

- saving you time by:
  - not researching the same idea twice
  - helping you to link ideas from one day or week to the next when you return to your work
  - suggesting guidance and pathways for research and queries that are related
- providing links between concepts in the future that are currently unknown
- providing answers (or partial answers) to future issues, conceptual blocks and questions
- developing understanding of unrelated concepts that become pertinent later.

The content and recordings in your journal will not be in a logical conceptual order, as are your class notes or a textbook. This is because you are researching unfamiliar knowledge and understanding, and you will not know how it all fits together until the research is complete.

Before you add an entry to your journal, make sure you have recorded the date and provided your own title for the entry. As you take notes from different sources, always record the source information; for example, the title, web address, author and page. Figure 1.14.1 is an example of a journal entry.



28.4.18

Reaction rates of chlorate – iodine clock reaction
Source: Rafaela T. P. Roberto Faria
doi.org/10.1371/journal.pane.0083706

first time ozone has been shown to produce non-linear behaviour

reaction very complex - about 19 proposed reactions

Experiment performed by spectrometer measurement of amount of iodine still present λ = 460 nm

As amount of iodine ↓ absorbance of 460 nm ↓

Graphs shown are average of many reactions at same conditions.

Reactants mixed in same order + quickly with accurate micropipettes – so methodology reproducible + consistent differences in rates not due to errors/changes in method.

Temperature 250°C so again constant.

Rates often temperature-dependent so this is important that researchers controlled + measured it.

O_3 reaction

Ozone reacts with iodine to form HOI
As ozone concentration increased to 4×10^{-6} mol L^{-1}
the reaction no longer flat at top + reactions
at 5 & 6 × 10^{-6} mol L^{-1} look very similar so
limit of reaction rate reached.

FIGURE 1.14.1 This journal entry records the source information, for example, the title, web address, author and page, and interpretation of the data.

Research background information

Information taken from a source should be rewritten or summarised in your own words in the journal. Avoid copying information verbatim so you are not tempted to plagiarise when you write up the research investigation.

It can be difficult sometimes to reword sources into your own words, especially if it is already expressed well and concisely. Before you write and record the research information, read the material and grasp its understanding. Without referring back to the source, write notes in your journal, in your own words. Use multiple sources and a dictionary as references for information. In your journal, your notes should be detailed with extended explanations that you will fully comprehend when you refer back to them at a future date. See the worked example below for a few examples of rewording information.

Worked example 1.14.1

REWORDING INFORMATION FROM SOURCES

Resource for information states that:

'It can be expected that the days will be warm and the nights will be cool over the next three days due the high pressure system moving across the region.'

Reword this information into your own words.

Thinking	**Working**
Can swap the order of information, the object (the warm days and cool nights) with the reason (high pressure system).	Due to the high pressure system moving across the region, the days will be warm while the nights will be cool.
Can change numerous words such as adjectives, verbs, adverbs and nouns.	It has been predicted that during the daytime it will be warm while the evenings will be cool for about three days, as a high pressure system travels over the region.
Can elaborate on the information to explicitly outline explanations.	The weather forecast expects (made a prediction based on computer simulations) that the temperature during the day will be warm (22–25°C) and the evenings will be cool (below 12°C) due to a high pressure system (weak high pressure system), which dissipates clouds and allows direct sunshine to warm the Earth's surface; however, in the evenings this system brings cool air from higher altitudes, which is no longer warmed by the sun.

➤ Try yourself 1.14.1

REWORDING INFORMATION FROM SOURCES

A resource for information states that:

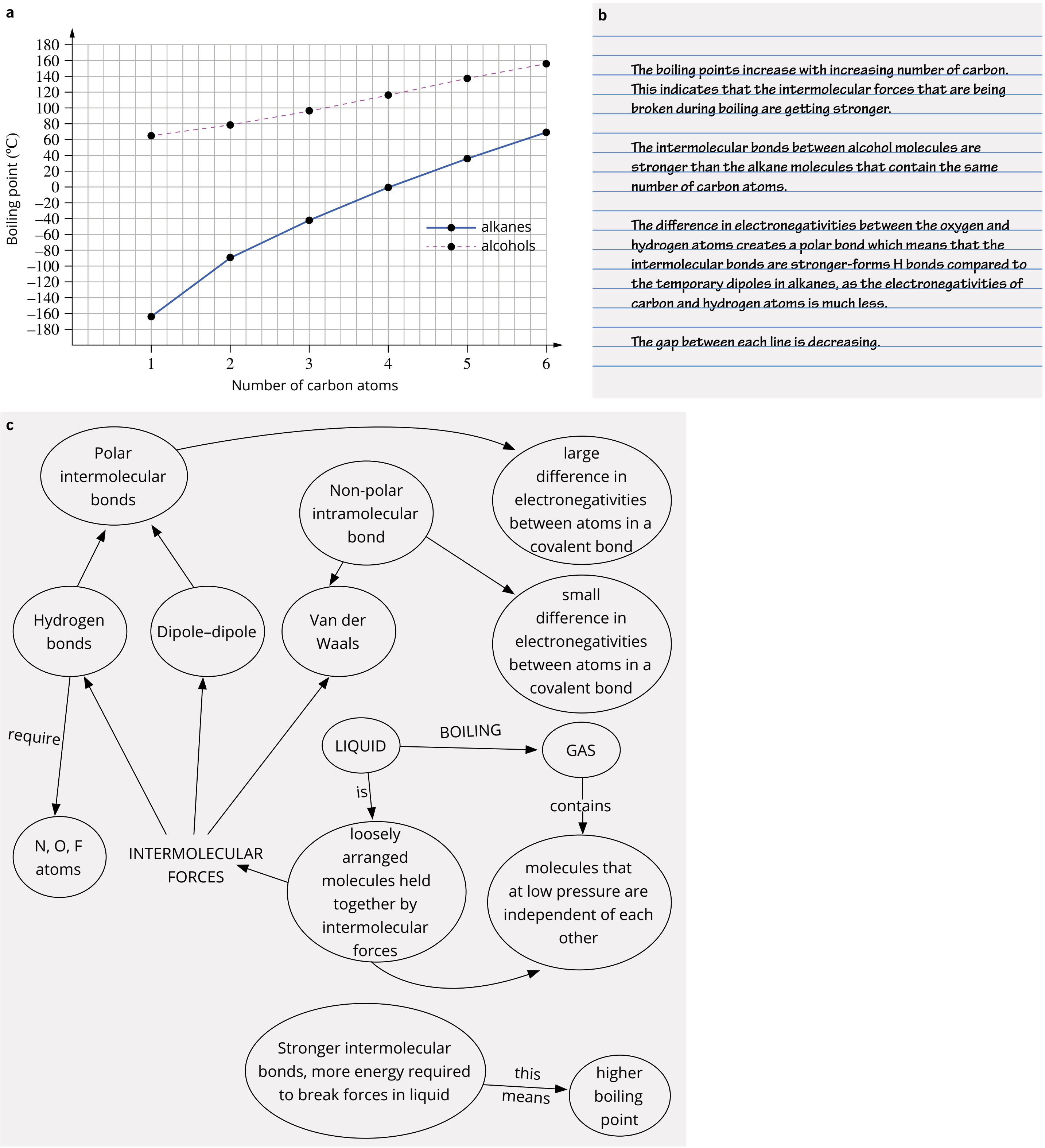
'This will be followed by a low pressure system that will bring cooler days and not so cool nights.'

Reword this information in your own words.

Personal explanations and interpretation

The journal is a valuable tool that is personal to you. As such, it should be written and displayed in formats that you understand and that make meaning of information to produce the research investigation report. It can include writing, drawings and schematics (e.g. flow charts).

Sometimes, one form of recording information provides more details than others. Details are not limited to just facts, they can be the order of events, their location or their physical orientation. Therefore use a variety of ways to record or present the information to encapsulate more comprehensive details. Figure 1.14.2 presents the same information in a few different ways.



The boiling points increase with increasing number of carbon. This indicates that the intermolecular forces that are being broken during boiling are getting stronger.

The intermolecular bonds between alcohol molecules are stronger than the alkane molecules that contain the same number of carbon atoms.

The difference in electronegativities between the oxygen and hydrogen atoms creates a polar bond which means that the intermolecular bonds are stronger-forms H bonds compared to the temporary dipoles in alkanes, as the electronegativities of carbon and hydrogen atoms is much less.

The gap between each line is decreasing.

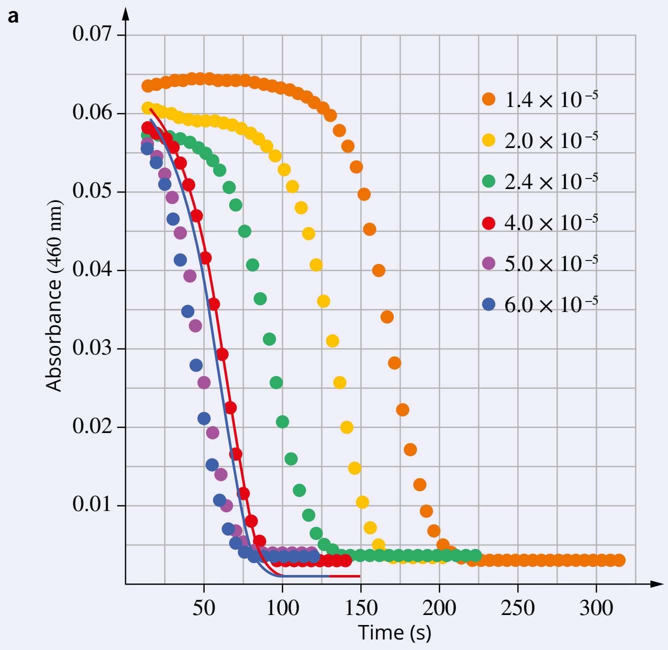
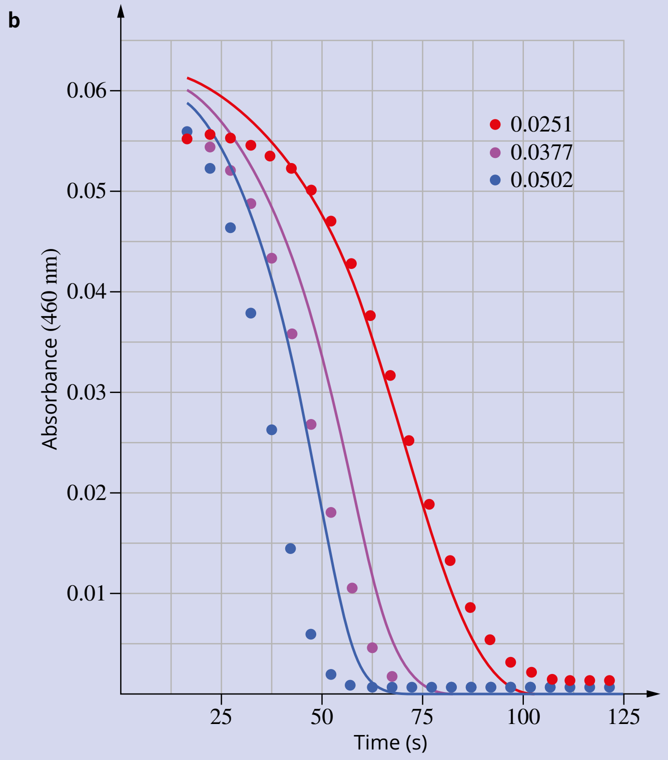
FIGURE 1.14.2 (a) Graphical representation of the boiling point trend for alcohols and straight chain alkanes. (b) Written comparison of the boiling point trend. (c) Concept map linking the electronegativities of atoms to the boiling points

Results, data and evidence

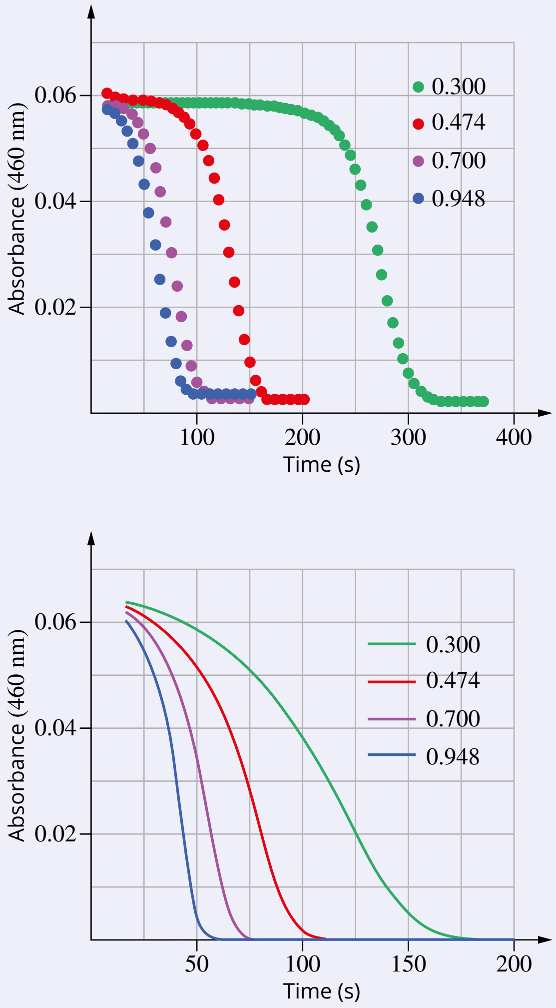
When you read scientific results, data or evidence in a resource, it is important to record the values or reproduce them in the journal. Table 1.14.1 shows examples of research evidence and how they can be reinterpreted for a journal entry. Along with the results, data or evidence includes:

- specific details and your own interpretation of significant values
- the trend or pattern
- the comparison or difference between one set of values and another
- the statistics used to establish significance and also the author's interpretation.

TABLE 1.14.1 Examples of interpretations of research evidence

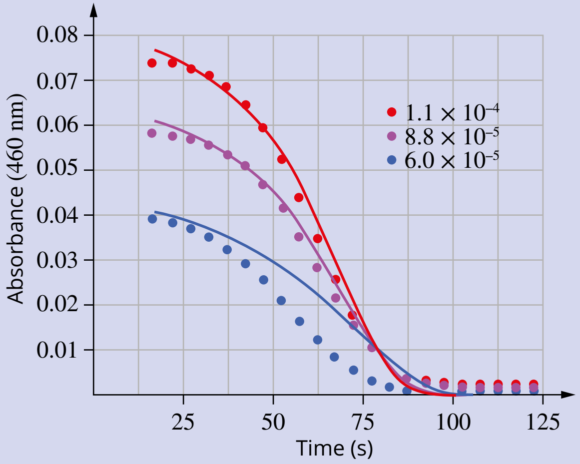
	Effect of the ozone concentration on the clock time. The initial concentrations after mixing the reactants in the cuvette were as follows: $[O_3]$ as indicated in the figure (mol L^{-1}), $[NaClO_3] = 0.0251$ mol L^{-1}, $[HClO_4] = 0.474$ mol L^{-1}, and $[I_2] = 8.8 \times 10^{-5}$ mol L^{-1}. Experimental data (symbols); modelling results (continuous line).
	Effect of the chlorate concentration on the clock time. The initial concentrations after mixing the reactants in the cuvette were as follows: $[NaClO_3]$ as indicated in the figure (mol L^{-1}), $[O_3] = 3.0 \times 10^{-6}$ mol L^{-1}, $[HClO_4] = 0.474$ mol L^{-1}, $[I_2] = 8.8 \times 10^{-5}$ mol L^{-1}. Experimental data (symbols); modelling results (continuous line).

c



Effect of the acid concentration on the clock time. The initial concentrations after mixing the reactants in the cuvette were as follows:
[$HClO_4$] as indicated in the figures (mol L^{-1}),
[O_3] = 2.0 × 10^{-6} mol L^{-1},
[$NaClO_3$] = 0.0251 mol L^{-1},
[I_2] = 8.8 × 10^{-5} mol L^{-1}.
Experimental data (symbols); modelling results (continuous line).

d



Effect of the iodine concentration on the clock time. Initial concentrations after mixing the reactants in the cuvette:
[I_2] as indicated in the figure (mol L^{-1}),
[$NaClO_3$] = 0.0251 mol L^{-1},
[$HClO_4$] = 0.474 mol L^{-1},
[O_3] = 4.0 × 10^{-6} mol L^{-1}.
Experimental data (symbols); modelling results (continuous line).

When making journal entries it is important for your understanding and analysis of statistical method (e.g. R^2 value, **confidence interval** or standard deviation) that you express information in your own words. You will most likely have to conduct further research to understand the statistics, their meaning, and the author's interpretation of the results as you will come across new statistical calculations you have not seen previously.

Recording the origin of information

As you conduct your investigation and research, it is important to write down the source of the information. This will enable you to return to the source later to continue researching, collect further information or recheck details. You will also be required to produce a reference list in the report, and it will save time if you have already recorded the source of the information.

Record the following information about sources you use.

TABLE 1.14.2 Essential information about your sources

Books	Online
• authors • title • date of publication • publisher • place of publication • pages	• authors or name of organisation • title • date website was written or updated • date website was accessed • website address (URL)

1.14 Review

SUMMARY

- There is no single method for taking and organising notes.
- A scientific journal enables scientists to record and organise their notes.
- Scientists continually revise and refine their experiments based on previous experimentation and results.

KEY QUESTIONS

Retrieval

1 State how you record information from a source in your journal.

Comprehension

2 Explain the benefit of keeping a scientific journal.

Analysis

3 Apply your skill of paraphrasing and rewrite the following text in your own words.

In 2010, the Deepwater Horizon (DWH) oil spill released five million barrels of oil and gas into the Gulf of Mexico. The West Falmouth oil spill in 1969 was on a smaller scale but resulted in petroleum contaminants that persisted for decades in the sediments. Global and local environments are threatened by oil spills over both short and long time scales. The impacts of crude oil and refined petroleum products on the environment deserve significant investigation. Most of these studies focus on the major components in crude oil, i.e. the non-polar hydrocarbons that contain only carbon and hydrogen.

1.15 Writing a report for the research investigation

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- write a report for your research investigation
- write a literature review for your research investigation.

There are several genres that can be used to report a research investigation. As a written report, the genres include a literature review, empirical essay and annotated bibliography. The report must be 1500–2000 words long. Alternatively, a poster can be used as part of a multimodal presentation, of 9–11 minutes in length. This is not an exclusive list. Even though there are many ways to present your report, it must follow a scientific genre and meet the requirements of the QCAA Chemistry General Senior Syllabus.

Writing a research investigation report requires scientific communication and will need to follow scientific genre conventions. An explanation of these requirements has been provided in Module 1.11. A report will communicate information in a logical sequence (introduction, body and conclusion) and may contain subheadings. Many of the characteristics of the student experiment are closely related to the research investigation. The focus of the student experiment is on interpreting data collected during a scientific investigation, while the focus of the research investigation is to explore a claim.

Characteristics of the genres for presenting the research investigation are summarised in Table 1.15.1. This module will focus on the literature review.

TABLE 1.15.1 Summary of features of presentation genres for the research investigation

Science presentation genre	Brief description	Features
literature review	• a report that evaluates information found in a publication, related to the selected topic • gives a theoretical base for the claim (question), and analyses and interprets information/data related to the claim • objective is to point out strengths and weaknesses of the claim	• abstract • introduction to topic providing context • discussion of information/data • analysis of data/information • evaluation of information/data • conclusion related back to claim • presented as paragraphs that flow on in a logical development of ideas • uses subheadings
annotated bibliography	• notes, comments and explanations about article/s • evaluation of a claim after investigating how other sources treat the claim	
poster	• presents ideas concisely and clearly • an oral presentation accompanies the poster • ideas are elaborated upon in the oral presentation	• includes all the above • visual, oral and text presentation • use of font size, colour, dot points, subheadings, logical flow of information to effectively deliver information

LITERATURE REVIEW

A literature review usually includes an introduction, a body and a conclusion. It critically analyses information and convinces the reader of the significance or importance of the topic being investigated. This is achieved through the presentation of information in a logical sequence in which the author guides the reader through the material to understand its significance. The research question provides the foundational direction and guidance of the literature review.

Qualities found in a literature review

Not all qualities and elements of a full literature review will be appropriate for the research investigation. The literature review will be limited to the word count and ISMG descriptors in the syllabus. Depending on the research question and the topic being investigated, choose to cover some of these qualities of a literature review.

- Determine the current understanding on the topic.
- Provide an overview of key concepts (relevant to the research question).
- Identify important relationships between variables (specific to variables that influence the independent or dependent variables stated in the research question).
- Identify strengths and weaknesses of evidence in the information used for the above points.
- Identify any gaps in the research.
- Identify any conflicting evidence.

One of the key qualities of a literature review is that it is a critical analysis of the evidence to communicate a true understanding of the 'big picture' about a topic. It does not just summarise information. Science is more about models, theories laws and principles, and these are continually developing. The literature review is used to critically analyse evidence, both strengths and weaknesses, as well as the gaps and conflicts to convince the reader about the current state of a large jigsaw puzzle conceptual relationship.

Qualities of a literature review include:

- critically analysing the evidence—establish what it means and account for the statistical processing of data used (be sure to analyse the methodologies, samples, results and data processing and analysis)
- discriminating between relevant and irrelevant sources and content—only use the relevant in the literature review
- organising the structure of the review to logically inform and convince the reader—plan the introduction, body and conclusion.

There is no hard and fast rule about the number of articles that is sufficient to consult for the research investigation. The term *sufficient* is defined by the QCAA Chemistry General Senior Syllabus as 'enough or adequate for the purpose'. The purpose of the research investigation is to evaluate the claim; therefore, if the number of articles enables a justified conclusion to be drawn about the research question and explores both sides of the argument, then the number of sources is sufficient. Relevant research is that which is connected to the rationale and unit under study.

A thorough analysis requires complete attention to every detail and therefore the research should include analysis and synthesis from different sources. It should identify patterns, trends and relationships that are related to the investigation. Also, it should be explicitly connected to the research question as well as clarifying where the sources agree and disagree.

The analysis should provide a complete identification of the limitations of the research as this may have a bearing on the extent to which information (primary or secondary) is valid to the research question. It is important that when analysing data and its limitations, it is not done so selectively, to demonstrate what the author wants to show. Such bias will result in producing an inappropriate and erroneous conclusion. Quality scientific analysis is open to any result.

Justified scientific arguments are those that are supported by sound reasons or evidence. Therefore, you must apply your scientific understanding and conceptual knowledge to the evidence that you have examined. The following factors should be examined and discussed.

- State whether a pattern, trend or relationship was observed between the independent and dependent variables.
- Describe what kind of pattern it was and specify under what conditions it was observed.
- Note and explain any **deviations** in the data or information.
- Identify any limitations in the data or information researched. Why and how do these limitations affect the validity of the research question or conclusion?

An analysis of information may also include an evaluation of the methodology used by authors to obtain their data or information.

Conclusion and evaluation

The discussion needs to include an understanding of the features of the evidence that limit its ability to be used. Is the data from measured samples or from estimated models? Did the studies occur under different conditions or categorise the independent variables differently so comparison was impossible?

The above questions require a justified discussion evaluating the reliability and validity of evidence. Therefore, it is important to discuss the limitations of each source. You can do this by:

- evaluating the method of evidence collection
- identifying issues that could affect validity, accuracy, precision and reliability of evidence
- stating sources of systematic and random errors
- recommending improvements to the evidence to improve validity.

The discussion must recommend improvements to the investigation that are linked to the evidence and would address the limitations and gaps in knowledge that have been identified during the research investigation. The suggestions presented must be connected to the claim and allow further investigation.

Communication

This section will provide a guide to writing an appropriate scientific report and some of the general conventions that need to be followed for a literature review.

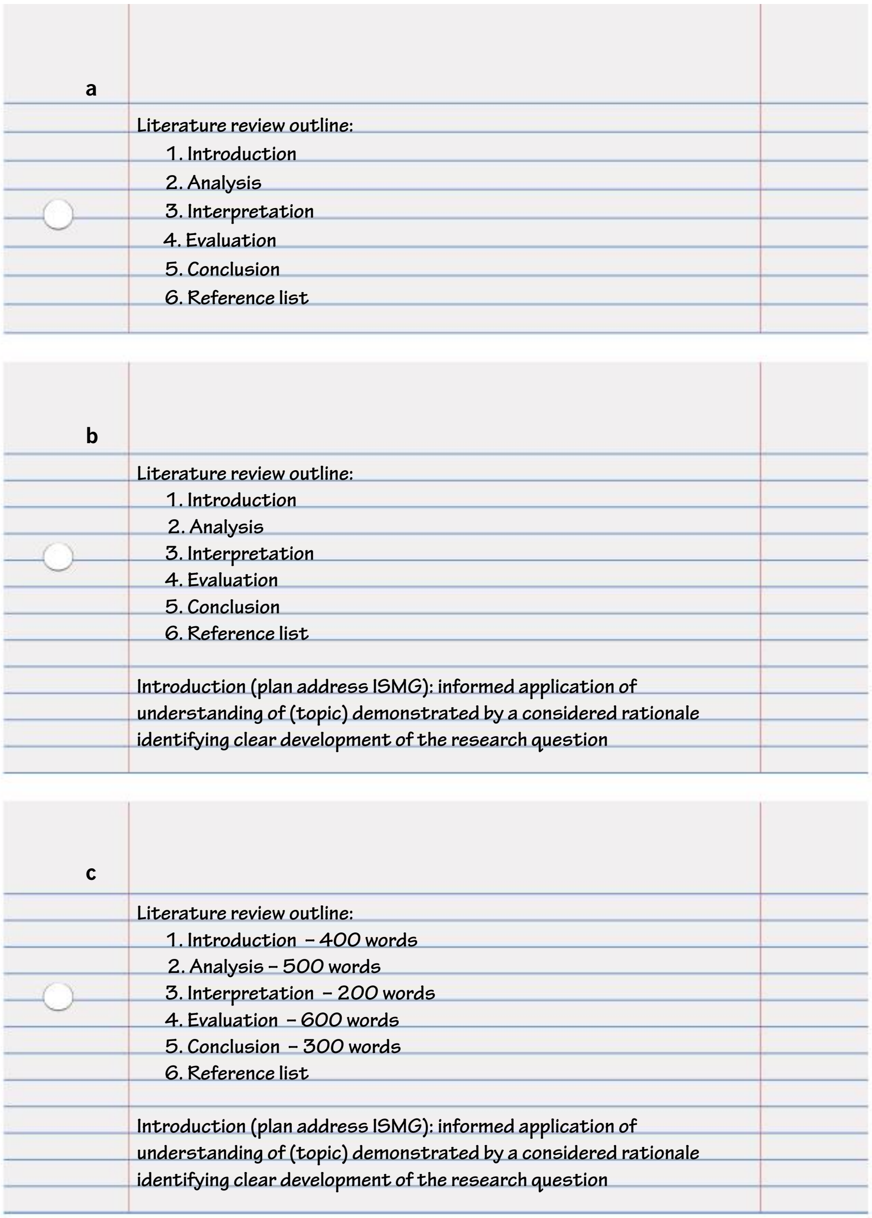
Once a plan for the literature review or your preferred presentation mode has been developed, word limit guides can be assigned to each section (Figure 1.15.1 on page e124). The notes in your research journal will give you a good idea of which sections will require more words for explanations than others. The journal notes will provide information so you can make decisions on which figures, data and evidence to include and also how many words may be required to elaborate on the evidence. Distribute the total word count across all sections you plan to include in the report.

The word guide is not binding. As you complete your review you can alter the word guide and distribution of the word count if you feel it is necessary. However, make sure the limit outlined in the syllabus is never exceeded.

Planning the research investigation report will help address the syllabus ISMG.

Structure of a written report

Although the use of headings in scientific reports is essential to guide and direct scientists to particular information, there is no single correct convention for a scientific report. A typical structure includes an introduction, body, conclusion and reference list.



a

Literature review outline:
1. Introduction
2. Analysis
3. Interpretation
4. Evaluation
5. Conclusion
6. Reference list

b

Literature review outline:
1. Introduction
2. Analysis
3. Interpretation
4. Evaluation
5. Conclusion
6. Reference list

Introduction (plan address ISMG): informed application of understanding of (topic) demonstrated by a considered rationale identifying clear development of the research question

c

Literature review outline:
1. Introduction – 400 words
2. Analysis – 500 words
3. Interpretation – 200 words
4. Evaluation – 600 words
5. Conclusion – 300 words
6. Reference list

Introduction (plan address ISMG): informed application of understanding of (topic) demonstrated by a considered rationale identifying clear development of the research question

FIGURE 1.15.1 Planning the report: (a) a planned outline of the research investigation report to ensure the syllabus ISMG is addressed; (b) ISMG descriptors assigned to the sections in the planned outline; (c) the word count guide planned

Edit the report after you have completed it. Editing is an important part of the process. After editing your report, save new drafts with a different file name and always back up your files in another location. Pretend you are reading your report for the first time when editing. When reading your own work, do not read it as what you intended to write, but what you have actually written. Reading the report out aloud to yourself will help you read it more critically.

When editing, look for content that is:

- ambiguous or unclear
- repetitive
- awkwardly phrased
- too lengthy
- not relevant to your research question
- poorly structured
- lacking evidence
- lacking a reference (if it is another researcher's work).

1.15 Review

SUMMARY

- There are several genres that can be used to report a research investigation, such as a literature review, empirical essay, poster presentation and annotated bibliography.
- A literature review critically analyses information and convinces the reader of the significance or importance of the topic being investigated.

KEY QUESTIONS

Retrieval

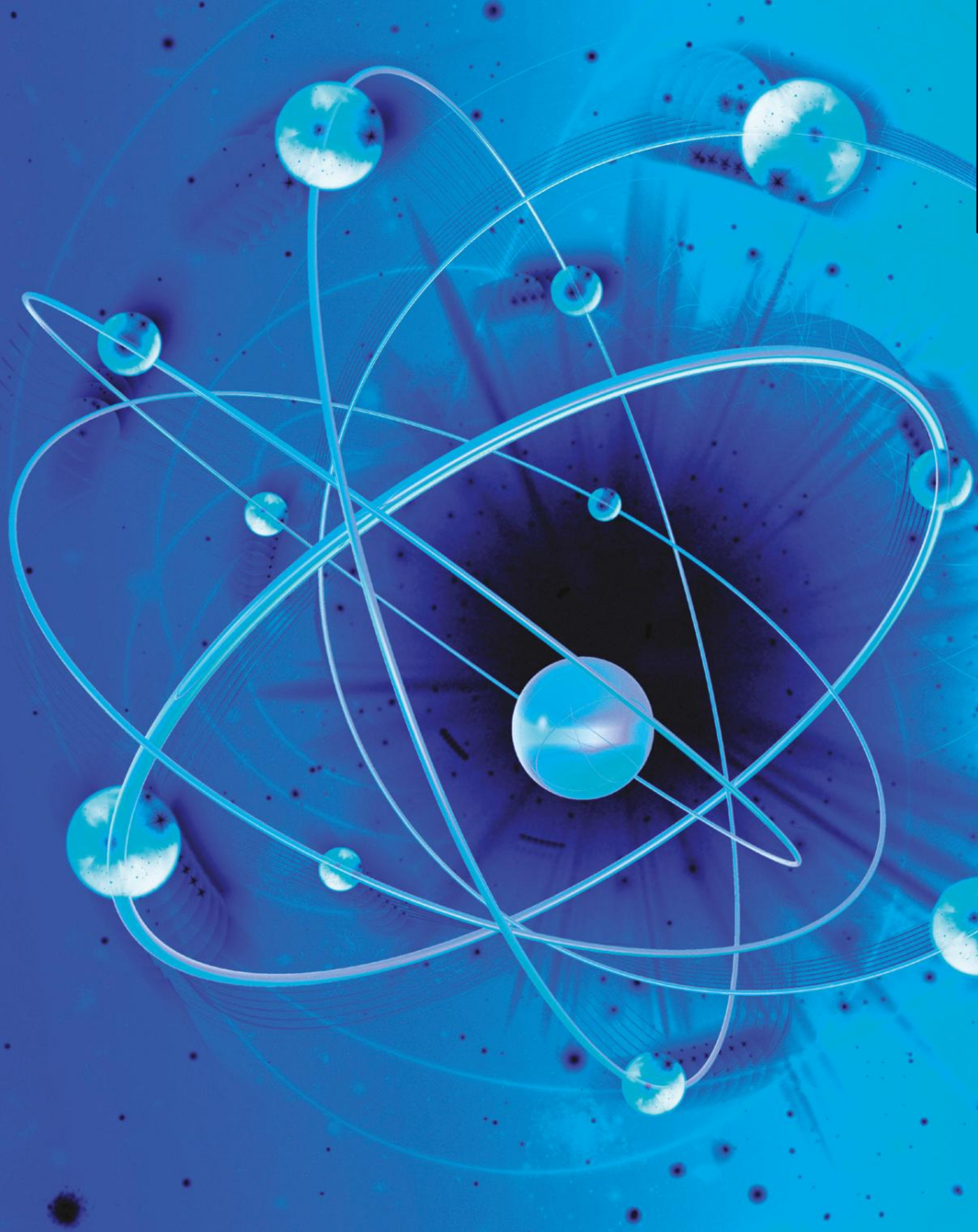
1 Describe the structure of a research investigation report.

Comprehension

2 Explain the purpose of the research investigation.

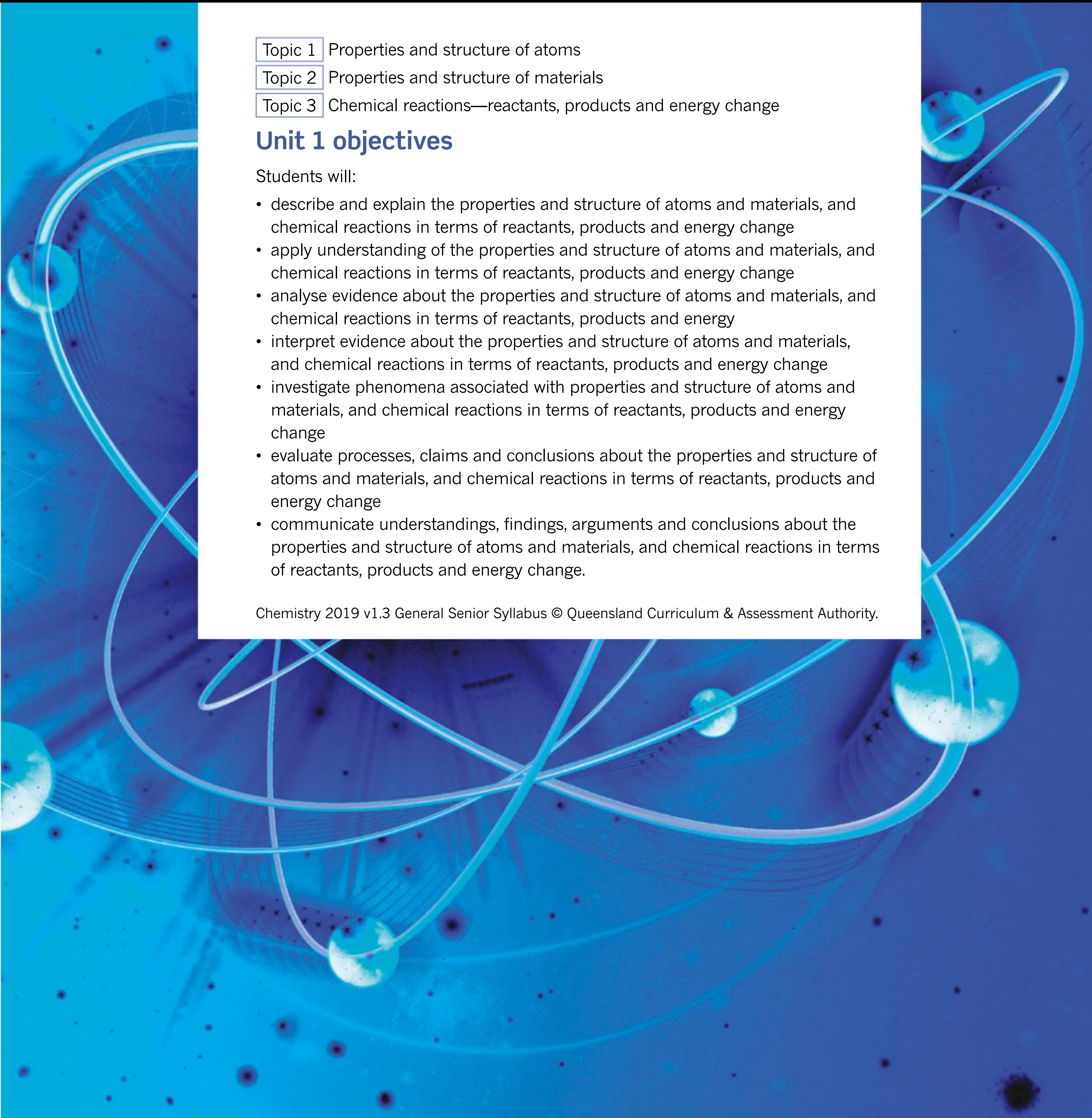
Analysis

3 Determine how you would achieve an analysis of the methodologies, samples and results in a research report.



UNIT 1

Chemical fundamentals—structure, properties and reactions



Topic 1 Properties and structure of atoms

Topic 2 Properties and structure of materials

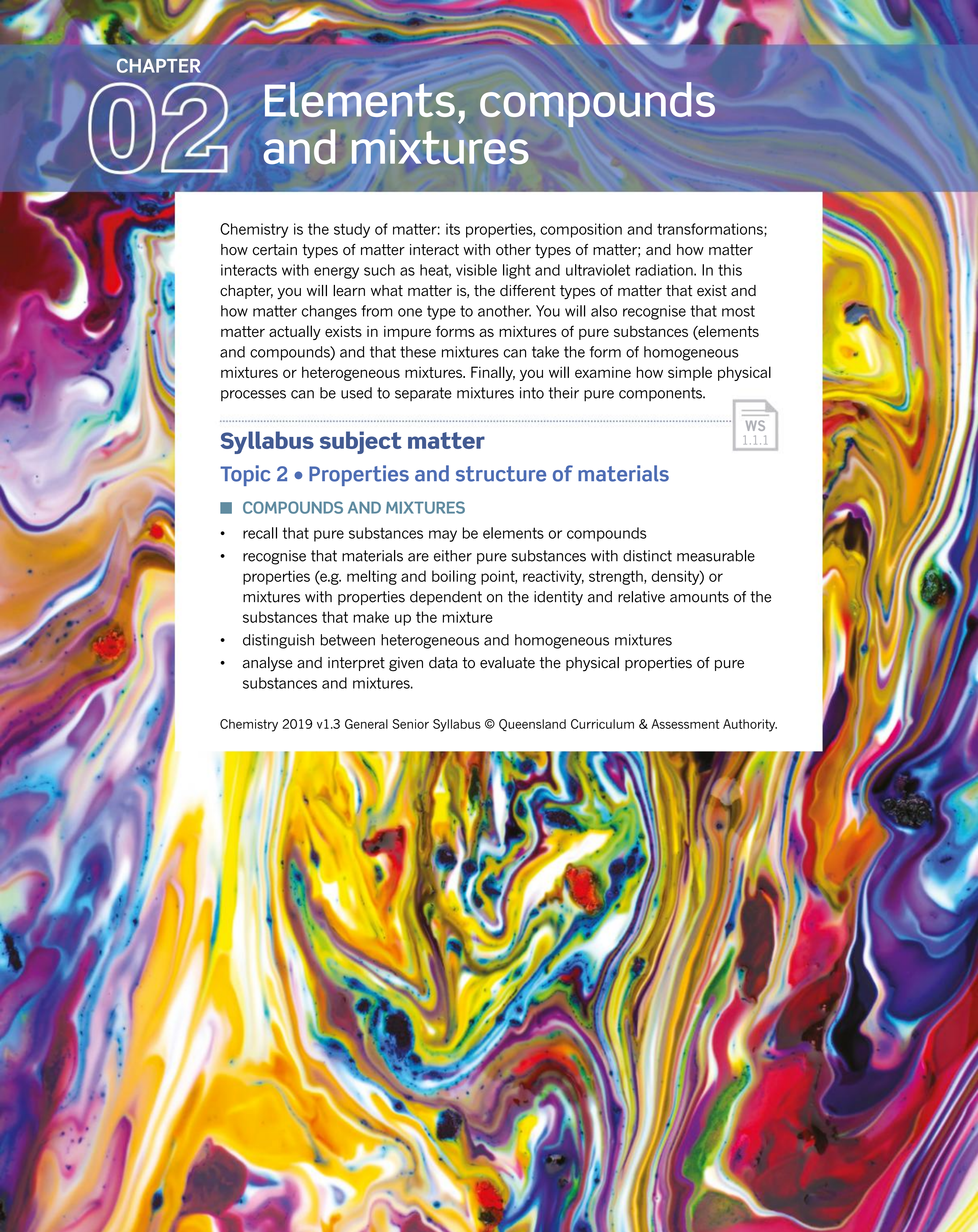
Topic 3 Chemical reactions—reactants, products and energy change

Unit 1 objectives

Students will:

- describe and explain the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
- apply understanding of the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
- analyse evidence about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy
- interpret evidence about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
- investigate phenomena associated with properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
- evaluate processes, claims and conclusions about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
- communicate understandings, findings, arguments and conclusions about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.



CHAPTER

02 Elements, compounds and mixtures

Chemistry is the study of matter: its properties, composition and transformations; how certain types of matter interact with other types of matter; and how matter interacts with energy such as heat, visible light and ultraviolet radiation. In this chapter, you will learn what matter is, the different types of matter that exist and how matter changes from one type to another. You will also recognise that most matter actually exists in impure forms as mixtures of pure substances (elements and compounds) and that these mixtures can take the form of homogeneous mixtures or heterogeneous mixtures. Finally, you will examine how simple physical processes can be used to separate mixtures into their pure components.

WS 1.1.1

Syllabus subject matter

Topic 2 • Properties and structure of materials

COMPOUNDS AND MIXTURES

- recall that pure substances may be elements or compounds
- recognise that materials are either pure substances with distinct measurable properties (e.g. melting and boiling point, reactivity, strength, density) or mixtures with properties dependent on the identity and relative amounts of the substances that make up the mixture
- distinguish between heterogeneous and homogeneous mixtures
- analyse and interpret given data to evaluate the physical properties of pure substances and mixtures.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

2.1 Characterising matter

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that matter can be characterised by its purity
- understand that most matter you encounter in your everyday life is a mixture of pure substances
- recall that mixtures may be homogeneous or heterogeneous
- recognise that mixtures are materials where the properties are dependent on the identity and relative amounts of the substances that make up the mixture
- recall that pure substances are either elements or compounds
- recognise that pure substances have a definite and distinct set of physical and chemical properties.

Chemistry is the study of **matter**, so it is important to understand the different types of matter that exist. You know from everyday experience that a tree, a rock, a glass of water and a piece of gold are examples of matter. You also intuitively know that there are fundamental differences in the observable properties of trees, rocks, water and gold that tell us that they are different types of matter. However, you can identify some properties common to all types of matter. For example, you can see the effect of matter on other matter; think of the book you are reading or the screen you are viewing; think of the wind on your face, the sand between your toes or the water in your bath tub. All are examples of matter.

Another characteristic feature of matter is that you can measure it. Matter has **mass** and you can measure this **physical property**; matter also occupies space and you can measure its **volume**. The following statement is a good working definition of matter that will suit our purposes for studying chemistry.

> ℹ Matter can be described as anything that has mass, occupies space and can be perceived by our senses.

PURITY OF MATTER

Matter can be classified, or characterised, in different ways. One way is to look at the purity of matter. It turns out that most of the matter you encounter in your everyday life—including the food you eat, the air you breathe and the water you swim in—is not chemically pure. Most matter actually consists of **mixtures** or **pure substances**. For example, the air you breathe is a mixture of oxygen and nitrogen with trace amounts of other gases, including carbon dioxide, water vapour and argon. Even tap water may appear to be pure but it actually contains trace amounts of dissolved minerals.

The relationship between pure substances and mixtures is shown in Figure 2.1.1. It shows that the matter you observe in your everyday life is ultimately composed of either **elements** or **compounds**. Collectively, elements and compounds are known as pure substances. Pure substances can be physically combined to produce mixtures. Mixtures can either be **homogeneous mixtures** or **heterogeneous mixtures**. The differences between these two types of mixtures will be discussed in detail in Module 2.2 of this chapter.

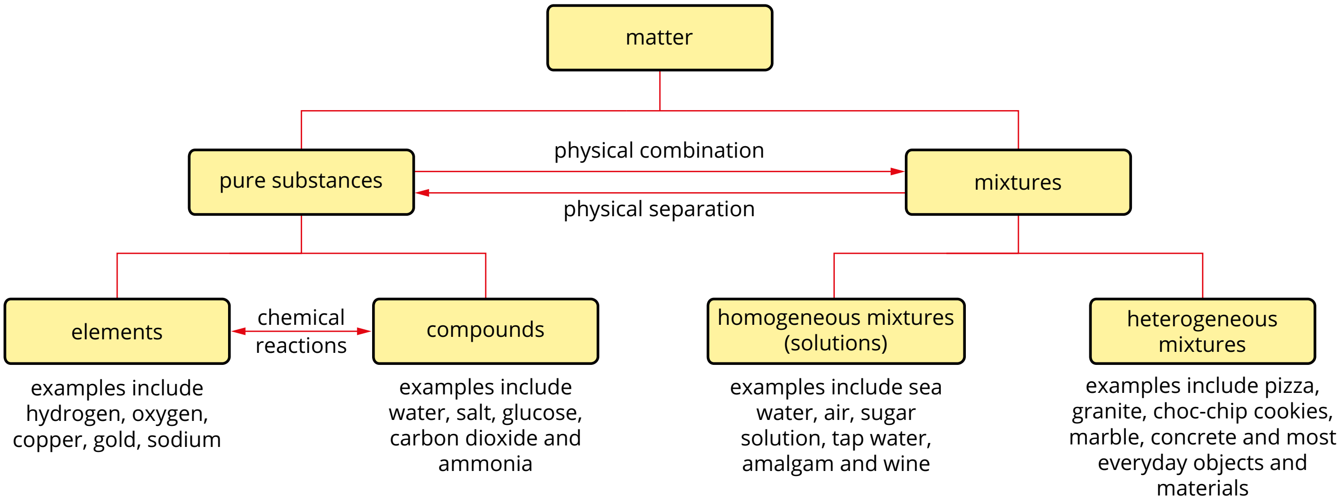


FIGURE 2.1.1 Classification of matter according to purity, showing the relationship between elements, compounds and mixtures

Figure 2.1.2 shows four examples of different types of matter.

- A slice of pizza (Figure 2.1.2a) contains a mixture of carbohydrates, fats and oils, as well as water and dissolved minerals and nutrients. It is a physical mixture of a wide range of pure substances. It also contains visibly distinct 'chunks' that are different from other parts, such as the pepperoni slices. This gives us the hint that a slice of pizza is a heterogeneous mixture.
- Food colouring dissolved in water (Figure 2.1.2b) is also a physical combination of two or more pure substances and is, therefore, a mixture. In this case, however, there are no distinct 'chunks' of matter that are visibly different from the rest of the coloured solution. The homogeneous nature of a solution of food colouring gives us the hint that it is a homogeneous mixture. Homogeneous mixtures are also known as **solutions**.
- The salt crystal (Figure 2.1.2c) is a pure substance and is not a physical combination of different substances. It is the compound sodium chloride (NaCl) and consists of elements chemically combined in a fixed ratio (i.e. sodium and chlorine in a 1:1 ratio).
- The sample of copper wire (Figure 2.1.2d) is also a pure substance and is an example of an element.

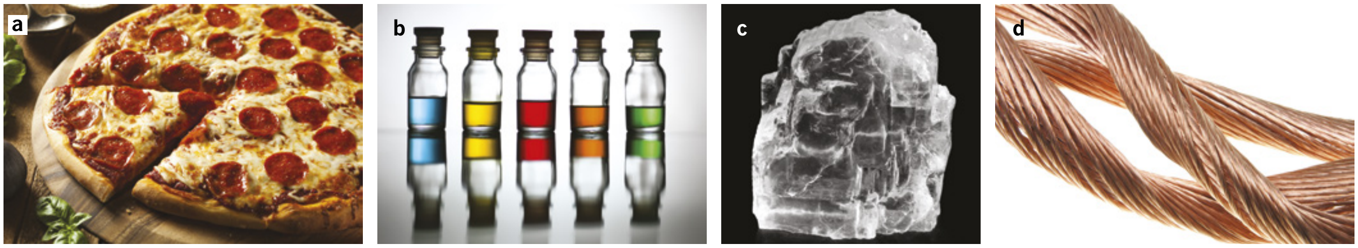
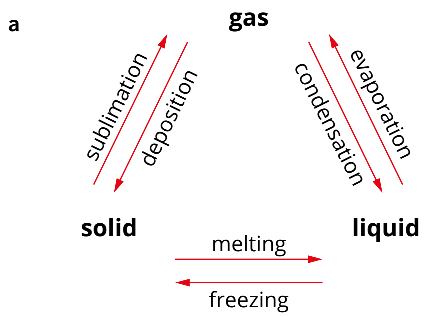


FIGURE 2.1.2 Examples of different types of matter: (a) a slice of pizza (heterogeneous mixture), (b) food colouring dissolved in water (homogeneous mixture and also known as a solution), (c) a salt crystal, which is a pure substance composed of the compound sodium chloride (NaCl) and (d) copper metal, which is an example of an element

PHYSICAL AND CHEMICAL CHANGES IN MATTER

In chemistry, you need to understand how matter can change from one form to another. A change in the form of matter can occur via physical changes and/or chemical changes. Figure 2.1.1 shows that combining pure substances to create mixtures requires a physical change. Separating mixtures into their pure components also requires a physical change. Figure 2.1.1 further shows that to create compounds or to decompose them into their elemental components requires a chemical change or **chemical reaction**.



b $$H_2O(s) \underset{-\Delta}{\overset{+\Delta}{\rightleftharpoons}} H_2O(1) \underset{-\Delta}{\overset{+\Delta}{\rightleftharpoons}} H_2O(g)$$

FIGURE 2.1.3 (a) Processes describing the changes of state between solids, liquids and gases. (b) The melting, boiling, condensing and freezing of water can be represented using chemical formulas and chemical equations. The bracketed letters, (s), (l) and (g), represent the solid, liquid and gaseous states of water. The '+Δ' and '−Δ' symbols refer to the need to 'add heat' or 'subtract heat' to induce the changes of state indicated.

Physical changes in matter

A **physical change** in matter is a process where the form of matter may be changed without changing its chemical identity or its chemical composition. No new substances are formed during physical changes. Cutting a piece of paper, grinding a tablet and bending an iron bar are examples of physical changes.

A change in **physical state** is one of the most important types of physical changes. The melting of ice to produce liquid water or the heating of water to produce gaseous steam are examples of physical changes of state. No new substances are formed in changes of state. The processes involved in changes of state are summarised in Figure 2.1.3a and represented using chemical equations in Figure 2.1.3b.

Chemical changes in matter

Chemical changes of matter (i.e. chemical reactions) involve a change in chemical composition where one or more kinds of matter are transformed into a new kind of matter (or several new kinds of matter). In other words, chemical reactions involve the production of new substances. You can see the results of chemical reactions around you every day. The burning of wood, the spoiling of milk, the digestion of food and the growth of plants via photosynthesis are all examples of chemical reactions.

In simple terms, a chemical reaction can be described as a rearrangement of **atoms**. The combustion of hydrogen (H_2) in the presence of oxygen (O_2) to produce water (H_2O) is one of the simplest chemical reactions. It can be represented using the balanced **chemical equation** below:

$$2H_2(g) + O_2(g) \xrightarrow{\text{heat}} 2H_2O(l)$$

Here two elements, hydrogen (H_2) and oxygen (O_2), chemically combine to produce the compound water (H_2O). The equation is balanced so the four hydrogen atoms and two oxygen atoms on the left-hand side of the equation are rearranged and incorporated into the two water molecules on the right-hand side of the equation.

MIXTURES

As Figure 2.1.1 on page 5 suggests, a mixture is a physical combination of two or more pure substances. This means there can be mixtures of:

- two or more elements (such as mercury–gold **amalgam**)
- mixtures of two or more compounds (such as salt water)
- mixtures of elements and compounds (such as oxygen dissolved in water).

Figure 2.1.1 also suggests that mixtures such as these can be physically separated into their pure components by simple physical processes. Processes such as cutting, crushing, sieving, filtration, distillation or centrifugation can produce pure substances from complex mixtures. The ability to separate mixtures into their pure components is crucial in many industrial, environmental and biomedical applications. Different separation methods will be discussed in detail in Module 2.3.

Mixtures can vary in composition from sample to sample with different types and amounts of substances being present. Since the composition of mixtures can vary, it follows that the chemical and physical properties of mixtures can also vary depending on the type and amount of substances present.

Figure 2.1.4 shows how the **boiling point** and **freezing point** of water change with small additions of sodium chloride (NaCl). The boiling point of pure water is 100°C and the freezing point of pure water is 0°C. Both change when other substances are mixed with water. The increase in boiling point is known as **boiling point elevation**. The more salt dissolved, the greater the change in boiling point. Likewise, the decrease in freezing point is known as **freezing point depression** and, again, the more salt dissolved, the greater the change in freezing point.

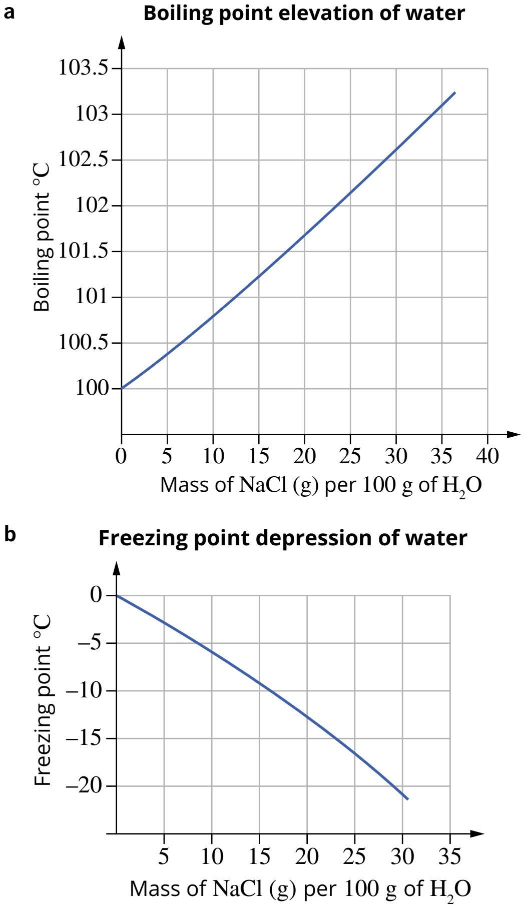


FIGURE 2.1.4 (a) Boiling point elevation—the boiling point of water increases with increasing amounts of NaCl. (b) Freezing point depression—the freezing point of water is lowered with increasing amounts of NaCl.

Changes in physical and chemical properties, like those shown in Figure 2.1.4, are useful for distinguishing between mixtures and pure substances. Mixtures will have different physical and chemical properties depending on the type and amount of substances present. Pure substances, on the other hand, do not vary in composition and therefore do not vary in chemical or physical properties. Some of the chemical and physical properties that can be used when characterising different types of matter are shown in Table 2.1.1.

TABLE 2.1.1 Examples of chemical and physical properties that can be used to characterise different types of matter. These properties can be used to distinguish mixtures from pure substances.

Chemical properties	Physical properties
combustibility/flammability	freezing point
reactivity in water	melting point
reactivity with acids	colour
reactivity with bases	viscosity
oxidisability	density
pH (specifically changes in pH)	solubility
toxicity	electrical conductivity
radioactivity	thermal conductivity
decomposition with heat	malleability/ductility

WS
1.2.2

PURE SUBSTANCES

A pure substance (or simply a substance) is matter that has a definite and distinct set of physical and chemical properties that do not vary in composition from sample to sample. In general, any two samples of matter that have identical chemical and physical properties are said to be the same substance. Therefore, chemical and physical properties (such as those outlined in Table 2.1.1 on page 7) can be used to identify a particular sample of matter. For example, a shiny, silver-coloured metal that has a melting point of 660.3°C, a **density** of 2.70 $g\,cm^{-3}$ and reacts with acid to produce hydrogen gas (H_2) can only be the element aluminium (Al). This is because only aluminium has this definite and distinct set of chemical and physical properties.

There are two types of substances: elements and compounds. As Figure 2.1.1 on page 5 shows, elements combine by chemical reactions to form compounds, while compounds can be decomposed into elements by chemical reactions. Unlike mixtures, substances cannot be separated into other kinds of matter by simple physical processes such as filtration, distillation and centrifugation.

ELEMENTS

Monatomic elements are those made up of only one atom. Diatomic elements are comprised of two atoms. The prefixes *mon* (or *mono*) and *di* are frequently used in chemistry. They mean 'one' and 'two' respectively.

A molecule is a definite and discrete group of atoms chemically bonded together. The atoms in molecules are non-metallic atoms bonded to other non-metallic atoms.

Elements are the simplest form of matter that exists. They cannot be broken down into other substances by simple physical processes, nor can they be broken down into other substances by chemical reactions. Elements are the building blocks of matter since they can combine chemically to form millions of different compounds. The defining feature of elements is that they are substances that contain only one type of atom. The monatomic gases helium (He), neon (Ne) and argon (Ar) are examples of elements; the diatomic **molecules** oxygen (O_2), nitrogen (N_2), hydrogen (H_2) and bromine (Br_2) are also examples of elements; so too are the metals sodium (Na), copper (Cu), aluminium (Al) and iron (Fe).

Most non-metallic elements form molecules with a definite number of atoms. Sulfur, for example, is composed of molecules with eight sulfur atoms (S_8). However, some non-metals form **covalent network lattices** or **giant molecules**. Carbon is an example of such a non-metallic element. Diamond and graphite are both examples of covalent network lattices formed by carbon. Graphene is a giant molecule formed by carbon. (You will learn more about covalent network lattices formed from carbon in Chapter 8.) Representations of the sulfur molecule and a carbon covalent network lattice are shown in Figure 2.1.5. Metallic elements form a different type of network lattice structure, which you will look at in detail in Chapter 6.

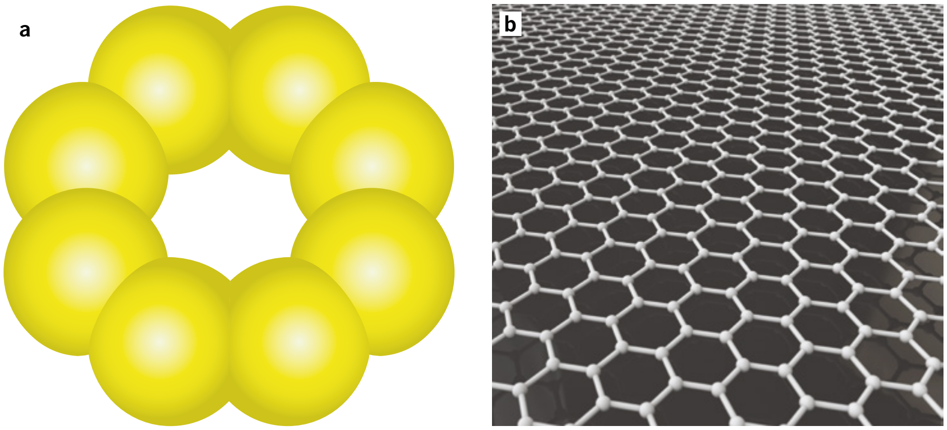


FIGURE 2.1.5 (a) Most non-metal elements, such as sulfur, form molecules. (b) Other elements, such as carbon, form covalent network lattices or giant molecules, given by the example here of graphene.

Element names, symbols and numbers

At present there are 118 known elements, 92 of them naturally occurring, while the other 26 have been synthesised in laboratories and are very unstable. Each element is assigned a unique name and **chemical symbol**. Chemical symbols are typically either a single capital letter (e.g. H for hydrogen) or a single capital letter followed by a lower-case letter (e.g. Ne for neon). Most chemical symbols make sense from their names (e.g. C for carbon or Mg for magnesium). Others symbols make less sense as their symbol may be derived from their Latin name (e.g. Au for gold, from the Latin *aurum*, or K for potassium, from the Latin *kalium*). Table 2.1.2 shows an alphabetical listing of some common elements along with their chemical symbols and some observable physical properties.

TABLE 2.1.2 Alphabetical listing of common elements including names (Latin name in brackets), symbols and physical properties

Element	Chemical symbol	Physical properties
aluminium	Al	silvery metal
barium	Ba	silvery metal
bromine	Br	reddish liquid
calcium	Ca	silvery metal
carbon	C	soft, black solid (graphite)
chlorine	Cl	greenish gas
chromium	Cr	silvery metal
cobalt	Co	silvery metal
copper (cuprum)	Cu	reddish metal
fluorine	F	pale yellow gas
gold (aurum)	Au	soft, yellow metal
helium	He	colourless gas
hydrogen	H	colourless gas
iodine	I	bluish-black solid
iron (ferrum)	Fe	silvery metal
lead (plumbum)	Pb	bluish metal
magnesium	Mg	silvery metal
manganese	Mn	grey metal
mercury (hydrargyrum)	Hg	silvery liquid
neon	Ne	colourless gas
nickel	Ni	silvery metal
nitrogen	N	colourless gas
oxygen	O	colourless gas
phosphorus	P	yellowish solid (white phosphorus)
potassium (kalium)	K	soft, silvery metal
silver (argentum)	Ag	silvery metal
sodium (natrium)	Na	soft, silvery metal
sulfur	S	yellow solid
zinc	Zn	bluish-white metal

Along with a name and chemical symbol, each element is also assigned a number, called the **atomic number**. The atomic number identifies the number of protons in the atom. For our purposes, atomic numbers range from 1 (for hydrogen) up to 92 (for uranium), i.e. the 92 naturally occurring elements. At this stage you should become familiar with the first 20 elements (Table 2.1.3 on page 10).

TABLE 2.1.3 The first 20 elements listed in order of increasing atomic number

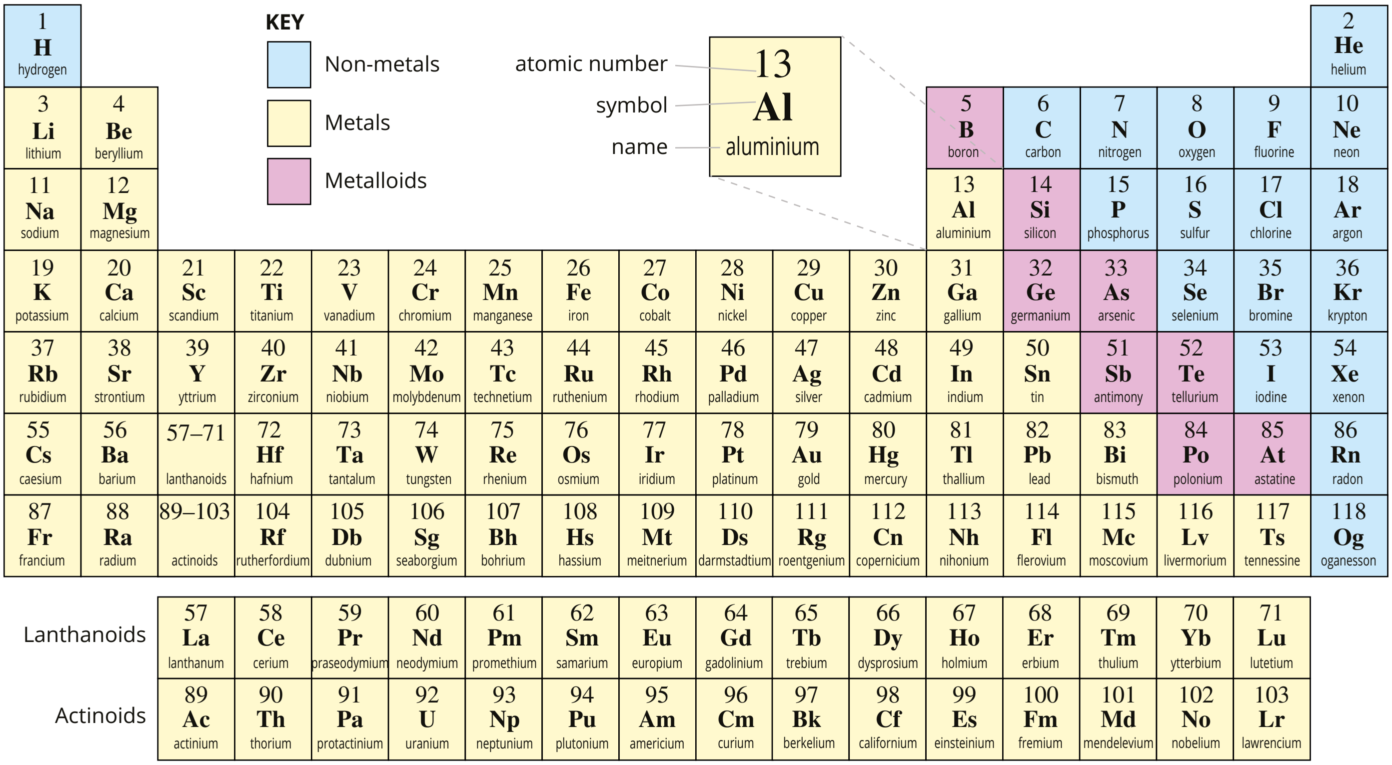
Atomic number	Name	Symbol	Atomic number	Name	Symbol
1	hydrogen	H	11	sodium	Na
2	helium	He	12	magnesium	Mg
3	lithium	Li	13	aluminium	Al
4	beryllium	Be	14	silicon	Si
5	boron	B	15	phosphorus	P
6	carbon	C	16	sulfur	S
7	nitrogen	N	17	chlorine	Cl
8	oxygen	O	18	argon	Ar
9	fluorine	F	19	potassium	K
10	neon	Ne	20	calcium	Ca

The periodic table

Figure 2.1.6 shows that elements can be listed in a special way in the **periodic table of elements**. The periodic table groups elements with similar chemical and physical properties into vertical columns called **groups**. Most elements are metals, which appear on the left-hand side of the periodic table, while the non-metals appear towards the upper-right of the periodic table.

You will generally have a copy of the periodic table at hand during your chemistry studies so detailed memorisation is not normally required. However, being able to recall specific information about the first twenty elements or so will be very useful. It is very important to learn how to use the periodic table since it is the most useful tool in chemistry. There are many useful trends in the periodic table that you will learn more about in Chapter 4.

ℹ The periodic table is an arrangement of the elements in order of increasing atomic number in which elements of similar chemical and physical properties are placed in vertical columns known as groups.



1 H hydrogen																	2 He helium
3 Li lithium	4 Be beryllium											5 B boron	6 C carbon	7 N nitrogen	8 O oxygen	9 F fluorine	10 Ne neon
11 Na sodium	12 Mg magnesium											13 Al aluminium	14 Si silicon	15 P phosphorus	16 S sulfur	17 Cl chlorine	18 Ar argon
19 K potassium	20 Ca calcium	21 Sc scandium	22 Ti titanium	23 V vanadium	24 Cr chromium	25 Mn manganese	26 Fe iron	27 Co cobalt	28 Ni nickel	29 Cu copper	30 Zn zinc	31 Ga gallium	32 Ge germanium	33 As arsenic	34 Se selenium	35 Br bromine	36 Kr krypton
37 Rb rubidium	38 Sr strontium	39 Y yttrium	40 Zr zirconium	41 Nb niobium	42 Mo molybdenum	43 Tc technetium	44 Ru ruthenium	45 Rh rhodium	46 Pd palladium	47 Ag silver	48 Cd cadmium	49 In indium	50 Sn tin	51 Sb antimony	52 Te tellurium	53 I iodine	54 Xe xenon
55 Cs caesium	56 Ba barium	57–71 lanthanoids	72 Hf hafnium	73 Ta tantalum	74 W tungsten	75 Re rhenium	76 Os osmium	77 Ir iridium	78 Pt platinum	79 Au gold	80 Hg mercury	81 Tl thallium	82 Pb lead	83 Bi bismuth	84 Po polonium	85 At astatine	86 Rn radon
87 Fr francium	88 Ra radium	89–103 actinoids	104 Rf rutherfordium	105 Db dubnium	106 Sg seaborgium	107 Bh bohrium	108 Hs hassium	109 Mt meitnerium	110 Ds darmstadtium	111 Rg roentgenium	112 Cn copernicium	113 Nh nihonium	114 Fl flerovium	115 Mc moscovium	116 Lv livermorium	117 Ts tennessine	118 Og oganesson

Lanthanoids	57 La lanthanum	58 Ce cerium	59 Pr praseodymium	60 Nd neodymium	61 Pm promethium	62 Sm samarium	63 Eu europium	64 Gd gadolinium	65 Tb trebium	66 Dy dysprosium	67 Ho holmium	68 Er erbium	69 Tm thulium	70 Yb ytterbium	71 Lu lutetium
Actinoids	89 Ac actinium	90 Th thorium	91 Pa protactinium	92 U uranium	93 Np neptunium	94 Pu plutonium	95 Am americium	96 Cm curium	97 Bk berkelium	98 Cf californium	99 Es einsteinium	100 Fm fremium	101 Md mendelevium	102 No nobelium	103 Lr lawrencium

FIGURE 2.1.6 The periodic table groups elements according to their chemical and physical properties.

COMPOUNDS

Compounds are substances formed from two or more elements in which the elements are always combined in the same fixed proportion. This means the composition of compounds does not vary, no matter how much of the compound there is. Water is a compound in which hydrogen and oxygen are always combined in the ratio of 2:1 and is represented by the **chemical formula** H_2O.

A chemical formula is a shorthand notation that uses elemental symbols from the periodic table, with numerical subscripts to convey the relative proportions of atoms of the different elements in the compound. You will note that the oxygen atom in the formula H_2O has no subscript. When an element in a chemical formula has no subscript, the subscript is presumed to be the number one.

Compounds are substances that can be broken down by chemical reactions to form other substances. To determine whether a pure substance is an element or a compound, you must determine if the substance can be broken down into elements. For example, when heated, mercury(II) oxide (HgO) decomposes to liquid mercury (Hg) and oxygen gas (O_2) (Figure 2.1.7). If it were not a compound, the mercury(II) oxide would not break down. As oxygen is a colourless gas, you cannot see it.

FIGURE 2.1.7 The red powder in this test-tube is mercury(II) oxide (HgO). If you look closely at the test-tube, you will see beads of liquid mercury forming from the decomposition of the compound.

Types of compounds

There are two major types of compounds: **molecular compounds** and **ionic compounds**. Molecular compounds are composed of molecules all of which are alike and have non-metallic elements chemically bonded to other non-metallic elements in a fixed ratio. They tend to have relatively low boiling points and melting points. Examples of common molecular compounds include water (H_2O), methane (CH_4), ammonia (NH_3), benzene (C_6H_6), ethanol (C_2H_6O) and carbon dioxide (CO_2). Note how each example contains only non-metallic elements.

Ionic compounds form when metallic elements bond to non-metallic elements. Ionic compounds are composed of **ions** arranged in a rigid three-dimensional lattice. They contain positively charged ions (called **cations**) and negatively charged ions (called **anions**), which are attracted to each other by the electrostatic attraction of charges of opposite sign. They tend to have relatively high melting points and boiling points compared to molecular compounds. Examples of common ionic compounds include sodium chloride (table salt, NaCl), calcium carbonate (limestone, $CaCO_3$) and calcium oxide (lime, CaO). Note how each example contains a metallic cation and a non-metallic anion.

You will look at ionic compounds and molecular compounds in more detail in Chapters 7 and 8.

2.1 Review

SUMMARY

- Matter can be characterised and classified according to its purity.
- Pure substances are materials with definite and distinct chemical and physical properties.
- Mixtures are physical combinations of pure substances whose properties are dependent on the identity and relative amounts of the substances that make up the mixture.
- Changes in matter are brought about by physical changes or chemical changes: physical changes do not produce new substances; chemical changes result in the formation of new substances.
- Pure substances may be elements or compounds.
- Every element has a unique name, atomic number and chemical symbol.
- Elements are organised into the periodic table.
- Compounds are formed from two or more elements combined in the same fixed proportion.
- Molecular compounds are composed of non-metals bonded to other non-metals.
- Ionic compounds are composed of metals bonded to non-metals.

KEY QUESTIONS

Retrieval

1 Define the term 'matter'.

2 Name two types of:
  a mixtures
  b pure substances.

3 Describe a physical change in matter.

4 Define a pure substance.

5 Identify the common name of each of the following elements from its Latin name.
  a ferrum
  b kalium
  c argentum
  d plumbum
  e hydrargyrum

6 Define the term 'compound' and list the two major classes of compounds.

7 Select the correct terms to complete the following sentence.
Molecular compounds/ionic compounds are composed of non-metals bonded to non-metals, whereas *molecular compounds/ionic compounds* are composed of metals bonded to non-metals.

8 Name two physical properties that could be used to distinguish between these substances.
  a water and methanol
  b gold and copper
  c oxygen gas and chlorine gas

Comprehension

9 Describe the change of state associated with each of the following processes.
  a Water is made into ice cubes.
  b The inside of your car window fogs up.
  c Mothballs in the wardrobe disappear with time.
  d Wet washing dries.

10 A certain substance is a silver-grey coloured metal that melts at 420°C. When it is placed in dilute sulfuric acid, hydrogen is given off and the metal dissolves. It has a density of 7.13 g cm^{-3} at 25°C and reacts slowly with oxygen to form a metal oxide. Describe the physical and chemical properties of the substance referred to above.

11 Determine if the following is a physical or chemical change.
  a A sample of mercury(II) oxide was heated in a reaction vessel to produce mercury metal and oxygen gas.
  b A glowing wood splint was thrust into the reaction vessel and the splint burst into flame.

12 Explain the differences between an element, a compound and a mixture.

Analysis

13 The following are properties of a certain element. Classify them as physical or chemical.
  a In powdered form, it burns brilliantly on ignition.
  b Bulk metal does not react with steam even when red hot.
  c It has a density of 1.85 g cm^{-3} at 20°C.
  d It is a relatively soft, silvery-white metal.

14 Classify each of the following as a physical change or chemical change.

- **a** the evaporation of water
- **b** the rusting of iron
- **c** the grinding of salt crystals into powder
- **d** the burning of wood in a fireplace

15 Classify each of the following as an element, a compound or a mixture.

- **a** copper
- **b** sand
- **c** water
- **d** carbon dioxide
- **e** muddy water
- **f** sodium chloride
- **g** gold
- **h** lemonade

16 About 3.5% (3.5 g per 100 g) of the mass of sea water is the result of dissolved salts, mainly sodium chloride. Determine the freezing point of sea water using the graph in Figure 2.1.4b on page 7.

17 Classify each of the following elements on the periodic table on page 10 as a metal, metalloid or non-metal and represent each element using its chemical symbol.

- **a** magnesium
- **b** manganese
- **c** silver
- **d** mercury
- **e** neon
- **f** arsenic
- **g** sulfur
- **h** silicon

18 Classify the following as ionic compounds or molecular compounds using the periodic table on page 10.

- **a** $NaCl$
- **b** H_2S
- **c** PF_3
- **d** Fe_2O_3

19 Identify an element that has similar physical and chemical properties to potassium, K. Explain your reasoning.

2.2 Homogeneous and heterogeneous mixtures

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- distinguish between homogeneous mixtures and heterogeneous mixtures
- understand that the defining feature of a heterogeneous mixture is the presence of visually distinguishable phases that have different physical and chemical properties
- understand that liquid homogeneous mixtures, also known as solutions, are composed of solutes dissolved in a solvent.

You have already noted that most samples of matter are not chemically pure and consist of a physical combination of two or more pure substances called a mixture. You have also noted that there are two types of mixtures—homogeneous mixtures and heterogeneous mixtures.

The terms **homogeneous** and **heterogeneous** have Greek origins: *homo*, meaning 'same', *hetero*, meaning 'different', and *genes*, meaning 'of a kind'. Homogeneous therefore translates to 'of the same kind' and heterogeneous translates to 'of a different kind'.

FIGURE 2.2.1 Some examples of mixtures: (a) This sample of granite shows at least three visibly distinct regions—white quartz, orange feldspar and black mica minerals. (b) A choc-chip cookie has at least two visibly distinct regions. (c) Some salad dressings are made from oil and water.

In some instances, mixtures are easily recognised. For example, consider a piece of granite, a choc-chip cookie and salad dressing (Figure 2.2.1). In these examples, you can see that different kinds of substances are present. In other cases, it is not so easy to recognise mixtures. For example, the air you breathe, sea water and sterling silver jewellery (Figure 2.2.2) may all appear to be pure but each consists of different substances. Air is a mixture of elements such as nitrogen (N_2) and oxygen (O_2) combined with compounds such as carbon dioxide (CO_2) and water vapour (H_2O); sea water is mostly a mixture of the compounds water (H_2O) and sodium chloride ($NaCl$); while sterling silver is a mixture of the elements silver (Ag) and copper (Cu). It is the uniformity of these mixtures and the lack of visibly different materials that makes it hard for us to recognise them as mixtures.

FIGURE 2.2.2 Examples of matter not easily recognised as mixtures. (a) Air is a colourless mixture of nitrogen, oxygen and some trace gases. (b) Sea water is a colourless mixture of salt and water. (c) Sterling silver is a mixture of silver and copper but appears to be a single lustrous silver-coloured metal.

HETEROGENEOUS MIXTURES

The piece of granite, the choc-chip cookie and the salad dressing shown in Figure 2.2.1 are examples of heterogeneous mixtures. These samples of matter are not uniform throughout and you can clearly observe the presence of different types of materials. You also know from experience that the different parts of each of these mixtures have different properties, such as colour, taste and hardness.

Heterogeneous mixtures consist of two or more substances that have visibly distinguishable regions, called **phases**, which have different physical and chemical properties. A heterogeneous mixture is not uniform throughout, so two small samples obtained from different parts of the mixture would be different in composition.

Heterogeneous mixtures may have phases in the same physical state or in different physical states. Granite is a heterogeneous mixture of three solid phases—the white quartz mineral (i.e. silica, SiO_2), the orange feldspar mineral and the black mica mineral. The oil and water phases of salad dressing are also both in the same physical state—the liquid state. On the other hand, a sample of muddy water consists of solid dirt particles physically mixed with liquid water.

You will see in Module 2.3 that the different phases in a heterogeneous mixture can be readily separated using simple mechanical separation techniques.

A phase is a region of matter that is physically and chemically uniform in composition and properties. It is physically distinct from other regions of matter and is mechanically separable from other phases.

HOMOGENEOUS MIXTURES

Homogeneous mixtures consist of a physical combination of two or more substances but have only one visibly distinct phase which has uniform properties. A homogeneous mixture is uniform throughout and samples taken from different parts of the mixture would be identical in composition. The air, sea water and sterling silver shown in Figure 2.2.2 are all examples of homogeneous mixtures where only one visibly distinct phase is observable.

Many homogeneous mixtures are also called solutions and have one substance dissolved in another. The substance present in the greatest amount is called the **solvent** and all other substances present in the mixture are called **solutes**. Solutes are said to be dissolved in the solvent.

The most common solutions you will encounter in your chemistry studies will be solid salts dissolved in liquid water (for example, sea water). The salts are the solutes and the water is the solvent. A solution in which water is the solvent is given the special name of an **aqueous solution**—the name being derived from the Latin *aqua*, meaning 'water'. Examples of some common solutions are shown in Table 2.2.1, which shows that solutions can involve mixtures across all three states of matter.

TABLE 2.2.1 Examples of common solutions

Example	States of matter involved	Solvent	Solute(s)	Physical appearance
air	gas–gas	nitrogen	oxygen, carbon dioxide, argon, water vapour	clear colourless gas
soft drinks	liquid–gas	water	carbon dioxide gas	coloured liquid
vinegar	liquid–liquid	water	ethanoic acid	clear colourless liquid (white vinegar)
sea water	liquid–solid	water	sodium chloride plus other trace salts	clear colourless liquid
sterling silver	solid–solid	silver	copper	lustrous silver-coloured solid metal

Even though any single sample of a homogeneous mixture will be uniform throughout, the composition may vary from sample to sample, depending on the relative ratio of the substances in the solution. For example, two samples of salt water may be prepared by dissolving, firstly, one gram of salt in a litre of water and then, secondly, 10 grams of salt in a litre of water. Both salt water samples will be homogeneous throughout but each sample will have different physical and chemical properties including density, electrical conductivity and boiling point.

You will see in Module 2.3 that the different components of a heterogeneous mixture are often separated using techniques that involve a change of state.

2.2 Review

SUMMARY

- Most matter is not chemically pure and consists of a physical combination of two or more pure substances that form a mixture.
- Mixtures can be classified as heterogeneous mixtures or homogeneous mixtures.
- Heterogeneous mixtures have two or more visibly distinguishable regions, called phases, which have different physical and chemical properties. They are not uniform throughout.
- Homogeneous mixtures may be known as solutions and have only one visibly distinct phase. They are uniform throughout.
- Many homogeneous mixtures are known as solutions.
- Solutions consist of solutes dissolved in a solvent. When the solvent is water, the solution is known as an aqueous solution.
- Both heterogeneous and homogeneous mixtures can vary in composition from sample to sample and therefore can vary in chemical and physical properties.

KEY QUESTIONS

Retrieval

1 **a** Define the term ‘phase’ with respect to heterogeneous and homogeneous mixtures.
  b State how many phases all solutions have.

2 Define the terms ‘solution’, ‘solvent’ and ‘solute’.

3 Indicate all possible answers that apply to each scenario from the list below.

compound
element
heterogeneous mixture
homogeneous mixture

  a a sample of matter that is not uniform throughout
  b a sample of matter that is uniform throughout
  c matter that can vary in composition

4 Name the two different types of mixtures that involve one substance being dissolved in another substance.

Comprehension

5 Explain the difference between a heterogeneous mixture and a homogeneous mixture using suitable examples.

Analysis

6 Classify each of the following as a homogeneous or heterogeneous mixture.
  a muesli
  b sand
  c wet sand
  d vinegar
  e an apple
  f sea water
  g a tree
  h air

7 Differentiate between a *state* of matter and a *phase* of matter.

8 Identify which of the following can have a varied composition. Explain your answer in each case.
  a element
  b compound
  c homogeneous mixture
  d heterogeneous mixture

9 Compare and contrast the variation in composition between a heterogeneous mixture and a homogeneous mixture.

10 Classify each of the following as an element, a compound, a heterogeneous mixture or a homogeneous mixture. For each, write down the different phases present.
  a printing ink solution with tiny particles of carbon black
  b iodine crystals and their vapour
  c a salt solution with salt crystals at the bottom of the flask
  d a blue solution of copper(II) sulfate
  e beach sand

11 Differentiate the following samples of matter containing sodium and/or chlorine by classifying them as elements, compounds, heterogeneous mixtures or homogeneous mixtures.
  a sodium metal
  b chlorine gas
  c sodium chloride crystals
  d sodium chloride dissolved in water
  e table salt (i.e. sodium chloride with added sodium iodide)

12 Identify the following metallic substances as elements or mixtures.
  a brass
  b sterling silver
  c bronze
  d mercury
  e gold
  f iron

2.3 Separating mixtures

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recall separation techniques used to separate both heterogeneous and homogeneous mixtures
- understand that separation techniques use differences in the physical properties of the components to separate them from each other
- understand that separating phases of a heterogeneous mixture involves mechanical separation techniques
- understand that separating components of a homogeneous mixture typically involves a change of state.

Removing impurities from samples of matter or separating mixtures into pure components are crucial processes in many biomedical, environmental and industrial applications. For example, ensuring the purity of pharmaceutical drugs or removing impurities in drinking water supplies have significant health-related consequences.

If you want to separate mixtures, you can generally use the differences in the physical properties of the components of the mixture to separate the components from each other. Since mixtures are physical combinations of substances, relatively simple physical processes can be used to separate them.

SEPARATING HETEROGENEOUS MIXTURES

Heterogeneous mixtures have different phases that are physically distinct and mechanically separable from each other. Depending on the nature of the mixture, mechanical separation can take several forms.

Hand sorting

Hand sorting is perhaps the simplest separation technique and can be used when there are relatively few objects to sort that have differing physical properties such as size, colour and texture. Separating seashells from sand would be a relatively straightforward process that could be done by hand. Similarly, if you needed to isolate the white quartz crystals from a sample of granite, you could use simple physical processes, such as cutting, grinding or crushing, followed by hand sorting—with the aid of a pair of tweezers if required.

FIGURE 2.3.1 Queensland beaches are cleaned by tractors dragging a rotating mesh drum to sieve large objects such as rocks, shells and other rubbish.

Sieving

Sieving can be used as a separation technique if there are many objects to sort and they are of different sizes. Beach-cleaning tractors are used daily on Queensland beaches to sift the sand to remove rocks, shells and rubbish, leaving behind clean sand (Figure 2.3.1). In a similar process, the primary treatment of sewage wastewater uses large rotating mesh screens to screen out (or sieve) large particles or objects, such as food scraps, that are part of the wastewater mixture.

FIGURE 2.3.2 Filtration of muddy water. Solid mud and clay particles suspended in water can be separated from the liquid water by the process of filtration.

Filtration

Fine grade separation of solids and liquids can be achieved in the laboratory using filter paper and a filter apparatus. In Figure 2.3.2, you can see the result of passing a muddy water sample through a filter apparatus. Muddy water consists of solid particles of dirt and clay suspended in water. During filtration, the solid particles are trapped by the filter paper while the clear liquid water easily passes through, resulting in the separation of the solid and liquid phases. This type of filtration technique is good for separating small samples of undissolved solids from a liquid.

PA 1.2.1

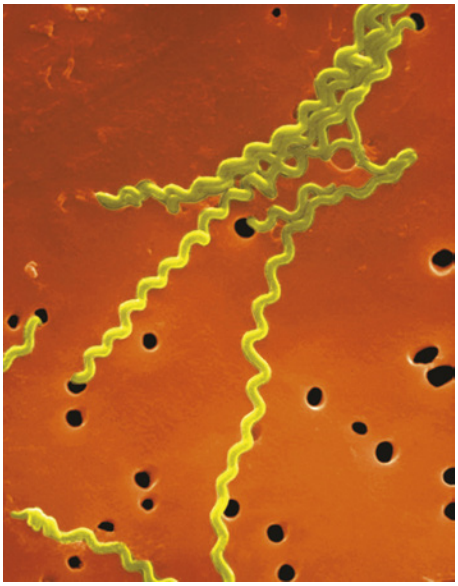


FIGURE 2.3.3 Polycarbonate filter membranes with well-defined pore size allow bacterial cells to be separated from water samples. The bacterial cells shown are the species *Leptospira interrogans*.

Different types of filter paper are used for different types of samples but most are manufactured from ashless paper, nitrocellulose or polycarbonate membranes. They can also be manufactured to separate particles of definite size. Figure 2.3.3 shows a close-up image of a polycarbonate filter membrane with pore sizes of the order of 0.2 μm. This filter membrane will block the passage of particles larger than this size and can effectively sterilise water by filtering bacterial cells from water samples.

Centrifugation

Another way to separate solid particles from liquids is to centrifuge them. **Centrifugation** is a separation technique that uses the centrifugal force of rotational motion to promote rapid settling of solid particles in a heterogeneous solid–liquid mixture. One of the most common uses of centrifugation involves the separation (or fractionation) of blood. Blood samples are placed into centrifuge tubes and rotated at very high speeds (Figure 2.3.4a). The solid components of the blood mixture are forced towards the bottom of the centrifuge tube. The end result is the separation of blood into different fractions with red blood cells settled at the bottom of the tube, white blood cells and platelets forming a layer above the red blood cells, and the liquid blood plasma sitting on top of the other layers (Figure 2.3.4b).

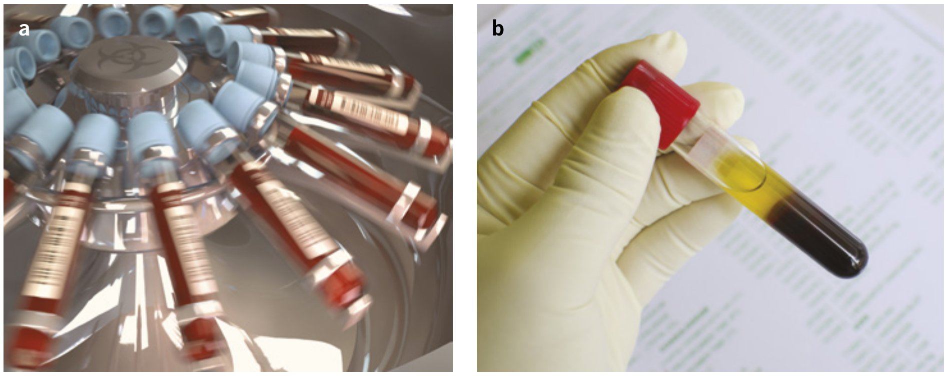


FIGURE 2.3.4 Centrifugation of whole blood is a common technique used in pathology laboratories. (a) Blood samples in centrifuge tubes are placed in a centrifuge and rotated at very high speeds. (b) Blood can be separated into different fractions: red blood cells, white blood cells, platelets and blood plasma.

Flotation

Flotation techniques take advantage of the differences in density of materials to separate heterogeneous mixtures. For example, a mixture of sawdust and sand can be easily separated by placing the mixture in water. The dense sand will sink to the bottom while the less dense sawdust will float on top, allowing it to be skimmed from the surface. This technique is used in the mining industry, in wastewater treatment plants and in paper recycling plants to separate complex mixtures.

Decantation

Decanting is another technique used for separating components of different densities. Decanting involves carefully pouring off the top liquid layer of a heterogeneous mixture. The mixture could be a liquid phase lying over a solid phase or it could be a liquid phase lying over another liquid phase. You can easily separate a mixture of sand and water by decanting the water, leaving behind the sand in the bottom of the container. Similarly, the liquid oil phase of salad dressing is easily decanted from the top of the more dense water layer underneath. Figure 2.3.5 shows that a separating funnel can also be used to separate two immiscible liquids, with the denser bottom layer being drained from the heterogeneous mixture.

FIGURE 2.3.5 A separating funnel is used to separate two immiscible liquids.

Magnetic separation

Magnetic separation can be used to separate heterogeneous mixtures where some components have magnetic properties. Figure 2.3.6 shows that a mixture of fine sand and iron filings can be readily separated using a magnet.

FIGURE 2.3.6 A magnet is used to separate iron filings from the non-magnetic sand particles.

SEPARATING HOMOGENEOUS MIXTURES

Homogeneous mixtures (or solutions) have one or more solutes dissolved in a solvent with only one visibly distinct phase that has uniform properties throughout. Separating solutions, therefore, requires more sophisticated techniques than the mechanical separation methods used for heterogeneous mixtures. One way to separate solutions is to employ changes of state to take advantage of the differences in boiling points or melting points of the components in the mixture.

Evaporation

As mentioned earlier, the most common examples of homogeneous solutions are solids dissolved in a liquid solvent. Evaporation of the solvent is the most convenient way of removing the liquid component and recovering the dissolved solid. Figure 2.3.7 shows that heating a salt solution will evaporate the water, leaving behind the solid salt. This process takes advantage of the differences in boiling points of the two substances involved.

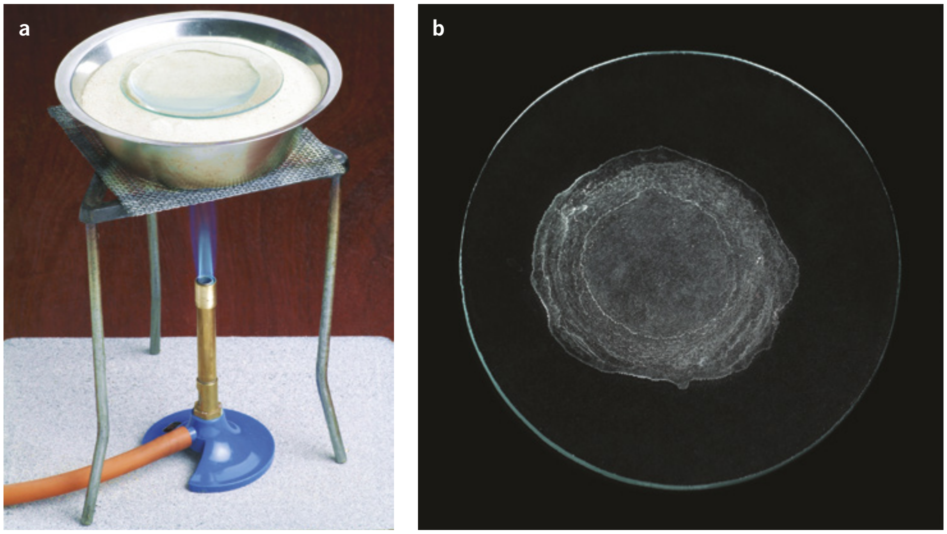


FIGURE 2.3.7 Evaporation of the solvent recovers the solid from a solid–liquid solution. (a) The salt solution is heated to drive off the water. (b) Solid salt residue remains after all of the water has evaporated.

Distillation

If both the solid solute and the liquid solvent need to be recovered then you need to use **distillation**. Distillation is a process of separating mixtures containing a liquid component by first evaporating the liquid to its gaseous state and then condensing it back to its liquid state. Figure 2.3.8 on page 20 shows a diagram of a typical distillation apparatus used in chemistry laboratories. This approach can be used to separate and recover both components of a salt water solution. A flask containing the salt solution is heated so liquid water evaporates. The water vapour is directed through a condensation tube, which is kept cool by a constant supply of cold water. The water vapour condenses to form pure liquid water, which is collected in the receiving flask. When all of the water has evaporated, a layer of pure salt will be retained on the inside surface of the distillation flask. In this way, both the pure salt and pure water are separated and recovered.

Volatility is a measure of how readily a substance will vaporise by going from its liquid state to its gaseous state. In general, substances with lower boiling points have higher volatility.

You can use the same distillation approach for separating a solution of two liquids by taking advantage of the difference in **volatility** and boiling points of the liquids. For example, a solution of ethanol and water could be separated via distillation by evaporating off the more volatile ethanol, leaving behind the higher boiling point water.

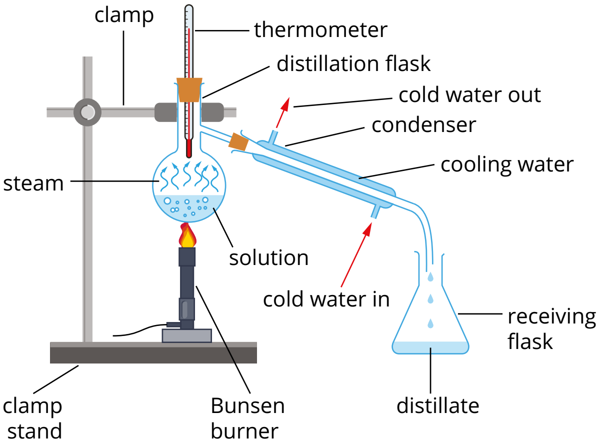


FIGURE 2.3.8 Typical distillation apparatus. A solution is vaporised by heating the distillation flask. Gaseous vapours are condensed in the condenser tube, which is kept cool by a constant supply of water. The pure liquid solvent is collected in the receiving flask and solutes remain in the distillation flask.

Fractional distillation of crude oil

Distillation is also used in one of the most important industrial processes of modern times—the production of petroleum products from crude oil. Crude oil is a complex mixture of different hydrocarbons with different boiling points. The differences in boiling points means that a process of **fractional distillation** can be used to separate and collect the different components (or fractions) of the mixture.

Fractional distillation of crude oil differs from ‘normal’ distillation in that a tall column, or tower, is situated above the liquid mixture with several condensers coming off at different heights. Figure 2.3.9a shows a typical distillation tower used in oil refineries, while Figure 2.3.9b shows a diagrammatic representation of the fractional distillation process used for crude oil. In this process, high temperature oil enters the distillation column at the bottom. As the mixture is vaporised it rises up the column and cools down with increasing height. Different components of the crude oil will condense at different temperatures, and therefore at different heights. Substances with high boiling points will condense at the high temperatures experienced at the bottom of the column; substances with low boiling points will condense at the lower temperatures experienced at the top of the column. Each of the different fractions is captured by condensers located at various heights.

The main fractions of crude oil that are collected include refinery gases (such as propane and butane), gasoline (i.e. petrol), naphtha, kerosene, diesel oil, fuel oil, and a residue containing paraffin wax, various oils and asphalt. Most of these fractions are used as fuels for heating or transport, while others are used as lubricants or in the manufacture of petroleum by-products such as plastics.

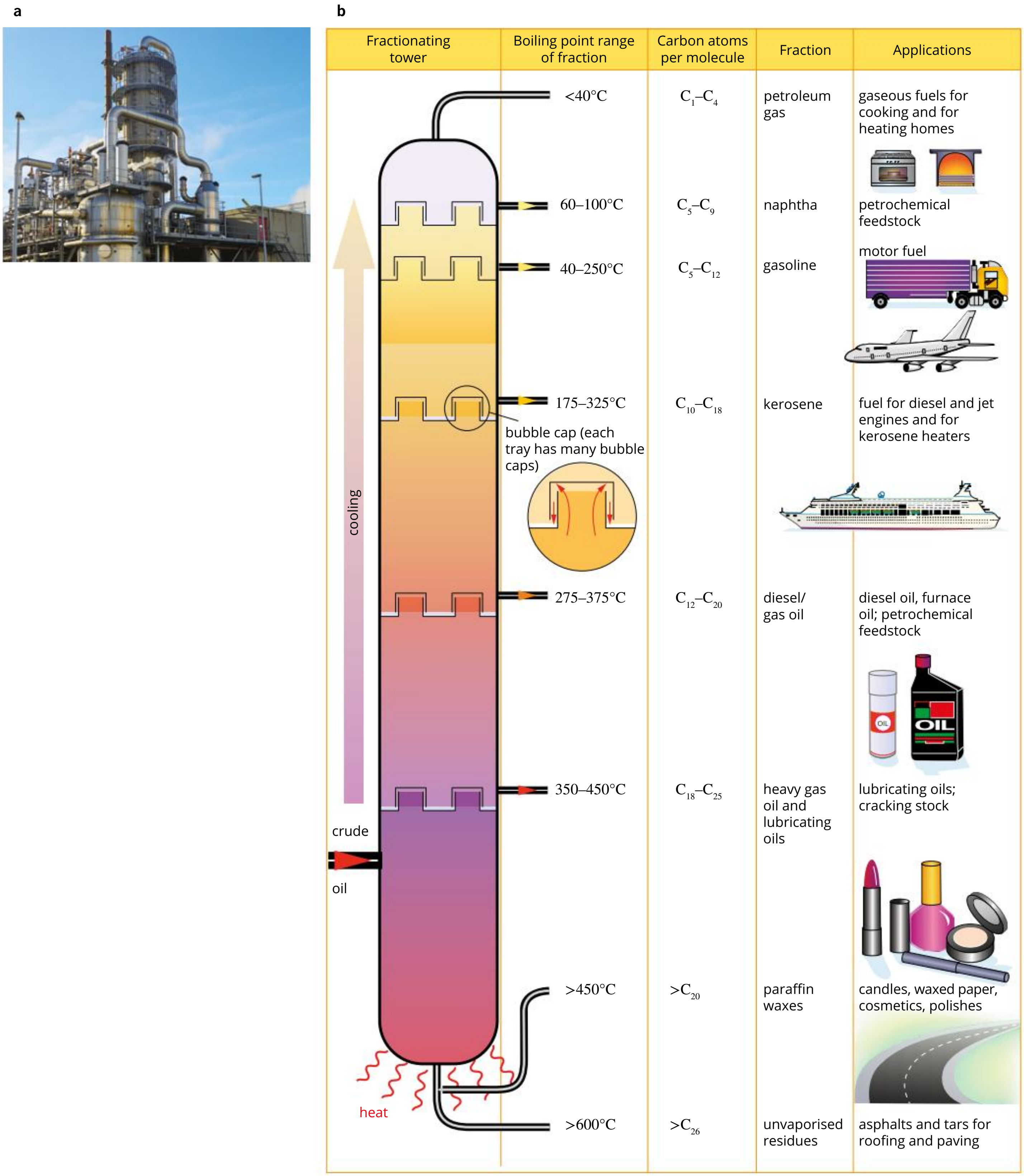


FIGURE 2.3.9 (a) A typical distillation tower used for the fractional distillation of crude oil in oil refineries. (b) A diagrammatic representation of the fractional distillation of crude oil. The crude oil is heated to around 400°C and piped into the bottom of the distillation tower. Different fractions are collected at different levels, depending on their boiling point.

2.3 Review

SUMMARY

- Separating mixtures uses relatively simple physical processes to take advantage of the differences in the physical properties of the components of the mixture.
- Hand sorting is a suitable separation method for mixtures containing a relatively small number of objects with visibly different properties.
- Sieving is a suitable separation method used for large numbers of different-sized particles.
- Filtration, centrifugation and decanting can all be used for separating undissolved solids from a liquid. Decanting can also be used for separating two immiscible liquids.
- Components of mixtures with differing magnetic properties can be separated using magnetic forces.
- Evaporation is a method for separating dissolved solids from the liquid solvent in a solid–liquid solution.
- Distillation is a suitable method for recovering both components of a liquid–solid solution.
- Distillation can also be used for separating two liquids with different boiling points.
- Fractional distillation is a method for separating complex mixtures of components with differing boiling points.

KEY QUESTIONS

Retrieval

1 Separating mixtures involves taking advantage of differences in the physical properties of the components that make up the mixture. Name one physical property that could be used to distinguish between the main components of these mixtures.

- **a** wine (main components are water and ethanol)
- **b** sterling silver (main components are silver and copper)
- **c** air (main components are oxygen gas and nitrogen gas)

2 Name the type of mixture that is separated into its constituent components by these processes.

- **a** sieving
- **b** filtration
- **c** flotation
- **d** distillation
- **e** evaporation

3 Recall the circumstances under which you would decant a mixture to separate its components.

4 Name the separation techniques that take advantage of differences in density.

Comprehension

5 Explain the physical properties you would take advantage of to separate the following mixtures. State the separation techniques you would employ.

- **a** iron filings and sand
- **b** salt and water
- **c** water and ethanol

6 Explain under what circumstances you would use distillation to separate an aqueous salt solution instead of simply evaporating the solvent.

7 Show your understanding of separation techniques by matching each scenario to the most appropriate technique.

Separation technique	Scenario
a sieving **b** filtration **c** evaporation **d** separating funnel **e** distillation **f** fractional distillation **g** centrifugation	**i** separation and recovery of each component in a complex aqueous solution of several different alcohols **ii** production of sea salt from salt water **iii** isolation of suspended solid particles from the Brisbane River water for laboratory analysis **iv** separation of the layers in an oil–water based salad dressing **v** separation of seashell fragments from sand **vi** separation and recovery of both components of a salt solution **vii** separation of the different fractions of whole blood

8 Explain the difference between distillation and fractional distillation when applied to the refining of crude oil.

Analysis

9 Identify the separation techniques that would be best used to separate and recover the following components within mixtures.

- **a** sand and gravel
- **b** boiled potatoes from the water they were cooked in
- **c** boiled rice from the water it was cooked in
- **d** silt particles from muddy water
- **e** hydrocarbon components in crude oil
- **f** salt and water from sea water

10 Compare and contrast the methods of evaporation and distillation for separating the components of a saltwater solution. Describe the advantages and disadvantages of each separation technique.

Chapter review

02

KEY TERMS

amalgam
anion
aqueous solution
atom
atomic number
boiling point
boiling point elevation
cation
centrifugation
chemical change
chemical equation
chemical formula
chemical reaction
chemical symbol
compound
covalent network lattice
decanting
density
distillation
element
fractional distillation
freezing point
freezing point depression
giant molecule
group
heterogeneous
heterogeneous mixture
homogeneous
homogeneous mixture
ion
ionic compound
mass
matter
mixture
molecular compound
molecule
periodic table of elements
phase
physical change
physical property
physical state
pure substance
sieving
solute
solution
solvent
volatility
volume

KEY QUESTIONS

Retrieval

1 Select the response that best describes a sample of matter that has these three characteristics.

- It is uniform throughout.
- It cannot be separated into other substances by physical processes.
- It can be decomposed into other substances by chemical processes.

A a heterogeneous mixture
B a homogeneous mixture
C an element
D a compound

2 Identify the following as either a chemical property or physical property of matter.

a boiling point
b melting point
c combustibility
d toxicity
e density

3 Select the correct terms to complete the following sentence.

A *physical change/chemical change* involves the formation of new substances, whereas a *physical change/chemical change* does not.

4 Name the vertical columns of the periodic table. State their significance.

Comprehension

5 Explain whether the composition of each of the following can vary. Explain your answer in each case.

a element
b compound
c homogeneous mixture
d heterogeneous mixture

6 Determine whether each sample of matter listed is a heterogeneous mixture, a homogeneous mixture or a pure substance.

a iron ore
b copper wire
c wet sand
d distilled water

7 Determine which of the following are pure substances and which are mixtures. For each, list all of the different phases present.

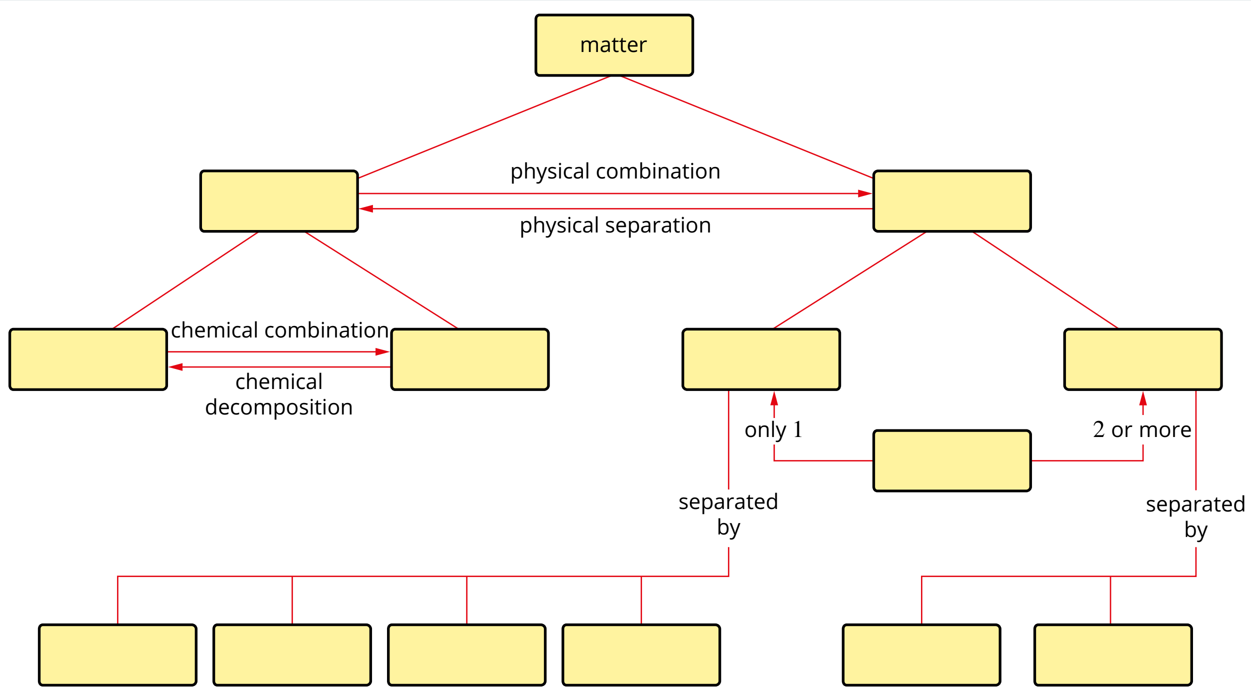
a alcohol and its vapour
b paint, containing a liquid solution and a dispersed solid pigment
c partially molten copper
d a sand containing quartz (silicon dioxide) and calcite (calcium carbonate)

8 The water in the Brisbane River is a mixture of water and suspended silt particles. A sample of Brisbane River water shows that the silt particles slowly settle to the bottom of a measuring cylinder under the action of gravity over a period of days. Describe two methods that could be used to rapidly separate the water and silt particles from the river water sample. Discuss the advantages and disadvantages of each method.

9 Describe the circumstances under which you would use distillation to separate an aqueous salt solution instead of simply evaporating the solvent.

10 Use the following list of terms and chart to complete a concept map that summarises the key ideas and their connections for this chapter.

centrifugation
compound
distillation
elements
evaporation
filtration
heterogeneous mixture
homogeneous mixture
mixture
phases
pure substance
sieving
sorting



Analysis

11 Classify each of the following pure substances as elements or compounds, based on the information given, or indicate that no such classification is possible because of insufficient information.

- **a** Analysis indicates that substance A contains two elements.
- **b** Substance B decomposes upon heating.
- **c** Heating substance C to 900°C causes no change.
- **d** Heating substance D to 400°C causes it to melt.

12 The following is a description of the element cadmium (Cd). Classify each descriptor as either a physical property or a chemical property.

- **a** It is a bluish-white coloured lustrous metal.
- **b** It has a melting point of 321°C.
- **c** When added to hydrochloric acid the metal dissolves and hydrogen gas is released.
- **d** It is highly toxic and can adversely affect the kidneys, lungs and bones.
- **e** It has a density of 8.65 g cm^{-3}.
- **f** It has a hardness of 2.0 on the Moh hardness scale.
- **g** If left in air it will form a layer of cadmium oxide (CdO) on its surface.

13 Classify each of the following changes as a physical change or chemical change.

- **a** the evaporation of ethanol
- **b** the rusting of steel
- **c** the grinding of sugar crystals into powder
- **d** the burning of coal in a fireplace

14 Classify each of the following changes as either a physical change or chemical change.

- **a** corrosion of zinc anodes on boats
- **b** the melting of iron in a blast furnace
- **c** the pulverising of a granite sample
- **d** digesting chocolate
- **e** the growth of plants via photosynthesis
- **f** explosion of TNT

15 Compare and contrast elements and compounds.

16 Identify the elements in the following molecular compounds, writing their name, symbol and atomic number.

- **a** water (H_2O)
- **b** ammonia (NH_3)
- **c** benzene (C_6H_6)
- **d** dinitrogen pentoxide (N_2O_5)
- **e** sulfur hexafluoride (SF_6)

17 Identify what you note about the nature of the elements in the compounds listed in Question **16** and what type of compound they are.

18 Identify the elements in the following ionic compounds, writing their name, symbol and atomic number.

- **a** sodium chloride (NaCl)
- **b** calcium fluoride (CaF_2)
- **c** aluminium oxide (Al_2O_3)
- **d** copper(I) sulfate (Cu_2SO_4)
- **e** iron(III) carbonate ($Fe_2(CO_3)_3$)

19 Identify what you note about the nature of the elements in the compounds listed in Question **18** and what type of compounds they are.

20 If you light a match under a cold metal spoon you may observe one or more of the following. Classify each observation as a physical change or a chemical change.

- **a** The match burns.
- **b** Carbon soot is produced.
- **c** The metal spoon gets warmer.
- **d** Water condenses on the metal spoon.
- **e** Carbon soot is deposited on the metal spoon.

21 Determine all possible answers for each scenario from the list below.

compound
element
heterogeneous mixture
homogeneous mixture

- **a** matter that cannot be broken down to simpler substances by chemical or physical means
- **b** matter that can be separated into its constituent components by physical processes
- **c** matter that can be separated into its constituent components by chemical processes

22 Classify each of the following as a mixture or pure substance. If it is a mixture, indicate if it is heterogeneous or homogeneous.

- **a** tomato juice
- **b** a laptop
- **c** chocolate-chip ice cream
- **d** air
- **e** bromine liquid
- **f** calcium carbonate
- **g** vinegar

23 Identify the solvent and solute(s) in the following solutions.

- **a** air
- **b** sea water
- **c** vinegar
- **d** white wine
- **e** fish tank water

24 When small amounts of the following solids are mixed with water, determine which mixture is most easily separated into its constituent components. Explain your answer.

copper(II) sulfate, salt, sand, sugar

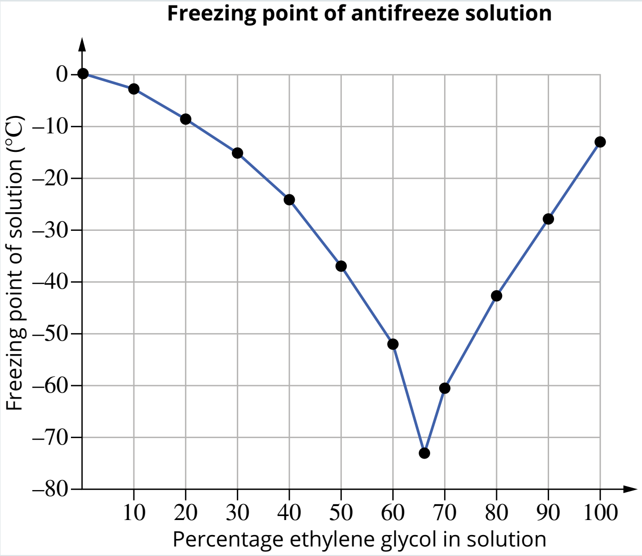
Knowledge utilisation

25 A glass contains a clear, colourless liquid that looks like water. Develop a test to describe how you can be sure that, if it is water, it is pure and does not contain any dissolved salts.

26 Propose how you could differentiate between a piece of pure silver jewellery and a piece of sterling silver jewellery.

27 Ethylene glycol is used as an antifreeze additive in vehicle radiators to stop the radiator water from freezing in cold weather. The lowest ever recorded temperature in Australia is –23.4°C recorded at Charlotte's Pass in the NSW Snowy Mountains on 29 June 1994.

- **a** Decide what percentage concentration of ethylene glycol would be required to be confident that a vehicle radiator would not freeze in Australia using the following graph.
- **b** The lowest recorded temperature in Canberra is –10°C. Determine what percentage concentration of ethylene glycol would be required to be confident that a vehicle radiator would not freeze in Canberra using the following graph.

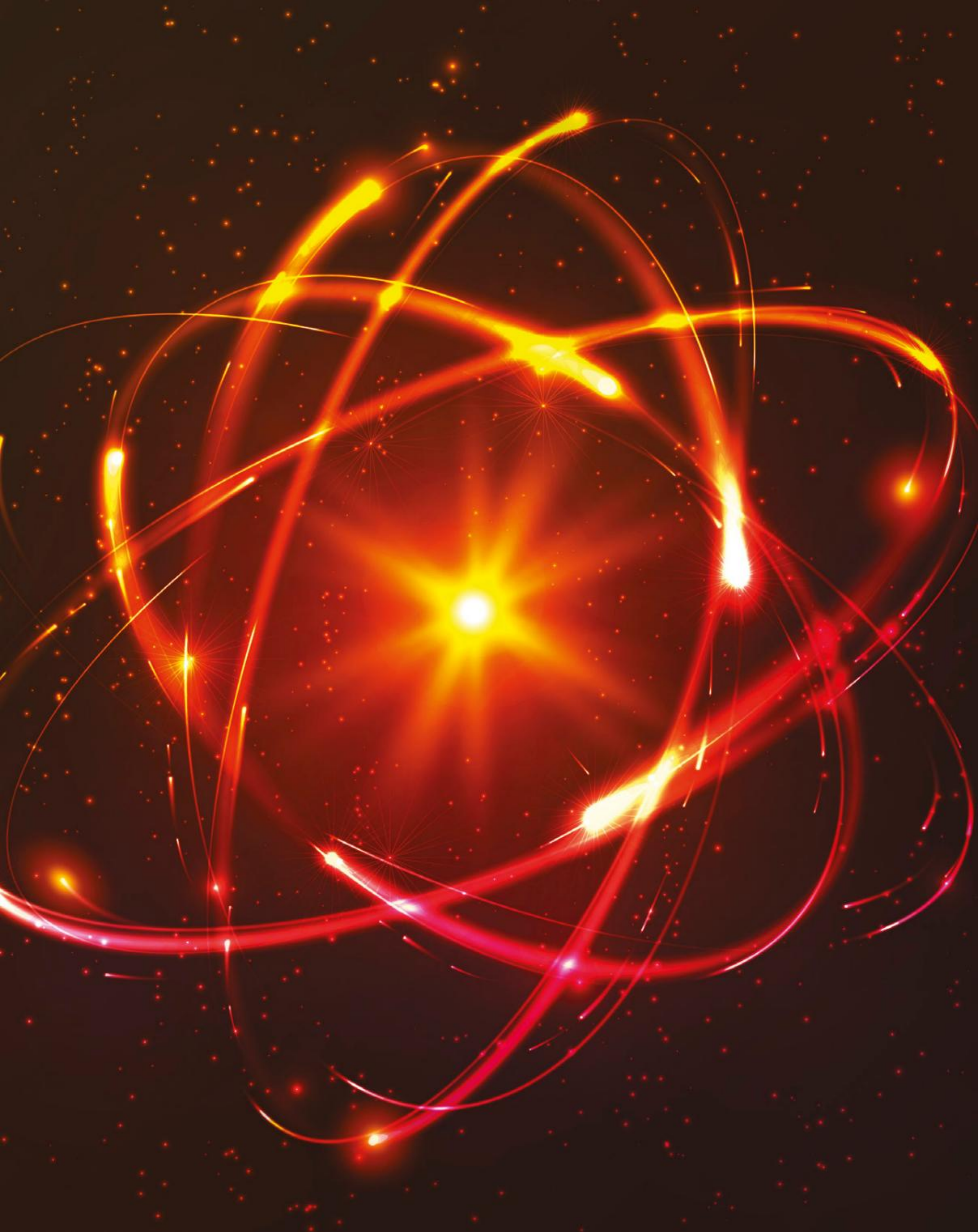


28 Design a test that would enable you to separate:

- **a** oil and water layers in salad dressing
- **b** rocks from sand.

29 Develop a test that would enable you to separate a mixture of sand and sugar crystals, ensuring you recover both the sand and sugar in their solid states.

30 A bottle of white wine (a mixture of ethanol and water) has been contaminated by dissolved salt, sand and iron filings. Develop an experiment that would enable you to separate and recover all components of the mixture. Present your answer in a visual way such as a concept map or flow chart.



CHAPTER

03

The atomic world

In this chapter, you will begin by examining how material science can create advanced materials by controlling matter on the atomic scale or as nanoparticles, with just a few hundred or thousand atoms. You will then develop a detailed knowledge of the structure of atoms, which is the foundation for all chemistry. Many of the details of this structure were described just 100 years ago when numerous ground-breaking experimental and theoretical observations were made by physicists and chemists—each one contributing to our understanding of the invisible world inside atoms.

WS
1.2.1

Syllabus subject matter

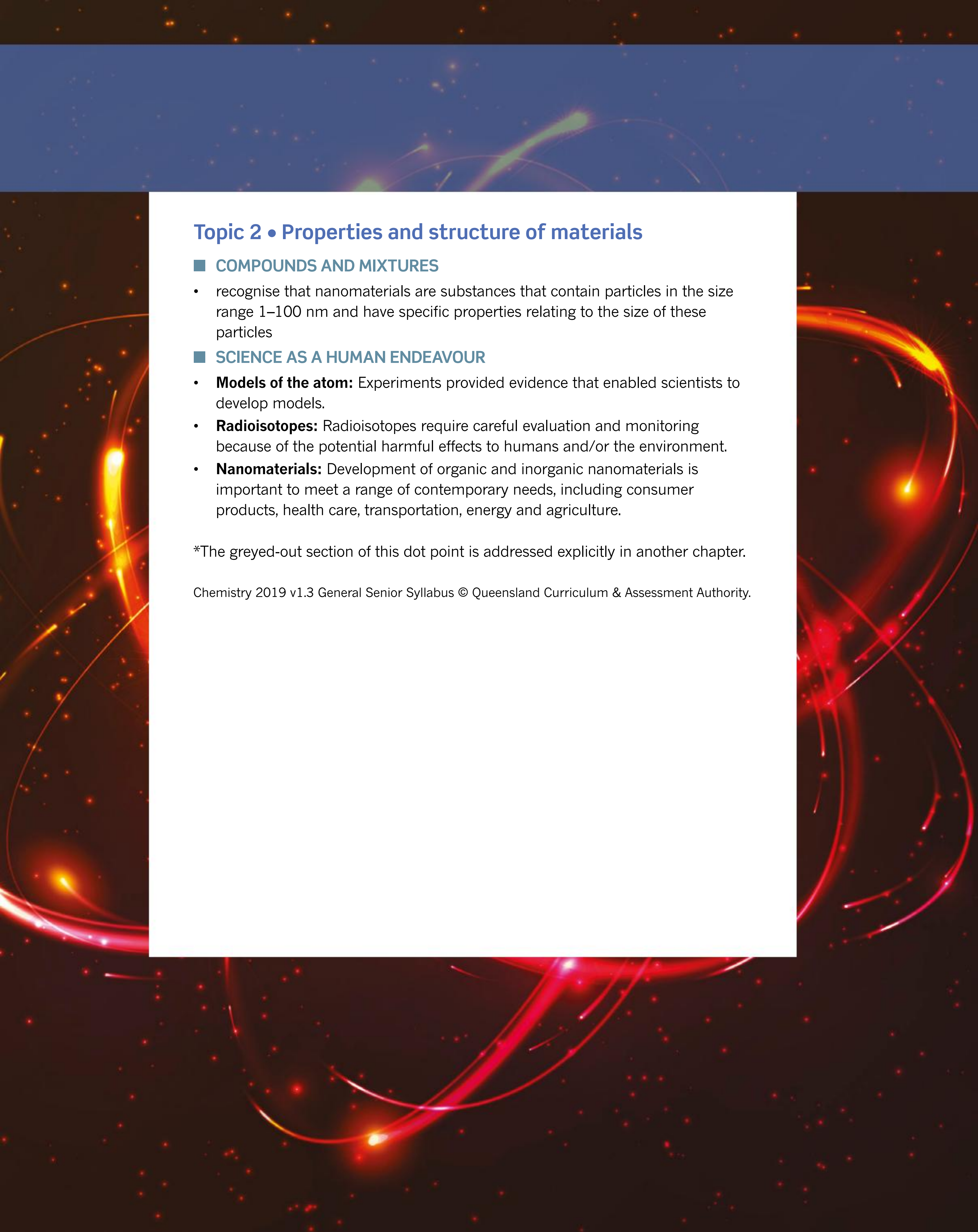
Topic 1 • Properties and structure of atoms

ATOMIC STRUCTURE

- understand that atoms can be modelled as a nucleus surrounded by electrons in distinct energy levels held together by electrostatic forces of attraction between the nucleus and electrons; the location of electrons within atoms can be represented using electron configurations; and the structure of the periodic table is based on the electron configuration of atoms
- use and apply the nuclear symbol notation $^{A}_{Z}\mathrm{M}$ to determine the number of protons, neutrons and electrons in atoms, ions and isotopes
- recall the relative energies of the s, p and d orbitals in energy levels to construct electron configurations for atoms and ions up to $Z = 36$ and recognise that the periodic table is arranged into four blocks associated with the four sub-levels—s, p, d and f*
- apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to $Z = 36$ and use orbital diagrams to represent the character and relative energy of orbitals
- recognise the electron configuration of Cr and Cu as exceptions.

ISOTOPES

- recall isotopes are atoms of the same element that have different numbers of neutrons and can be represented in the form $^{A}\mathrm{X}$ (IUPAC) or X-A
- recognise that isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties



Topic 2 • Properties and structure of materials

COMPOUNDS AND MIXTURES

- recognise that nanomaterials are substances that contain particles in the size range 1–100 nm and have specific properties relating to the size of these particles

SCIENCE AS A HUMAN ENDEAVOUR

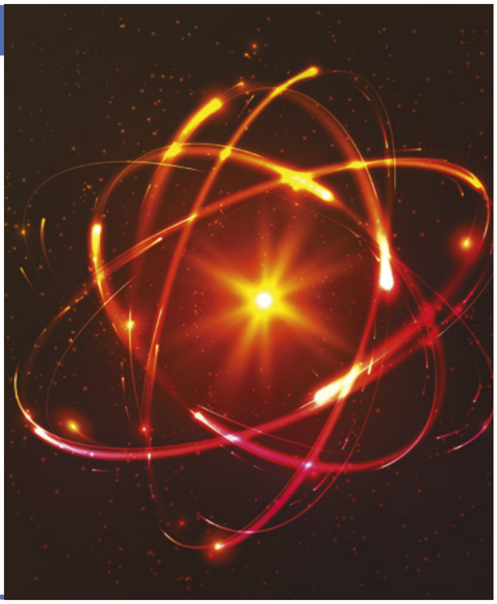
- **Models of the atom:** Experiments provided evidence that enabled scientists to develop models.
- **Radioisotopes:** Radioisotopes require careful evaluation and monitoring because of the potential harmful effects to humans and/or the environment.
- **Nanomaterials:** Development of organic and inorganic nanomaterials is important to meet a range of contemporary needs, including consumer products, health care, transportation, energy and agriculture.

*The greyed-out section of this dot point is addressed explicitly in another chapter.

3.1 Nanomaterials and nanoparticles

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand types of structures that are invisible to the eye and the scale of their size
- recognise the two approaches by which nanostructures are made
- understand that nanomaterials can have very different physical, chemical and optical properties from the bulk material
- understand that nanoparticles have a large surface area compared to their volume and so can be used to adsorb molecules, transport molecules and act as catalysts for chemical reactions
- recognise that knowledge about nanomaterials can be used to develop technologies that have projected economic, social and environmental impacts and to design sustainable processes and materials
- recognise that health and wellbeing precautions and safeguards are required due to the newness of nanomaterials technology and the ability of nanoparticles to be transported relatively easily into the human body.



One of the most significant advances in materials science in the twenty-first century has been in the area of nanoscience. **Nanoscience** looks at materials at a very small scale—the **nanoscale**. The nanoscale is larger than the atomic scale, but is still far too small to be seen with the naked eye. By modifying and engineering materials at the nanoscale, scientists can fundamentally change the properties of the materials and therefore create new technologies.

Nanoscience is the study of nanoparticles and nanotechnology. **Nanotechnology** is the use of technologies that manipulate and investigate the properties of materials on the nanoscale.

Materials engineered on the nanoscale are already an important part of your everyday life. For example, microchips in computers and phones are made up of nanoscale transistors that process information at billions of bits per second. Lasers and light-emitting diodes (LEDs) also use materials engineered on the nanoscale. Even some sunscreens are made of nanoparticles designed specifically to absorb ultraviolet radiation.

'Nanotechnology' is no longer just a term used in science fiction. It is here today, and it has fundamentally changed the world you live in.

THE NANOSCALE

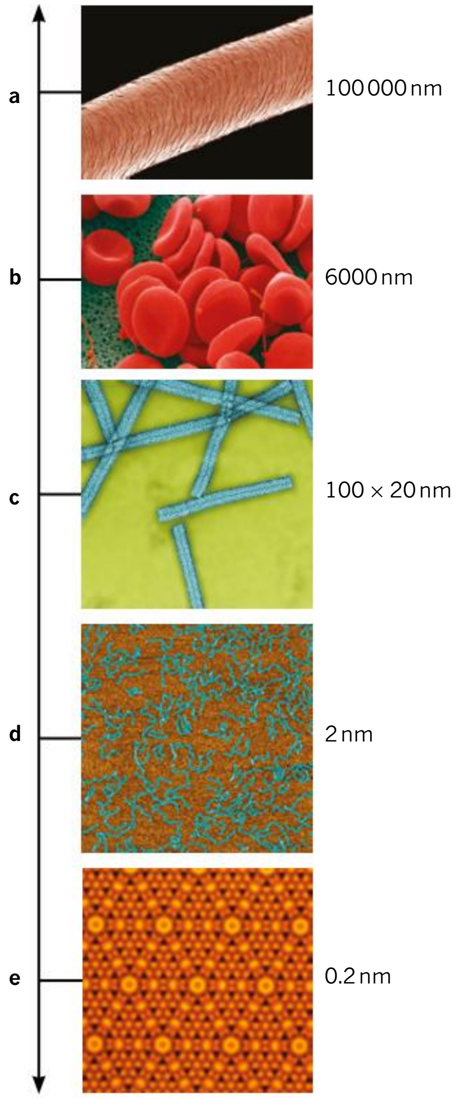


FIGURE 3.1.1 This scale compares the sizes of objects down to the nanometre size. (a) The width of a human hair is approximately 100 000 nm. (b) The diameter of red blood cells is 6000 nm. (c) Tobacco mosaic virus is 100 nm long and 20 nm wide. (d) DNA strands are about 2 nm wide. (e) Silicon atoms are 0.2 nm across.

The prefix 'nano' refers to one-billionth or 10^{-9} in scientific notation. For example, a nanometre (nm) is one-billionth of a metre or 10^{-9} m. This is much smaller than anything you can see with the naked eye. Table 3.1.1 compares metric units for length to the metre.

TABLE 3.1.1 Some metric units for length

Unit	Symbol	Relative to standard unit of length (m)
picometre	pm	10^{-12}
nanometre	nm	10^{-9}
micrometre	μm	10^{-6}
millimetre	mm	10^{-3}
metre	m	1
kilometre	km	10^{3}

Worked example 3.1.1

CONVERTING MILLIMETRES TO NANOMETRES

Convert 15 mm to nanometres.

Thinking	**Working**
Write the length in metres.	15 mm = 0.015 m
Write the length in scientific notation.	0.015 m = 1.5×10^{-2} m
Multiply by 10^{9} to convert to nanometres. (Hint: To do this, add 9 to the index on the number 10.)	$(1.5 \times 10^{-2}) \times 10^{9} = 1.5 \times 10^{(-2+9)}$ nm $= 1.5 \times 10^{7}$ nm

➤ Try yourself 3.1.1

CONVERTING MILLIMETRES TO NANOMETRES

Convert 2.43 mm to nanometres.

The width of a human hair or a pore on your skin is still thousands of times larger than the nanoscale features in **nanomaterials**. The nanoscale is used to describe objects that are about 1–100 nm wide. Figure 3.1.1 shows the sizes of different objects. You can see that tobacco mosaic virus and DNA belong to the nanoscale.

The image of the silicon atoms in Figure 3.1.1e was taken by a **scanning tunnelling microscope (STM)** developed by IBM researchers Gerd Binnig and Heinrich Rohrer. Using STMs, scientists confirmed that atoms are indeed spherical. STMs use an extremely sharp metal tip to detect atoms. The tip is scanned, line-by-line, across the surface of a crystal. As the tip moves, the tip measures minute height differences in the crystal's surface due to the individual atoms. This is similar to the way sight-impaired people use their finger to sense braille on a page. The data from the tip is then sent to a computer that constructs an image of the atoms.

NANOMATERIALS

Nanomaterials can have very different properties from those of the same material in bulk form. For example, carbon readily forms nanomaterials, such as the carbon nanotubes shown in Figure 3.1.2. These nanostructures are nothing like the soft, powdery charcoal that is normally associated with carbon. **Carbon nanotubes** are extremely strong, like diamond, and conduct electricity, like graphite. But unlike the carbon constructs of diamond and graphite, they are flexible and have a high tensile strength (resistance to stretching) and their conductivity can be engineered to make them insulating, conducting or semiconducting. You will look more closely at carbon nanomaterials in Chapter 8 and in Unit 2, Chapter 14.

The term ‘nanomaterial’ is usually associated with modern, high-tech materials. However, nanomaterials have always been present in nature. For example, milk contains nanoparticles of proteins and fats. These particles scatter light and make milk appear white. The surface of opals and butterfly’s wings diffract light to produce vibrant colours (Figure 3.1.3). Figure 3.1.4 shows samples containing gold nanoparticles of different sizes. All of the particles absorb white light, but they reflect different-coloured light, depending on the particle size.

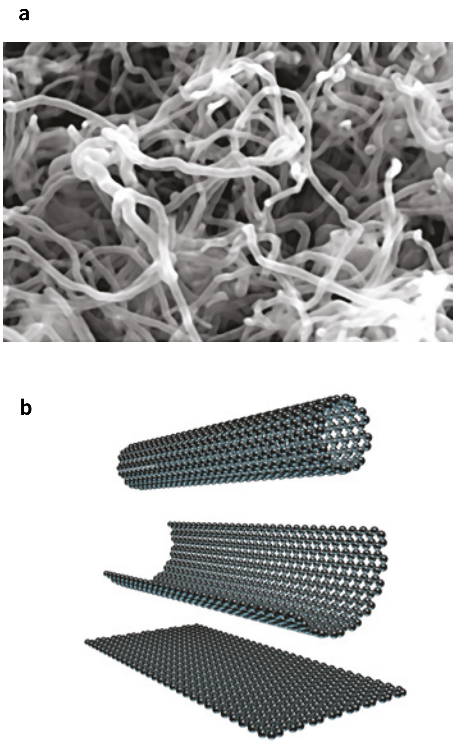


FIGURE 3.1.2 (a) These carbon nanotubes seen under a scanning electron microscope and (b) graphene (sheet version) have different properties from other materials made from carbon such as charcoal, graphite and diamond. Carbon nanotubes vary in size but are often 1 nm in diameter.

FIGURE 3.1.4 These vials contain gold particles of various sizes. The different-sized particles are different colours because they interact with light differently.

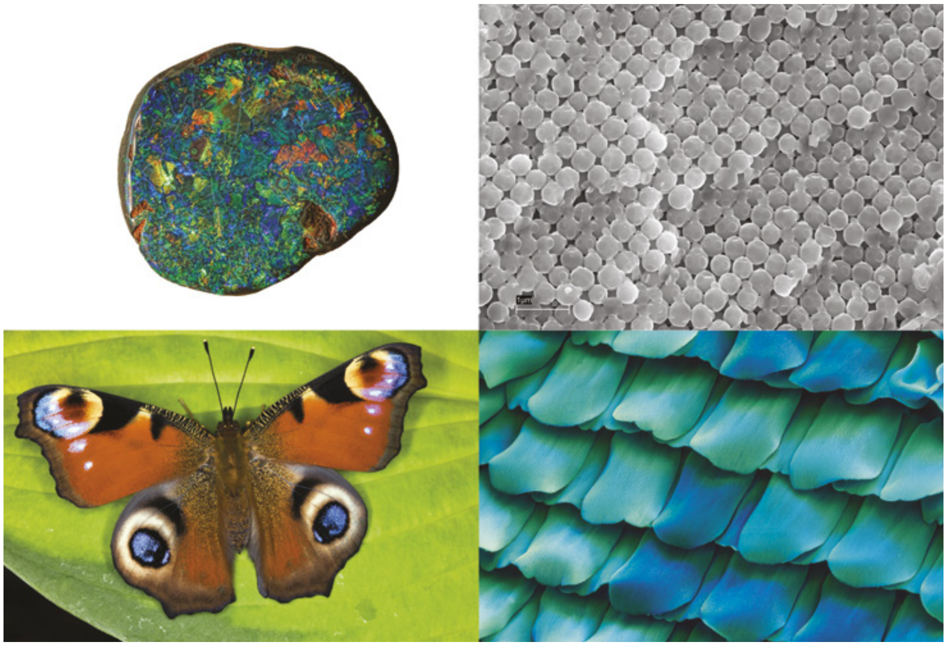


FIGURE 3.1.3 (top) An opal is made of tiny spherical nanoparticles of silica. (bottom) The colour of a butterfly’s wing is due to nanostructures in the wing. These spheres diffract light to produce spectacular flashes of red, green and blue.

Forming nanomaterials

Two processes are used to manufacture nanomaterials: the bottom-up method and the top-down method.

- The bottom-up method uses specially designed **molecules** or atoms in chemical reactions to gradually build up the new nanoparticle from the smaller atoms or molecular units.
- The top-down method uses the larger bulk material as the starting material. The bulk material is broken down into nanoparticles by mechanical or chemical means.

Even though the two methods start at opposite ends of the scale, the bottom-up and top-down approaches both produce materials that exist on the nanoscale and rely on physical and chemical processes to achieve these nanomaterials.

Manufactured nanomaterials

FIGURE 3.1.5 The nanostructure gives the silica gel an enormous surface area compared to its volume.

Scientists are learning more about how the structure of nanomaterials influences the properties of materials. This has allowed scientists to create new technologies based on nanomaterials. For example, materials such as silica gel can adsorb large amounts of liquids and gases. Silica gel sachets like the one in Figure 3.1.5 are often placed with tablets and capsules to keep them dry. The silica gel works by adsorbing water from the atmosphere so that humidity does not affect the medications. Silica gel can adsorb large amounts of water because of its nanostructure.

Materials with a **porous** structure contain billions of nanoscale holes that trap unwanted molecules such as water and gases. In activated charcoal, the holes are so tiny and so numerous that one grain can have a surface area of several square metres. This means that only a small amount of activated charcoal is needed to adsorb a large number of gas molecules. For this reason, activated charcoal is used to store natural gas. Activated charcoal is also used in gas masks to adsorb dangerous gases from the atmosphere. This type of mask was used during World War I to protect soldiers against poisonous gases.

Engineered nanomaterials

Scientists have been able to engineer nanoscale features on the surfaces of materials to give the materials new and useful properties. A simple example is the patterning of the surface of a CD-ROM. Information is encoded on the surface of the CD-ROM as a series of nano-sized 'pits' or hollows, as shown in Figure 3.1.6. Improvements in the technology have meant that smaller pits can be made so that more information can be stored on a surface. This has led to the development of DVD and Blu-ray disks.

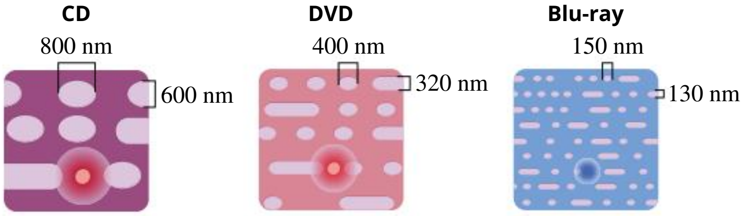


FIGURE 3.1.6 CD-ROMs store information as a series of nano-sized pits that are approximately 800 nm by 600 nm. DVDs can store more information because the pits are smaller: 400 nm by 320 nm. Blu-ray disks can store even more information because their pits are only 150 nm by 130 nm.

NANOPARTICLES

Nanoparticles are a specific type of nanomaterial with unique properties and a broad range of applications. Nanoparticles are usually spherical with diameters of 1–100 nm. They are used in chemotherapy and sunscreens and as industrial catalysts. Nanoparticles range in diameter from 1 to 100 nm.

The importance of surface area

One of the most important properties of nanoparticles is their large surface area compared to their volume. In chemistry, surfaces are extremely important since surfaces are often where chemical reactions take place. Substances made of nanoparticles have very large surface areas because the component parts are so small.

For example, consider a 1 m^3 cube. It has six sides, each exactly 1 m in length, and so has a total surface area of 6 m^2 and a volume of 1 m^3. If the cube is cut into eight cubes with side length 0.5 m, the total volume of all the cubes stays the same (1 m^3), as shown in Figure 3.1.7. However, the surface area has doubled. This is because parts of the cube that were originally on the inside of the cube have now become part of the surface. The same block divided into cubes with side length 100 nm has a surface area of 60 000 000 m^2 but the volume is still 1 m^3.

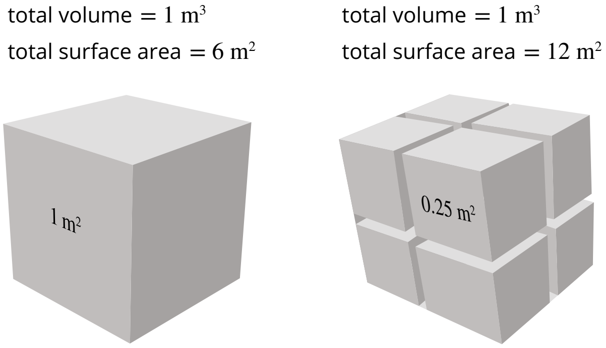


FIGURE 3.1.7 When a 1 m cube is divided into eight equal cubes, the volume stays the same but the surface area doubles.

A large surface area to volume ratio is one of the most important features of nanoparticles.

The large surface area of nanoparticles is important in three ways.

1 **Adsorption** of molecules: The large surface area of nanoparticles means that even a small volume of nanoparticles can adsorb a large number of molecules. Therefore, nanoparticles, such as silica gel and activated charcoal, can be used to remove unwanted chemicals and gases.
2 **Transportation** of molecules: Nanoparticles can be used to transport the molecules that are adsorbed onto the nanoparticle. The nanoparticles' small size and large surface area mean that they can transport chemicals through air, skin and even cells. This feature of nanoparticles has a variety of applications in medicine for transporting drugs into the body and is already being used in some chemotherapy treatments for cancer.
3 **Catalysts**: The surfaces of some nanoparticles can be used to increase the rate of particular chemical reactions. The nanoparticles are acting as catalysts because they speed up the reaction but are not consumed (used up) by the reaction. The reactant molecules adsorb onto the surface of the nanoparticles, which allows reactant molecules to combine to form the product. The large surface area of nanoparticles means that many of these reactions can take place at the same time, which increases the rate of the reaction significantly.

The term 'adsorption' is used when a molecule attaches to a surface of a solid or liquid molecule. This is different from 'absorption', where molecules are incorporated into the substance. Adsorption is reversible but absorption is often not reversible.

Other properties and applications

Nanoparticles have many other useful properties. Their small size and unique structural properties mean they can be used for developing **composite** (made up of several parts) nanomaterials. Composite nanomaterials incorporate nanoparticles into other materials.

The stain-resistant cotton shown in Figure 3.1.8 on page 34 is a composite nanomaterial. In this case, the cotton fibres are covered with water-resistant nanoparticles so that liquids cannot penetrate the material. This makes the fabric water-repellent and stain-resistant.

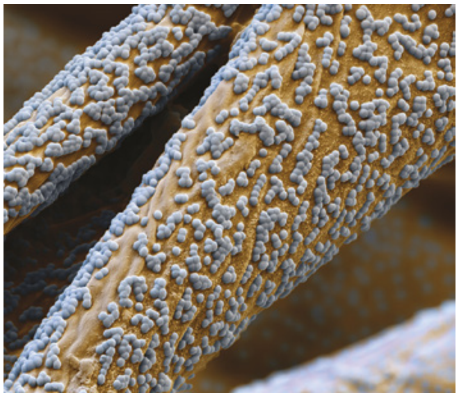


FIGURE 3.1.8 Water-repellent nanoparticles can be incorporated into cotton fibres to create a composite nanomaterial that is a water-resistant fabric.

When a substance is processed into nanoparticles, its physical, chemical and optical properties can change. For example, when gold is in the form of nanoparticles, it loses its metallic lustre. Instead, it can take on a variety of colours depending on the size of the gold particles. For this reason, gold nanoparticles are often used to colour glass.

The fact that nanoparticles interact differently with light makes them useful in sunscreens. Zinc oxide (ZnO) nanoparticles are commonly used in sunscreens. Zinc oxide absorbs a broad range of the ultraviolet spectrum and so provides excellent protection from the sun. Normally, sunscreens that contain metal oxides are milky white when applied to the skin. However, sunscreen that contains zinc oxide nanoparticles is clear, and therefore is preferred by some beach-goers.

Nanoparticles: health and environmental risks

Nanoparticles have opened up a range of technological possibilities. However, some people are concerned about the possible dangers to humans and the environment. Other useful materials such as asbestos have been found to have devastating side-effects. Scientists at Australia's Commonwealth Scientific and Industrial Research Organisation (CSIRO) and around the world are studying the potential dangers associated with nanoparticles and their applications.

Nanoparticles are so small that they can travel through the air, through skin and into your bloodstream and even into cells. Inside the body, the particles may interact with biomolecules to cause unwanted chemical reactions. This makes nanoparticles potentially dangerous if breathed in or if they come into contact with the skin (for example, in sunscreen, fabrics or cosmetics).

Scientists at CSIRO are looking at the zinc oxide nanoparticles used in sunscreens to determine whether they are safe. This research focuses on whether the nanoparticles can penetrate skin, their long-term health effects and how they might impact on the environment.

Initial studies suggest that small amounts of zinc oxide from sunscreens are absorbed into the body and can be detected in the blood and urine. It is still not clear whether the absorbed zinc oxide has any negative effects on the human body. The most recent research indicates that the cells of the immune system can break down these nanoparticles.

SCIENCE AS A HUMAN ENDEAVOUR

Nanomaterials: Australian Institute for Biotechnology and Nanotechnology

Professor Mark Kendall (Figure 3.1.9) was awarded the 2016 CSL Young Florey Medal by the Australian Institute of Policy and Science for his work in developing the Nanopatch™. This vaccinating device is at the interface of engineering, chemistry, biomaterials and immunology. Professor Kendall, with his team and collaborators, was able to achieve immune responses equivalent to a needle and syringe but using a painless skin patch and lower doses of vaccine. The Nanopatch (Figure 3.1.10) array of microprojections penetrates through the protective outer skin layer (stratum corneum) and delivers immune-activating material to the immune-cell rich layers just beneath the skin. The device is currently undergoing clinical trials, and is also being tested in collaboration with the World Health Organization for vaccination against poliovirus. This gives it the potential to improve the health of millions of people worldwide.

FIGURE 3.1.9 Professor Mark Kendall of the Australian Institute for Biotechnology and Nanotechnology

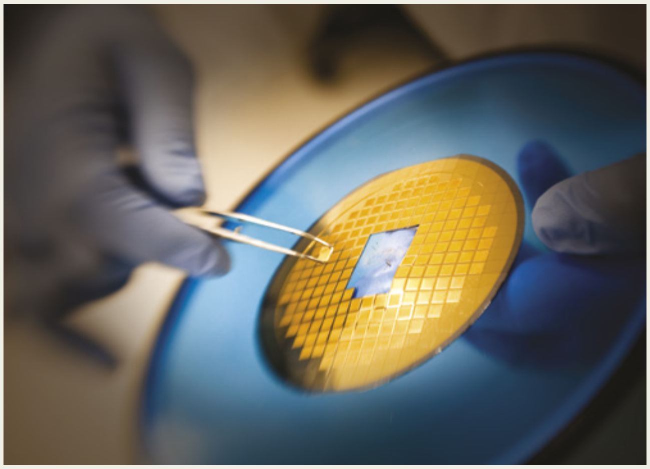


FIGURE 3.1.10 The Nanopatch consists of a silicon surface with approximately 20 000 microprojections per cm^2. The microprojections are invisible to the naked eye.

Professor Chengzhong Yu (Figure 3.1.11 on page 36) was awarded the 2015 Le Févre Memorial Prize by the Australian Academy of Science. Professor Yu and his team of scientists are investigating the use of nanoparticles to improve the delivery of oral medications. Oral medications are subjected to changes in pH (high acid levels in the gut) and enzymatic breakdown in the mouth and stomach. This acid and enzyme attack means that many protein-based medications break down before they have a chance to act on the specific target tissue. By associating medicines with nanoparticles, Professor Yu hopes to protect and prolong the activity of the drug when in the digestive system and other parts of the body. His team is developing these new, cost-effective nanomaterials to improve delivery of vaccines, genes and drugs for human and animal health care.

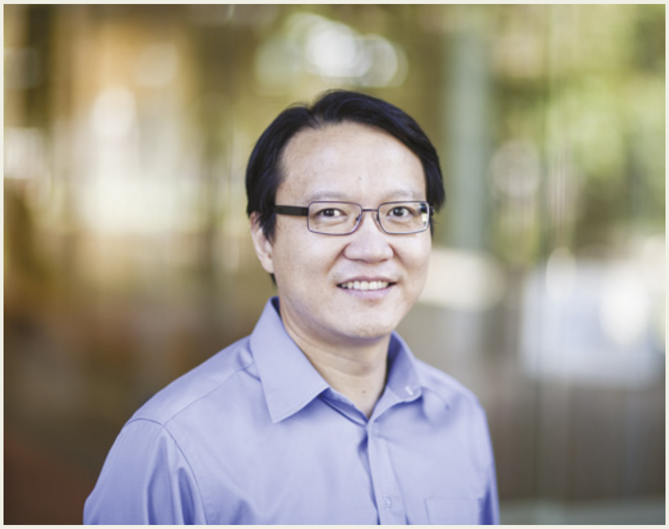


FIGURE 3.1.11 Professor Chengzhong Yu of the Australian Institute for Biotechnology and Nanotechnology

Another of the group's projects is the design of nanoparticles that mimic a virus to deliver medicines to diseased or damaged cells. Figure 3.1.12 shows the targeted destruction of a tumour by substances delivered by nanoparticles. In medicine, nanotechnology could improve the analysis and treatment of cancers, genetic disorders and viral infections once thought untreatable because cytotoxic drugs failed to enter target cells.

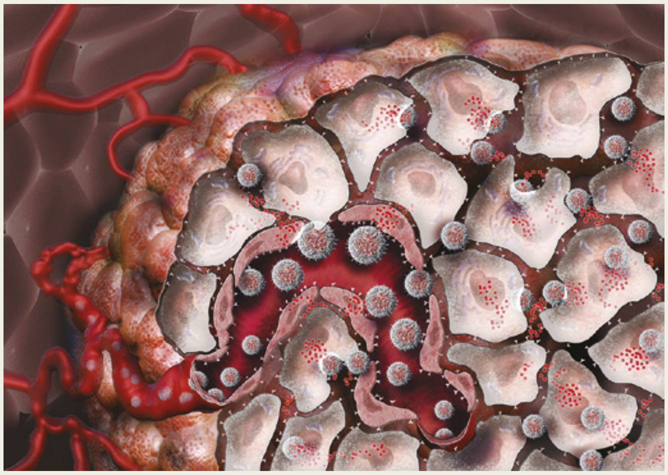


FIGURE 3.1.12 This computer illustration shows nanoparticles (spheres) containing cytotoxic drugs. The nanoparticles can leave the blood vessels (U-shaped structure in the centre of the diagram) and target tumour cells, leading to their destruction.

Review

1. In delivery of medications, researchers are developing techniques that attach a drug to a nanoparticle, have it taken up by a target cell and, finally, have the drug released from the nanoparticle.
   - **a** Using a carbon nanotube as an example, describe two possible ways that a drug molecule could be associated with the nanotube.
   - **b** Investigate virus structure and suggest why the nanoparticle in part **a** can mimic viral infectivity and enter cells.
2. The possibilities of carbon nanoparticles are of keen interest to many research groups wanting to treat disease. Review the properties of these particles and describe a reason that they are unlikely to damage body tissues. Propose an area of concern for long-term treatment of a patient when using nanoparticles.

3.1 Review

SUMMARY

- Our understanding of the range of structures that are invisible to the unaided eye has expanded from atoms and molecules to include nanoparticles and nanostructures.
- Nanoparticles have diameters of 1–100 nm. A nanometre is 10^{-9} metres.
- Nanoparticles can have very different physical, chemical and optical properties from the bulk material.
- Nanoparticles can be made in two ways: the bottom-up and top-down methods.
- Nanoparticles have a large surface area compared to their volume and so can be used to adsorb (hold on their surface) molecules, transport molecules and act as catalysts for chemical reactions.
- Nanoparticles can transport substances through the air and into the human body. This has raised concerns about their safety for human health and the environment.
- Investigations into properties of engineered particles aim to address the potential risks of nanomaterials.

KEY QUESTIONS

Retrieval

1 Convert 8.35 cm to nanometres.

2 Convert the following lengths to nanometres. Use scientific notation in your answers.

  a 1.35 cm

  b 4.2 mm

Comprehension

3 The thickness of a sheet of paper is approximately 100 000 nm. Determine how many sheets would be in a pile of paper 2.5 cm high.

4 A human hair is about 0.050 mm wide, whereas a tobacco mosaic virus is 20 nm wide. Determine how many times larger the width of a human hair is compared to the width of a tobacco mosaic virus.

5 The figure in the image below is holding a carbon nanotube. Use information provided in this chapter to determine an approximate value for the height of the figure in metres.

6 ‘The surface area to volume ratio decreases as the size of the object decreases.’ Determine the validity of this statement.

7 Explain why the size of nanoparticles makes them useful for transporting medicines into the body.

Analysis

8 A body of a flea (such as that enlarged in the figure below) is approximately 2×10^6 nm long and 10^6 nm wide. Deduce if the flea could hide behind any of the structures shown in Figure 3.1.1 on page 30.

3.1 Review *continued*

9 **a** For each shape in the following table, calculate the surface area, volume and then surface area to volume ratio.

Shape	**Surface area**	**Volume**	**Surface area to volume ratio**
cube 2 cm × 2 cm × 2 cm			
sphere radius 1.38 cm			
tube (cylinder) radius 1 cm height 2 cm			

Hint: the following formulas may be useful.

- surface area of a sphere = $4\pi r^2$
- surface area of a cylinder = $2\pi rh + 2\pi r^2$
- volume of a sphere = $\frac{4}{3}\pi r^3$
- volume of a cylinder = $\pi r^2 h$

b From the table, identify which shape has the greatest surface area to volume ratio.

c State which shape would be the most useful to hold the greatest volume.

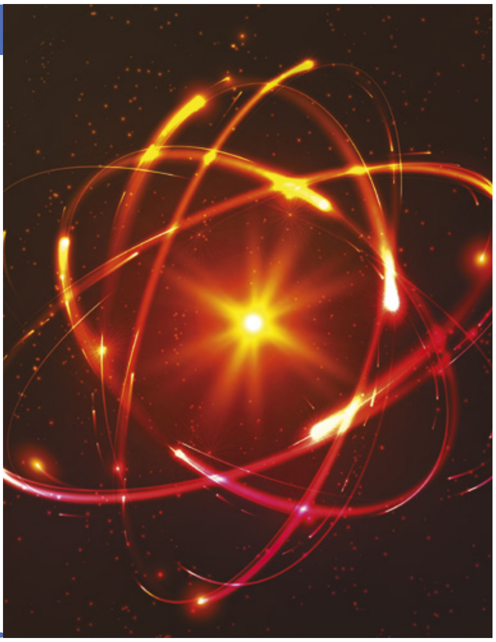
d Suggest an application where a large surface area to volume ratio is important and another application where a small surface area to volume ratio is important.

10 Some coloured compounds change colour when they absorb water, but do not change colour when they adsorb water. Apply your understanding of these two processes, and explain why this would be the case.

3.2 Inside the atom

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand how the current model of the atom shows that it is composed of three types of subatomic particles (protons, neutrons and electrons) held together by electrostatic attraction
- describe each subatomic particle's mass, charge and position in the atom
- recall that elements are defined by the number of protons—the atomic number. The mass number is the number of particles in the nucleus (protons and neutrons)
- understand that most of the mass is in the nucleus and that the cloud of electrons contains shells or layers, each one having a particular energy level
- understand that the lower energy shells are closer to the nucleus and that energy is given off or absorbed when an electron moves to a different shell of higher or lower energy
- recall the defined number of electrons that can exist in each shell (*n*)
- relate the size of ionisation energy to the shell of an electron.



The development of advanced materials such as nanoparticles is the result of centuries of scientific discovery. Over time, scientists have gained a deep understanding of the structure of atoms, which are the basic building blocks of **matter**. As scientists' understanding has increased, their ability to control matter on the atomic scale has also improved. However, atoms are too small to be seen and so much of what scientists know about atoms has come from theoretical models and indirect observations. The STM images obtained in 1981 were confirmation of the spherical shape of atoms—a **scientific model** that had been proposed nearly 200 years before by John Dalton. Further evidence suggested that atoms are made of smaller particles—more detail of these historical experiments is provided in Module 3.5.

STRUCTURE OF ATOMS

All atoms are made up of a small, positively charged nucleus surrounded by a much larger cloud of negatively charged electrons as shown in Figure 3.2.1. The **nucleus** is in turn made up of two types of **subatomic particles—protons** and **neutrons**. The protons are positively charged and the neutrons have no charge.

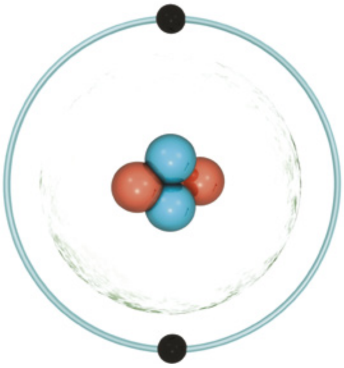


FIGURE 3.2.1 A simplified model of the atom. This shows a helium atom with two protons and two neutrons (in the central nucleus) and two electrons. The outer ring symbolises that the electrons are held in an orbit; however, evidence suggests they move throughout an area that is more like a cloud.

A scientific model is a description that scientists use to represent the important features of what they are trying to describe. This is based on experimental observations. They can then test the consistency of their observations against various predictions of the model.

The nucleus

The nucleus of an atom is approximately 10 000–100 000 times smaller than the size of the atom. To put this in perspective, if an atom were the size of the Woolloongabba Cricket Ground (Figure 3.2.2), then the nucleus would be about the size of a pea. Nonetheless, the nucleus contributes around 99.97% of the atom's mass. This means that atomic nuclei are extremely dense.

FIGURE 3.2.2 If an atom were the size of the Woolloongabba Cricket Ground, then the nucleus would be the size of a pea, and the electrons would be found in a stadium-sized cloud.

The subatomic particles in the nucleus, the protons and neutrons, are referred to collectively as **nucleons**. Protons are positively charged particles with a mass of approximately 1.673×10^{-27} kg or 1 atomic mass unit (amu). Neutrons and protons are almost identical in mass. Table 3.2.1 summarises the properties of protons, neutrons and electrons.

TABLE 3.2.1 Properties of the subatomic particles

Particle	Symbol	Charge	Mass relative to a proton	Mass (kg)
proton	p	+1	1	1.673×10^{-27}
neutron	n	0	1	1.675×10^{-27}
electron	e	–1	$\frac{1}{1800}$	9.109×10^{-31}

Electrons

Electrons are negatively charged particles. You can imagine them forming a cloud of negative charge around the nucleus. This cloud gives the atom its size and volume.

An electron is approximately 1800 times lighter than a proton or neutron. Therefore, electrons contribute very little to the total mass of an atom. However, the space occupied by the cloud of electrons is 10 000–100 000 times larger than the nucleus.

Positive particles attract negative particles. This is called **electrostatic attraction**. Electrons are bound to the nucleus by electrostatic attraction to the protons within the nucleus. The charge on an electron is equal to but opposite the charge of a proton. Electrons are said to have a charge of −1, whereas protons have a charge of +1. This means that all neutral **elements** have equal numbers of protons and electrons.

In some circumstances, electrons can be removed from an atom. For example, when you rub a rubber balloon on a woollen jumper or dry hair, electrons are transferred to the balloon and the balloon develops a negative charge. The negative charge is observed as an electrostatic force that can stick the balloon to a wall or even move an aluminium can. You will look more closely at the removal of electrons from atoms when learning about ionic compounds in Chapter 7.

The electricity that powers lights and appliances is the result of electrons moving in a current through wires. Sparks and lightning are also caused by electrons moving through air.

ELECTRON SHELLS

Beginning with the hydrogen atom and extending to many other elements, scientists in the early twentieth century observed energy being released in 'packets' or **quanta** (observed as particular wavelengths of light). Niels Bohr proposed that electrons were grouped in different **energy levels**, called **electron shells**.

In the **Bohr model**, the shell in which an electron moved depended on the energy of the electron; electrons with low energy are in shells close to the nucleus, while high-energy electrons are in outer shells (Figure 3.2.3). Due to the difference in energy levels between shells, if an electron were to move to a different shell, energy must be absorbed or released. Electrons in the same shell share these characteristics.

- They are about the same distance from the nucleus.
- They have about the same energy.

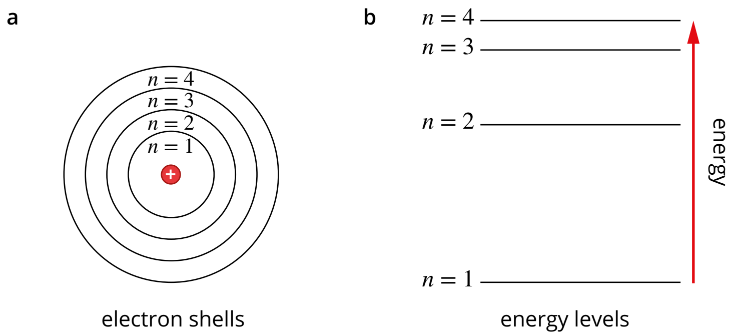


FIGURE 3.2.3 (a) In the Bohr model of the atom the electron shells are labelled using integers (n). The first shell is the shell closest to the nucleus and the radii of the shells increase as the shell numbers increase. (b) Each shell corresponds to an energy level that electrons can occupy. The first shell has the lowest energy level and the energy level increases as the shell number increases.

The lowest energy shell is the shell closest to the nucleus and is labelled $n = 1$. Shells with higher values of n correspond to higher energy levels (Figure 3.2.3). As the values of n increase, the energy levels get closer together.

When fireworks explode, they create a spectacular show of coloured lights (Figure 3.2.4 on page 42). The light is produced by metal atoms that have been heated by the explosion. Electrons, when heated, 'jump' to a higher energy level (**excited state**) and then release light energy as they return to their lower energy level (**ground state**). Because each metal element has a different electron arrangement a different energy (colour) of light will be released when electrons undergo this change. These unique **emission spectra** are further described in Chapter 5.

FIGURE 3.2.4 The spectacular colours from a New Year's Eve fireworks display in Brisbane. As the metal atoms have been heated to very high temperatures some of their electrons have moved to higher energy levels. When the electrons return to lower energy levels in the metal atom, light of a particular wavelength (colour) is emitted.

In all atoms, the electrons are as close to the nucleus as possible. This means that electrons will generally occupy inner shells before outer shells. For example, an atom of lithium has three electrons. Two electrons will occupy the $n = 1$ shell as that is all it can hold. The third electron is in the $n = 2$ shell.

A Bohr diagram is a simple diagram that shows the arrangement of electrons around the nucleus. In such diagrams, only the shells that are occupied are drawn. The Bohr diagram for lithium is shown in Figure 3.2.5.

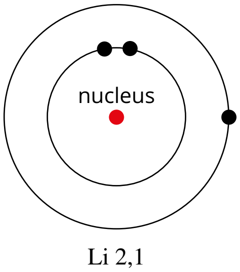


FIGURE 3.2.5 The Bohr diagram showing the electron shells of the lithium atom (three electrons). Protons and neutrons are not shown.

You have learnt that each shell has a particular energy level. These shells will only hold a particular number of electrons. The maximum number of electrons that a shell can hold can be calculated using $2n^2$, where n is the shell number. For example, for the $n = 1$ shell, $2n^2$ is two electrons; whereas for the $n = 2$ shell, $2n^2 = 8$ electrons. The arrangement of electrons in shells for a given atom is called its **electron configuration**.

It is possible to determine the electron configuration of any atom as each shell can only contain a maximum number of electrons. Table 3.2.2 shows how many electrons can occupy the first four shells.

TABLE 3.2.2 The number of electrons held by each shell of an atom. Shell 1 is the shell closest to the nucleus.

Electron shell (n)	Maximum number of electrons
1	2
2	8
3	18
4	32
n	$2n^2$

Oxygen has eight electrons, so $n = 1$ shell will hold two electrons and $n = 2$ shell will hold six electrons.

Figure 3.2.6 shows the Bohr diagram of a sodium atom, which has 11 electrons. The first two electrons fill $n = 1$ shell. Then the next eight electrons fill $n = 2$ shell. The remaining electron occupies the $n = 3$ shell.

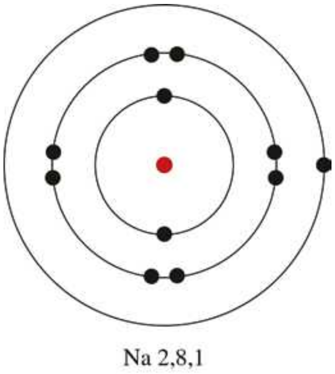


FIGURE 3.2.6 Bohr diagram of the electron configuration of a sodium atom

Worked example 3.2.1

ELECTRON CONFIGURATIONS FOR UP TO 18 ELECTRONS

Apply the rules of the shell model to determine the electron configuration of an atom with 16 electrons.

Thinking	**Working**
Recall the maximum number of electrons that each shell can hold.	shell (n) — maximum number of electrons 1 — 2 2 — 8 3 — 18 4 — 32
Place the 16 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed. Note shell 3 could hold 18 electrons, but in this atom it only has 6 electrons.	shell (n) — electrons in shell 1 — 2 2 — 8 3 — 6

➤ Try yourself 3.2.1

ELECTRON CONFIGURATIONS FOR UP TO 18 ELECTRONS

Apply the rules of the shell model to determine the electron configuration of an atom with 13 electrons.

WRITING ELECTRON CONFIGURATIONS

Diagrams originally used by Bohr show the arrangement of electrons in each energy level (shell) and are also called shell diagrams. However, the use of **electron configuration notation** is more efficient. Writing the number of electrons in each shell in order of filling and separated by a comma describes the electron configuration of the atom.

In Figure 3.2.6, you can see that the electron configuration of the sodium atom is 2,8,1. This indicates that the sodium atom has two electrons in the $n = 1$ shell, eight electrons in the second $n = 2$ shell and one electron in the $n = 3$ shell.

The electron configuration for this atom is written as 2,8,1 to indicate the arrangement of electrons in shells.

Worked example 3.2.2

ELECTRON CONFIGURATIONS FOR UP TO 18 ELECTRONS

Apply the rules of the shell model to determine the electron configuration of an atom of argon.

Thinking	**Working**
Recall the atomic number of argon.	18
Recall the maximum number of electrons that each shell can hold.	shell (*n*) — maximum number of electrons 1 — 2 2 — 8 3 — 18 4 — 32
Begin to place the 18 electrons in the shells beginning with the lowest energy ($n = 1$) to the highest energy. Shell 1 can only hold two electrons.	shell (*n*) — electrons in shell 1 — 2 2 3 4
Place all 18 electrons in this way by proceeding to shell 2 ($n = 2$) and shell 3 ($n = 3$). Do not exceed the maximum number of electrons allowed.	shell (*n*) — electrons in shell 1 — 2 2 — 8 3 — 8 4
Write the electron configuration by listing the number of electrons in each shell separated by commas.	The electron configuration is 2,8,8.

➤ Try yourself 3.2.2

ELECTRON CONFIGURATIONS FOR UP TO 18 ELECTRONS

Apply the rules of the shell model to determine the electron configuration of an atom of fluorine.

Valence electrons

The **outermost shell** of an atom is known as the atom's **valence shell**. The electrons in the outer shell are called **valence electrons**. These electrons require the least amount of energy to be removed from the atom. The valence electrons are involved in chemical reactions. Consequently, if you know the number of valence electrons in the atoms of an element, then you can predict the chemical properties of the element and the bonds that it will form.

In Chapters 7 and 8, you will learn how atoms tend to lose, gain or share valence electrons to achieve eight electrons in their outer shell when they bond with other atoms. This is known as the **octet rule**. The arrangement of elements in the **periodic table** is closely linked to the number of valence electrons (see Chapter 4).

Worked example 3.2.3

CALCULATING THE NUMBER OF VALENCE ELECTRONS IN AN ATOM

Determine the number of valence electrons present in an atom of magnesium, which has an atomic number of 12.

Thinking	**Working**
Recall the maximum number of electrons that each shell can hold.	shell (*n*) — maximum number of electrons 1 — 2 2 — 8 3 — 18 4 — 32
Place 12 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.	shell (*n*) — electrons in atom 1 — 2 2 — 8 3 — 2 4
Determine the number of electrons in the outer shell or valence shell.	The number of electrons in the valence shell is 2.

➤ Try yourself 3.2.3

CALCULATING THE NUMBER OF VALENCE ELECTRONS IN AN ATOM

Determine the number of valence electrons present in an atom of phosphorus, which has an atomic number of 15.

Ionisation energy

Evidence for the existence of energy levels in atoms was obtained from studies of successive ionisation energies in atoms of different elements. The **ionisation energy** is the energy needed to remove an electron from an atom. For example, a sodium atom contains 11 electrons and its electron configuration is 2,8,1. As each successive electron is removed, these ionisation energies can be measured. Figure 3.2.7 shows that the first electron to be removed has the lowest ionisation energy. It is easiest to remove because it is in the outermost shell, $n = 3$. The following eight electrons are in the $n = 2$ shell and are more difficult to remove. Finally, the last two require substantially more energy to remove as they are in the innermost shell, $n = 1$. In this position, closer to the nucleus, the electrostatic attraction is much greater and so much more energy needs to be used to remove each electron.

Once scientists had determined the successive ionisation energies for many elements, they concluded that electrons were grouped in different energy levels. Figure 3.2.7 shows three energy levels in the sodium atom with eight electrons in the middle energy level (the second to the ninth electrons). The evidence for this is that the second to ninth electrons need a similar amount of energy to be removed (ionised), and there is subsequently a large jump in energy required to remove the last two electrons from the inner shell, $n = 1$.

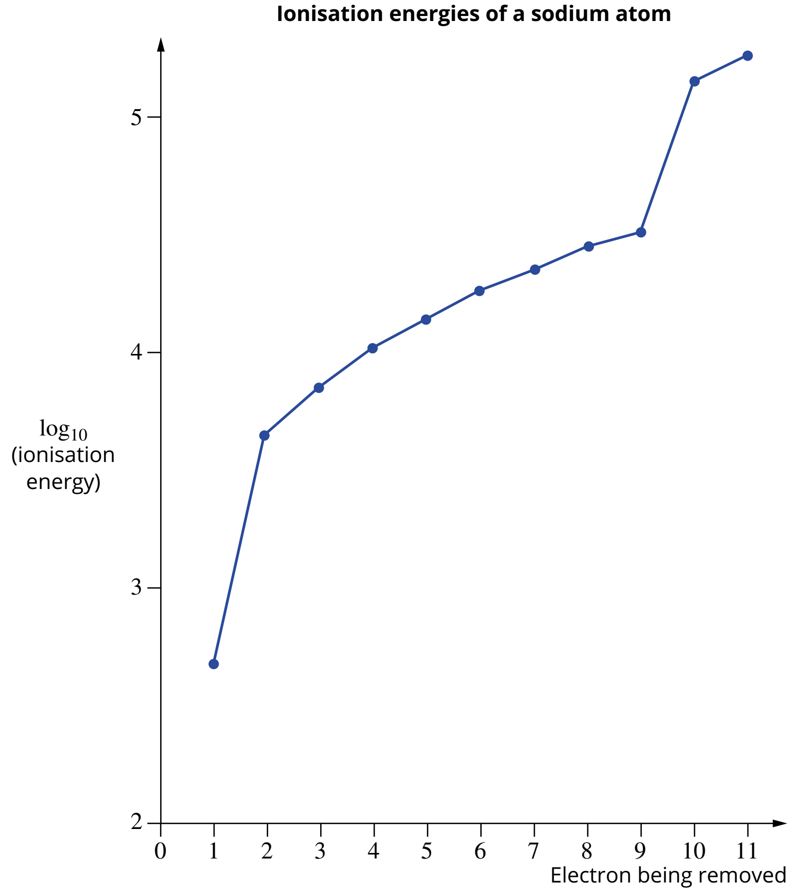


FIGURE 3.2.7 This graph of the ionisation energies of a sodium atom provides evidence for the shell model of electron configuration. (The logarithm of ionisation energy is used to provide a more convenient vertical scale.)

3.2 Review

SUMMARY

- All atoms are composed of a small, positively charged nucleus surrounded by a negatively charged cloud of electrons.
- The mass of an atom is mostly determined by the mass of the nucleus, while the size of an atom is determined by the cloud of electrons.
- The nucleus is made up of two subatomic particles called nucleons—protons and neutrons.
- Protons have a positive charge, electrons have a negative charge and neutrons have no charge.
- Protons and neutrons are similar in size and mass, while electrons are approximately 1800 times smaller.
- The Bohr model assumes that electrons can only exist in fixed, circular orbits of specific energies called energy levels or shells.
- Lower energy level shells are closer to the nucleus and other shells can be drawn like layers of an onion around the nucleus.
- The maximum number of electrons that will exist in a shell (n) is given by $2n^2$. This means that $n = 1$ shell, $n = 2$ shell and $n = 3$ shell could hold 2, 8 and 18 electrons respectively.
- The arrangement of electrons in shells is called an electron configuration and is shown as a series of numbers, e.g. 2,8,2.
- When an electron absorbs energy (e.g. heat or light), the electron can jump from one shell to a higher energy shell.
- When an electron falls from a higher energy shell to a lower energy shell, it emits energy in the form of light.
- The outer electron shell is called the valence shell.
- The ionisation energy (energy required to remove an electron) is greater for electrons in shells closer to the nucleus.

KEY QUESTIONS

Retrieval

1 State how many times larger an atom is compared to its nucleus.

2 Identify which subatomic particles make up most of the mass of an atom and where are they found.

3 Describe how electrons are held within the cloud surrounding the nucleus.

4 A sodium atom has 23 nucleons and an atomic number of 11. State how many neutrons this sodium atom will have.

5 State the number of electrons in each shell for the following atoms.

- **a** 5 electrons
- **b** 12 electrons
- **c** 20 electrons
- **d** 35 electrons

6 Write the name and symbol of the element with this electron configuration.

- **a** 2
- **b** 2,7
- **c** 2,8,3
- **d** 2,5
- **e** 2,8,7

7 Identify the elements that have the following electron arrangements.

- **a** two electrons in the first shell, six electrons in the second shell
- **b** two electrons in the first shell, eight electrons in the second shell and two electrons in the third shell
- **c** two electrons in the first shell, eight electrons in the second shell and eight electrons in the third shell
- **d** two electrons in the first shell and two electrons in the second shell

Comprehension

8 Draw the Bohr diagram for an atom of argon.

9 Determine how many valence electrons are in an atom with 18 electrons.

10 Determine the electron configuration of the following elements.

- **a** Be
- **b** S
- **c** Ar
- **d** Mg
- **e** Ga

3.2 Review *continued*

11 Determine the name and symbol of the element with this electron configuration.

- **a** 2
- **b** 2,7
- **c** 2,8,3
- **d** 2,5
- **e** 2,8,7

12 Describe the electron configuration of the following elements.

- **a** Ca
- **b** Sr
- **c** Na
- **d** Cl

Analysis

13 Identify two elements that have two electrons in their valence shell. Using the periodic table on page 10, describe another similarity between these elements.

14 Table 3.2.2 on page 42 shows the maximum number of electrons in the first four shells. Calculate the maximum number able to be held in the fifth and sixth shells.

15 An atom has two electrons in the first shell, six electrons in the second shell and eight electrons in the third shell. Identify the atom and state why this electron arrangement is unexpected. Propose a reason for this.

16 **a** Consider the ionisation energy data for aluminium shown below and represent this on a graph similar to the format used in Figure 3.2.7 on page 46.

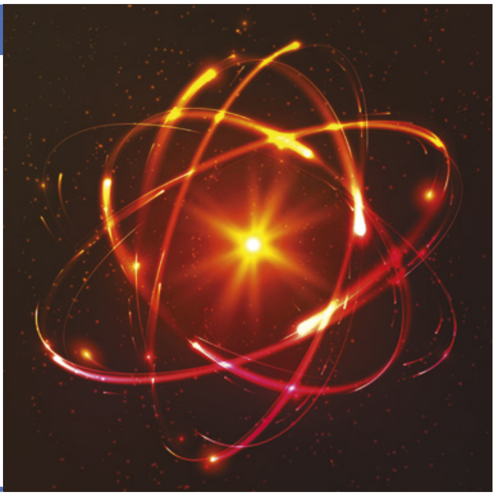
b Explain how this data reflects the electron configuration of aluminium.

Ionisation number	Ionisation energy (kJ mol^{-1})
1	577.5
2	1816.7
3	2744.8
4	11577

3.3 Using notation to describe atoms

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- write the nuclear symbol notation for elements using atomic number and mass number
- determine the number of subatomic particles in an atom from an element's atomic number and mass number
- recall isotopes are atoms of the same element that have different numbers of neutrons and can be represented in the form ^{A}X (IUPAC) or X-A
- recognise that isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties
- understand the origin of radioactivity and the usefulness of radioisotopes and some potential risks associated with their use.



An element is made up of one type of **atom**. The type of atom that makes up each element is determined by the number of protons in the nucleus, so, for example, an atom of carbon will always have six protons. In the unlikely event that a carbon atom loses a proton it is no longer a carbon atom—it has become a boron atom. However, this is not true for electrons and neutrons—sometimes these particles are lost or gained. An **isotope** is one such alternative form of an atom. To simplify the description of how many of each particle is present in an atom, scientists use nuclear symbol notation.

NUCLEAR SYMBOL NOTATION

The number of protons in an atom's nucleus is known as the **atomic number** and is represented by the symbol Z. All atoms that belong to the same element must have the same number of protons and therefore have the same atomic number, Z. For example, all hydrogen atoms have $Z = 1$, all carbon atoms have $Z = 6$ and all gold atoms have $Z = 79$.

The number of protons and neutrons in the nucleus is known as the **mass number** and is represented by the symbol A. The mass number is obtained by adding the number of protons to the number of neutrons. The periodic table provides the atomic mass of each element. However, the atomic mass shown on the periodic table is rarely an integer (whole number). When you use this information to obtain the mass number you round to the nearest whole number. This is due to the existence of isotopes, which is described on page 51.

As all atoms are electrically neutral, the number of protons in an atom is equal to the number of electrons in the atom. The atomic number therefore tells you both the number of protons and the number of electrons. For example, carbon atoms, with $Z = 6$, have six protons and six electrons.

ℹ In a neutral atom, number of protons = number of electrons.

Similarly, by definition, if the atomic and mass numbers are known the number of neutrons can be calculated. The number of neutrons is equal to the mass number minus the atomic number.

ℹ Number of neutrons = mass number − atomic number

The number of protons, neutrons and electrons defines the basic structure of an atom. **Nuclear symbol notation** represents an atom using its atomic and mass numbers as shown in Figure 3.3.1. This notation has the mass number as a superscript and the atomic number as a subscript.

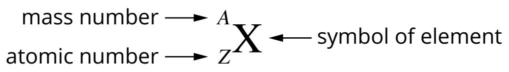


FIGURE 3.3.1 Nuclear symbol notation showing atomic number and mass number of an atom

$$^{27}_{13}\mathrm{Al}$$

FIGURE 3.3.2 This representation of an aluminium atom indicates the number of protons, neutrons and electrons in the atom.

For an aluminium atom, this would be written as shown in Figure 3.3.2.

From this notation, you can determine that:

- aluminium has 13 protons (13)
- aluminium has 14 neutrons (mass number minus the atomic number, $A - Z$)
- aluminium has 13 electrons because atoms have no overall charge, and the number of electrons must equal the number of protons.

Worked example 3.3.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this atomic symbol:

$$^{40}_{18}\mathrm{Ar}$$

Thinking	**Working**
The atomic number is equal to the number of protons.	the number of protons = Z = 18
Find the number of neutrons. Number of neutrons = mass number – atomic number	the number of neutrons = $A - Z$ = 40 – 18 = 22
Find the number of electrons. The number of electrons is equal to the atomic number because the total negative charge is equal to the total positive charge.	number of electrons = Z = 18

➤ Try yourself 3.3.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this atomic symbol:

$$^{238}_{92}\mathrm{U}$$

carbon
6
C
12.011

FIGURE 3.3.3 Presentation of atomic number and atomic mass on some periodic tables

Some periodic tables represent the atomic number and mass number differently. The atomic number and mass number are two key pieces of information provided in the periodic table square for each element. However, some periodic tables place the atomic number on the *top* of the **chemical symbol** and the mass number on the *bottom* of the square, as shown for carbon in Figure 3.3.3. The periodic table in the inside cover at the back of the book uses this format. Nuclear symbol notation places the mass number on the *top* and the atomic number on the *bottom*.

ISOTOPES

All atoms that belong to the same element have the same number of protons in the nucleus, and therefore the same atomic number, Z. However, not all atoms that belong to the same element have the same mass number, A. For example, hydrogen atoms can have mass numbers of 1, 2 or 3. In other words, hydrogen atoms may contain just a single proton, a proton and a neutron, or a proton and two neutrons as shown in Figure 3.3.4. Atoms that have the same number of protons (atomic number) but different numbers of neutrons (and therefore different mass numbers) are known as **isotopes**. In nature, different elements have different numbers of isotopes. Gold has only one isotope, whereas lead has four and mercury has seven isotopes.

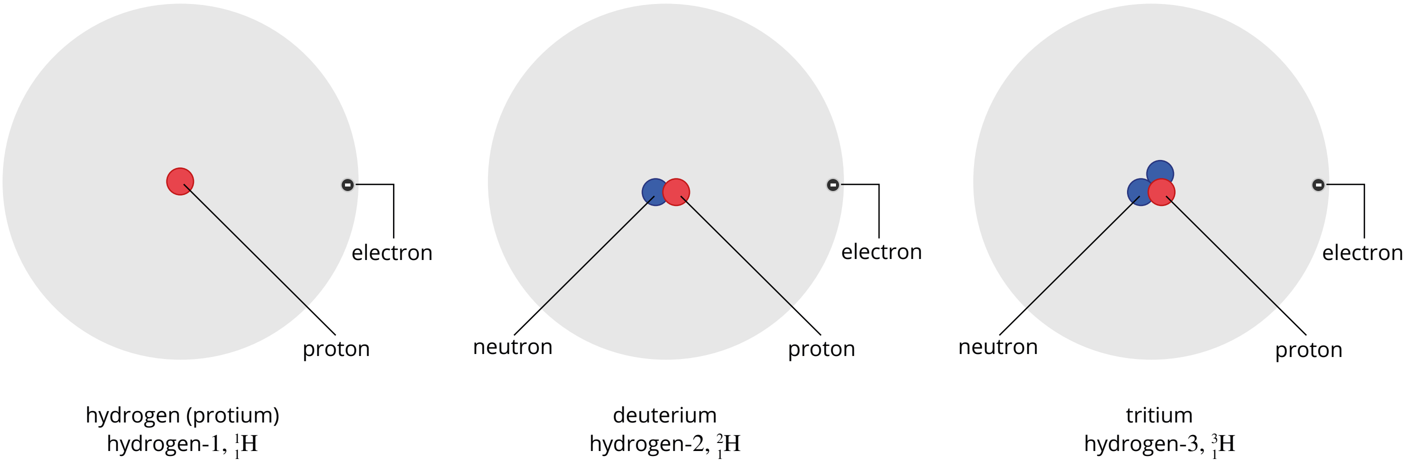


FIGURE 3.3.4 The three isotopes of hydrogen are given special names. A hydrogen atom with just one proton in its nucleus is known as hydrogen or protium. A hydrogen atom with one proton and one neutron is known as deuterium. A hydrogen atom with one proton and two neutrons is known as tritium.

An isotope will have an identical electron configuration as the element and so its **chemical properties** such as bonding and chemical reactions are the same. However, the change to the number of neutrons will alter the mass, density and **physical properties** related to mass or density.

The existence of isotopes creates a problem when identifying the atomic mass of an element because each isotope has a different mass. The accepted atomic mass for an element is a **weighted average** of all known isotopes. This is the reason that the atomic masses shown on the periodic table are not whole numbers even though you know atoms have whole numbers of protons and neutrons (each with a mass of 1). Chapter 5 describes this further and explains calculations involving determination of atomic mass.

Carbon has three naturally occurring isotopes, shown in Figure 3.3.5. In the 1950s and 1960s, nuclear weapons testing caused a spike in the abundance of the carbon-14 isotope in the atmosphere. This has been declining in the last 50 years.

$^{12}_{6}C$ carbon-12

$^{13}_{6}C$ carbon-13

$^{14}_{6}C$ carbon-14

FIGURE 3.3.5 Isotopes of carbon

Nuclear symbol notation is used to name radioisotopes because this clearly shows the number of protons and neutrons. However, it is common to only use the element name and mass number. For example, $^{14}_{6}C$ can be written as ^{14}C or carbon-14. An important use of this isotope is to determine the age of biological specimens. This technique uses the change in the nucleus of the **radioactive** isotope over time to estimate how long ago the organism lived.

Carbon-14 and the war against poachers

Isotopes are important in the war against ivory poaching in Africa. All living organisms take up carbon-14 and carbon-12 from the environment to build biologically important structures in the organism. The biochemistry of proteins, DNA, carbohydrates and lipids is based on this crucial element. Over time, the amount of carbon-14 in tissues of plants and animals decreases as the isotope radioactively decays. The levels of carbon-12 remain constant; so, the change to the ratio of carbon-14 to carbon-12 (^{14}C:^{12}C) decreases with time. By looking at the levels of the carbon-14 isotope in elephant tusks and ivory, scientists can determine how old the ivory is.

If the authorities know the age of the ivory, then legal action can be taken against the poachers and sellers of the product. Ivory products prior to 1989 are allowed to be traded. Since 1989, the trade on ivory has been made illegal worldwide.

Radioisotopes

The role of chemists who monitor carbon-14 in criminal investigations of poachers is one of many examples of the usefulness of these special isotopes. These are called **radioisotopes**. The forces holding the nucleus of an isotope together are unbalanced and over time particles and electromagnetic radiation are emitted from the nucleus. This is measured as **radioactivity**, and was first described by Henri Becquerel, Pierre Curie and Marie Curie in 1903. The three key particles and radiation emitted (Table 3.3.1) have provided the basis for valuable technologies but also revealed health risks to humans. After discovering and characterising the radioactive elements radium and polonium, Marie Curie eventually succumbed to the effects of radiation exposure. Even her famous notebooks are too radioactive to be viewed without protection.

TABLE 3.3.1 Radiation types showing the three types of particles and radiation emitted from radioactive substances and their penetrating ability. Note that alpha radiation is a helium nucleus, beta radiation is electrons and gamma radiation is electromagnetic radiation.

	Symbol	Identity	Penetrating ability	Examples of sources
alpha	α	helium nucleus $^{4}_{2}He$	low	polonium-210 radon-222
beta	β	electron	moderate	carbon-14 sulfur-35
gamma	γ	electromagnetic radiation	high	iodine-125 caesium-137

Radioactive particles and radiation can traverse matter and interact with biological molecules. The most vulnerable biological structure is the DNA held within the nucleus of every cell. Even a change to a single molecule can alter the way the DNA is transcribed or copied. At the protein level this may mean that the cell has uncontrolled growth or does not respond to normal physiological regulation. Some tumours develop in this way. Therefore, safe use of isotopes requires shielding equipment, such as lead panels, to protect the user.

The penetrating ability of radiation emitted by radioisotopes is used in quality control processes to measure the thicknesses of materials. In smoke detectors, americium-241 (an α emitter) ionises and detects any smoke particles in the air.

Figure 3.3.6 shows the penetrating capabilities of **alpha particles**, and **beta** and **gamma radiation**. This characteristic is used in medical imaging as shown in Figure 3.3.7 which shows **positron emission tomography (PET)** of iodine-131.

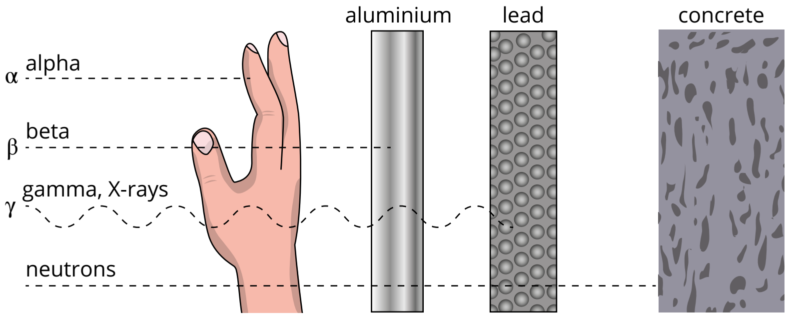


FIGURE 3.3.6 Diagrammatic representation of the penetrating ability of alpha particles, and beta and gamma radiation

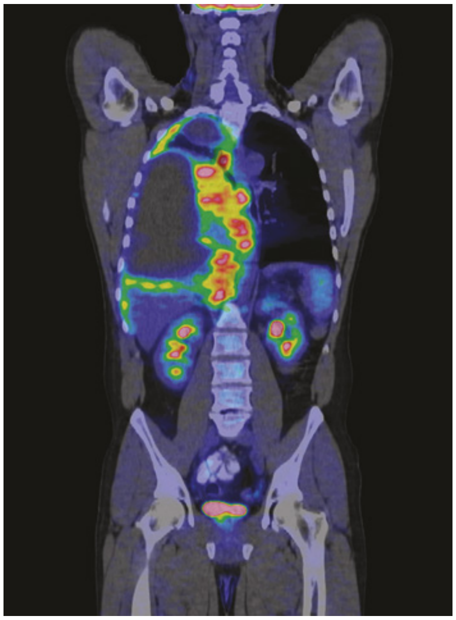


FIGURE 3.3.7 A composite image of a computed tomography (CT) scan and a coloured positron emission tomography (PET) scan of the thorax and abdomen. Localisation of the medical radioisotope indicates rapidly dividing cells.

Radioisotopes have a wide variety of uses, including carbon-14 for carbon dating in geology and palaeobiology; radioactive tracers such as iodine-131 in nuclear medicine; radioimmunoassays for testing constituents of blood, serum, urine, hormones and antigens; and radiotherapy that destroys damaged cells. Use of radioisotopes requires careful evaluation and monitoring because of the potential harmful effects to humans and/or the environment if their production, use and disposal are not managed effectively. Risks include unwanted damage to cells in the body, especially during pregnancy, and ongoing radiation produced from radioactive sources with long half-lives.

Half-life

The **half-life** of a radioisotope is the time taken for half the atoms in the sample to decay to a new element or isotope. This can be an extremely short time or millions of years. Figure 3.3.8 shows the change to the number of radioactive particles released from a sample when the half-life of the particle is two days. By determining the rate of decay (gradient of the graph) the age of the specimen can be determined.

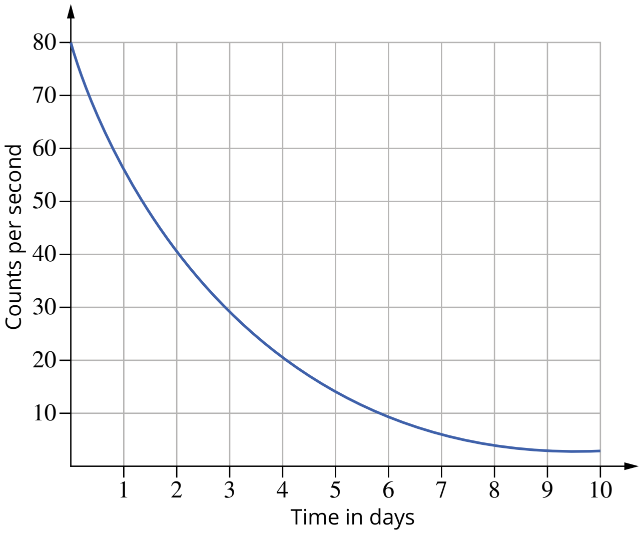


FIGURE 3.3.8 Change to the emission of radiation from a radioisotope over time

Radioisotopes: nuclear medicine

It is estimated that one-third to one-half of Australians will use radioisotopes and nuclear medicine for diagnosis or treatment of disease in their lifetime.

The use of these non-invasive chemicals in a way that targets a cell type or organ is increasing. Some are shown in Table 3.3.2. The usefulness of these isotopes for scanning and diagnostic purposes relies on the particles produced being lower energy β or γ radiation, and the isotope having a short half-life. The 'tracer' is often attached to a substance that is preferentially absorbed or used by the organ of interest. Conversely, radioisotopes used for radionucleotide therapy (RNT) should have a longer half-life, higher energy and must be targeted to cells that need to be destroyed such as tumour cells.

TABLE 3.3.2 Some common medical radioisotopes, their radiation emission type and half-life

Radioisotope	Radiation type	Half-life	Scan type
^{11}C	positron	20.4 minutes	PET—Alzheimer's, schizophrenia, depression
^{18}F	positron	110 minutes	PET—various brain and body scans
^{52}Fe	positron	8.3 hours	bone marrow scan
^{99}Tc	gamma	6 hours	bone, heart, brain scans
^{131}I	beta	8 days	thyroid imaging and radiotherapy
^{133}Xe	beta	5.3 days	lung function scan
^{177}Lu attached to a peptide	beta	160 days	radiotherapy—prostate cancer
^{192}Ir implant	gamma	94 days	radiotherapy—breast cancer
^{213}Bi attached to a specific antibody	alpha	45 minutes	radiotherapy—leukaemia, glioma, melanoma

Australia's only nuclear reactor at Lucas Heights in south-west Sydney is run by the Australian Nuclear Science Technology Organisation. It has increased the number of medical doses of radioisotopes produced from 550 000 in 2015 to many millions of doses to meet national and international demand. However, due to the very short half-life of some useful radioisotopes, these must be made in a cyclotron at a PET scanning facility (Figure 3.3.9). Only a few large hospitals in Australia have a cyclotron in their nuclear medicine department.

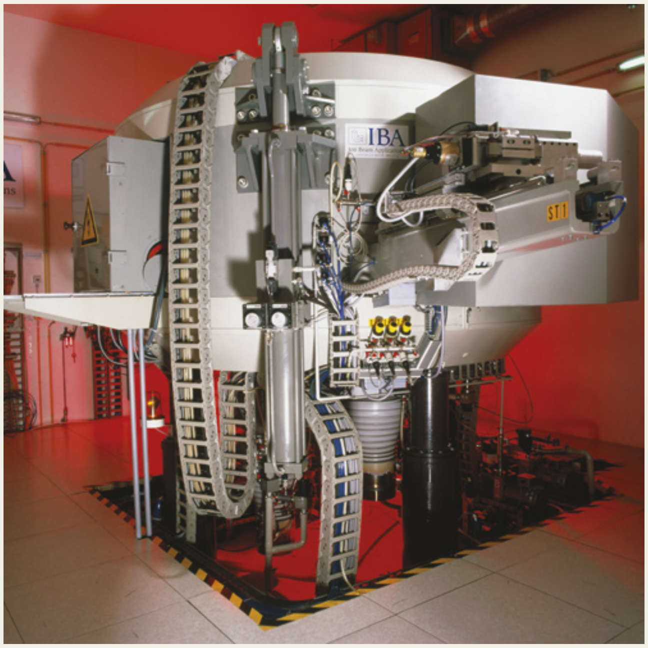


FIGURE 3.3.9 A cyclotron machine. The cyclotron is used to radioactively label chemicals for use in medicine. This is done by accelerating high energy particles and passing them through samples of the chemicals to be made radioactive.

The radioactive properties required determine whether a nuclear reactor or a cyclotron is used to produce a radioisotope.

- Atoms with extra protons in the nucleus are called neutron-deficient and are produced in a particle accelerator such as a cyclotron.
- Atoms with extra neutrons in the nucleus are called neutron-rich and are produced in a nuclear reactor.

Radionucleotides are potentially harmful to humans. As ionising radiation penetrates living matter, it gives up its energy through random interactions with atoms and molecules in its path, leading to the formation of reactive ions and free radicals. These biochemical changes can decrease the viability of cells or, more importantly, cause changes to DNA known as mutations. A change in the way a cell's growth is regulated may result in a cancer cell. Treatments involving high-energy radiation, and those where the radionucleotide accumulates in a particular body tissue, are situations where cell damage can be expected. In the case of breast cancer treatment using ^{192}Ir (iridium), targeted cell death occurs when the radioisotope is implanted next to tumour cells.

In practice, a variety of shielding structures are used, depending on the penetrating ability of the radiation (Figure 3.3.6 on page 53). These may be as easy to use as a sheet of perspex, or as cumbersome and heavy as a lead apron. Generally, scanning machines are automated to allow technicians to retire to an adjacent room or behind a barrier during the scan. It is vital that all recommended precautions are used, including daily monitoring of all surfaces used with a Geiger counter. Personal radiation badges and monitors are checked daily or weekly. These have traditionally contained film, but thermoluminescent dosimeters (TLDs), aluminium oxide based dosimeters, and an electronic personal dosimeter (EPD) are increasingly being used. The safe use, storage and disposal of radioisotopes is outlined in the Code of Practice of the Australian Radiation Protection and Nuclear Safety Agency (ARPNSA). These procedures are put in place to protect patients and hospital staff.

Radionucleotides must be carefully handled to avoid contamination. Scientists, doctors and nurses must wear devices to monitor levels of radiation in their workplace. Three ways that this is done are shown in Figure 3.3.10.

Review

1 Identify two characteristics of a radioisotope suitable for medical imaging.

2 Research a common radioisotope from Table 3.3.2 on page 54 and describe how it acts to provide information or treat a disease.

3 A nuclear medicine scan that medical practitioners use in their diagnosis and treatment of disease is available within days in most large cities in Australia. Discuss the advantages and disadvantages of the increasing use of this technology.

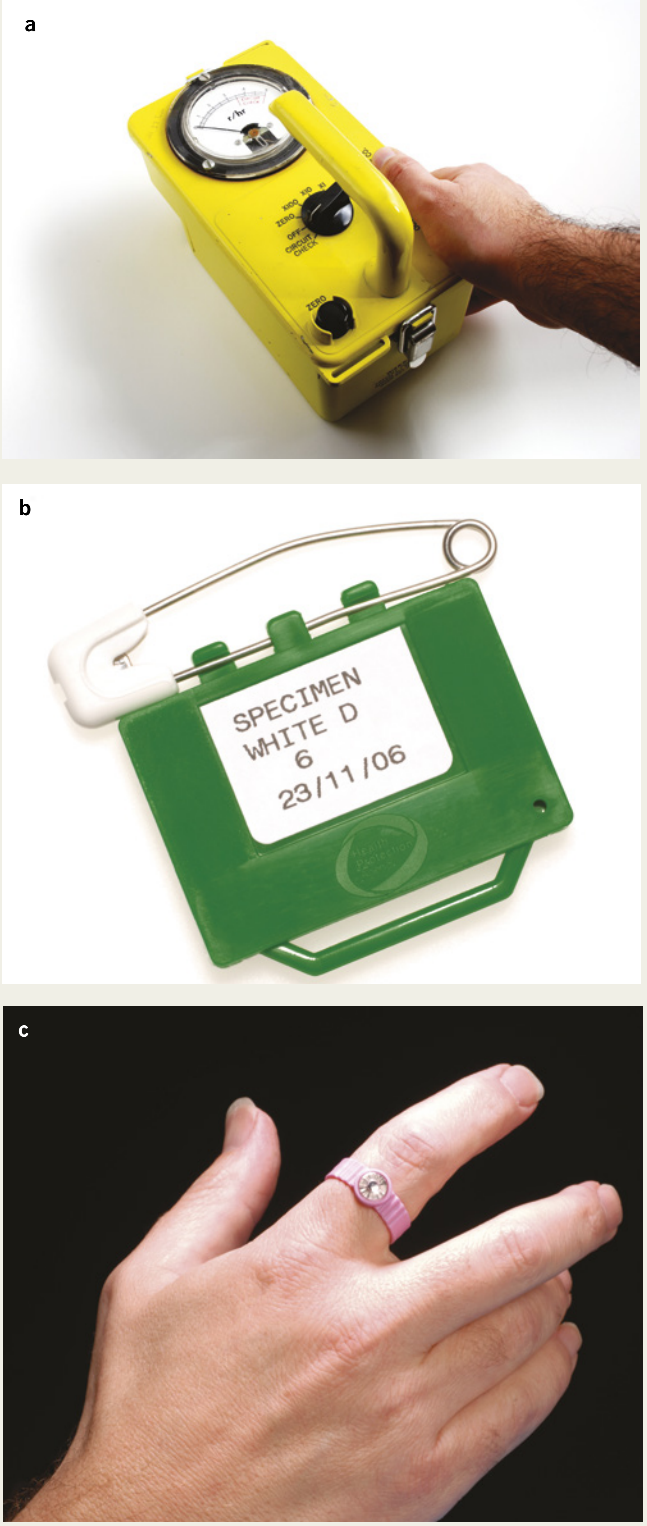


FIGURE 3.3.10 Monitoring radiation in a nuclear medicine department takes many forms: (a) a Geiger counter (b) a case film badge and (c) a radiation detector.

3.3 Review

SUMMARY

- Nuclear symbol notation is a way to descriptively write the symbol for an atom using the atomic number and mass number.
- The number of subatomic particles in an atom can be determined from an element's atomic number and mass number.
- The atomic number indicates how many protons or electrons an atom has.
- The mass number tells you how many protons and neutrons are in the nucleus of the atom. When the atomic number is subtracted from the mass number the number of neutrons is obtained.

  mass number (*A*) minus atomic number (*Z*)
- Isotopes are atoms with the same atomic number but different mass numbers, i.e. they have the same number of protons but different numbers of neutrons.
- Isotopes have the same chemical properties but different physical properties such as mass, density and radioactivity.
- The penetrating ability of radiation from unstable isotope nuclei has uses in our society but also poses some risk.

KEY QUESTIONS

Retrieval

1 State the numbers of protons, neutrons and electrons in the atom $^{31}_{15}P$.

2 Identify the correct name for an atom with seven protons and seven neutrons. Write its nuclear symbol notation.

3 Write the nuclear symbol notation of the following elements.
  - **a** potassium
  - **b** nickel
  - **c** boron
  - **d** lithium

4 Identify these atoms and state the number of neutrons each contains.
  - **a** $^{27}_{13}X$
  - **b** $^{19}_{9}X$
  - **c** $^{55}_{25}X$
  - **d** $^{4}_{2}X$

Comprehension

5 Oxygen has two known isotopes. One has nine neutrons and the other has 10 neutrons. Write their symbols using atomic symbol notation.

6 Explain the similarities and differences between isotopes of an element.

7 Use the periodic table in the inside cover at the back of the book to, complete the following table.

Element	**Nuclear symbol notation**	**Atomic number**	**Mass number**	**Number of protons**	**Number of neutrons**	**Number of electrons**
hydrogen						
hydrogen	$^{2}_{1}H$					
fluorine		9	19			
sulfur				16	17	
		16	34			
					36	29
	$^{16}_{8}O$					
				29	34	

Analysis

8 Determine any isotopes included in the table in Question **7** and explain your reasoning.

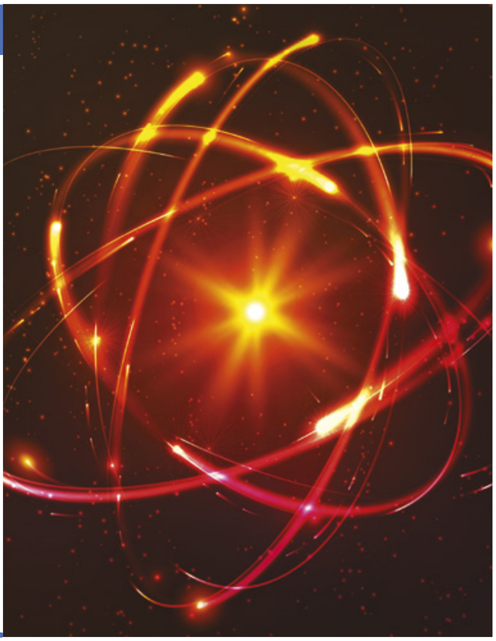
9 Determine the *Z* number for the following elements using the atomic masses shown in the periodic table in the inside cover at the back of the book. Compare and explain the answers obtained for parts **b** and **d**.
  - **a** zirconium
  - **b** calcium
  - **c** phosphorus
  - **d** argon

10 Different elements sometimes have the same number of neutrons but never the same number of protons. Identify any example(s) of this in the first 10 elements and use these as examples in your answer.

3.4 Electronic structure of the atom

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that the quantum mechanical model of the atom describes electrons residing in shells, subshells and orbitals
- understand that the location of electrons within atoms in their ground state can be represented using electron configuration indicating the shell, subshell and number of electrons
- understand the relative energies of the *s*, *p* and *d* subshells and the required order of filling of shells and subshells
- understand that the order of filling of subshells is different to the order of writing subshells in an electron configuration
- write electron configurations for atoms and ions up to $Z = 36$ using both full and condensed notation
- use Hund's rule and the Pauli exclusion principle to construct orbital diagrams to represent character and relative energy of orbitals for elements up to $Z = 36$
- recognise the electron configuration of Cr and Cu as exceptions.



The shell model of the atom described in Module 3.3 shows electrons in particular energy levels (shells) in an atom. Electronic structure is more complicated than this. The model that explains all experimental findings to date describes that electron shells contain **subshells** and **orbitals**. When scientists including Erwin Schrödinger adapted Bohr's model and used mathematics to describe the atom, the model became known as the **quantum mechanical model**. Interestingly, his mathematical modelling was only correct when the electron was considered to be a wave, not a particle. Scientists began to think about electrons in an entirely different way, suggesting that the physics inside atoms might be very different from the physics experienced in everyday life.

SHELLS, SUBSHELLS AND ORBITALS

The quantum mechanical model describes shells and subshells in the following way.

- There are major energy levels in an atom that are called shells. The shell number (n) is numbered 1, 2, 3, 4, 5, 6 or 7.
- The maximum number of electrons in a shell is $2n^2$.
- These shells contain separate energy levels of similar energy, called subshells, which are labelled *s*, *p*, *d* and *f*. Each subshell can only hold a certain number of electrons.
- *s*-subshells hold *two* electrons, *p*-subshells hold *six* electrons, *d*-subshells hold *10* electrons, *f*-subshells hold *14* electrons.

Electrons therefore exist within subshells, which themselves are within shells. The subshells for the first three shells are summarised in Table 3.4.1. It shows for shell 1 ($n = 1$) there is only an *s*-subshell holding a maximum of two electrons. However, in the second shell ($n = 2$) there is an *s*-subshell *and* a *p*-subshell. These subshells hold two and six electrons respectively, and so the *total* electron number in the shell is 8. Note that the total number of electrons is still $2n^2$.

TABLE 3.4.1 Subshells for the first three shells

Shell number (n)	Subshells present	Electrons in subshell	Name of subshell	Total electrons in shell
1	*s*	2	1*s*	2
2	*s* *p*	2 6	2*s* 2*p*	8
3	*s* *p* *d*	2 6 10	3*s* 3*p* 3*d*	18

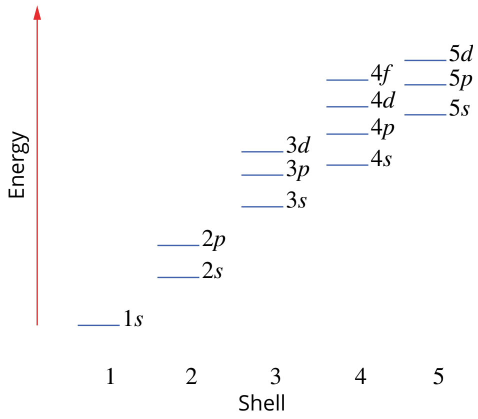


FIGURE 3.4.1 Subshells described in the quantum mechanical model showing energy levels

The name of each subshell tells us two things—the energy level and the type of subshell. For example, a '3*s*' subshell is in the third energy level ($n = 3$), and will hold two electrons.

A useful way to view the quantum mechanical model is to view all subshells relative to their energy level. Remember, the inner shells and subshells are lower energy. Figure 3.4.1 shows that the outer shells contain more subshells and are higher energy than the inner shells. For example, $n = 4$ shell contains *s*, *p*, *d* and *f* subshells, whereas $n = 2$ shell only contains an *s*- and a *p*-subshell. (Note: $n = 5$ shell on this diagram is not complete.)

The order of 'filling' electrons into these shells and subshells must be from lowest energy ($n = 1$) first, as described by the **Aufbau principle**. The order is (reading from the bottom of Figure 3.4.1) 1*s* 2*s* 2*p* 3*s* 3*p* etc. Using Table 3.4.1 or Figure 3.4.1, the number of electrons in each subshell can be determined for a given atom.

Worked example 3.4.1

ELECTRON CONFIGURATION USING SUBSHELLS

Apply the rules of the quantum mechanical model to determine the number of electrons in each subshell of an atom with 10 electrons.

Thinking	Working
Recall the maximum number of electrons that each *s*, *p*, *d* or *f* subshell can hold.	subshell — maximum number of electrons *s* — 2 *p* — 6 *d* — 10 *f* — 14
Recall the order of filling of subshells.	1*s* 2*s* 2*p* 3*s* 3*p* etc.
Start at the first subshell and place the correct number of electrons in the subshells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.	subshell — electrons 1*s* — 2 2*s* — 2 2*p* — 6
Write the number of electrons in each subshell.	The neon atom has 10 electrons. It has two electrons in the 1*s*-subshell, two electrons in the 2*s*-subshell and six electrons in the 2*p*-subshell.

➤ Try yourself 3.4.1

ELECTRON CONFIGURATION USING SUBSHELLS

Apply the rules of the quantum mechanical model to determine the number of electrons in each subshell of an atom with 13 electrons.

ELECTRON SUBSHELL CONFIGURATION FOR ATOMS WITH MORE THAN 18 ELECTRONS

Look closely at the section of Figure 3.4.1 where the subshells in the third ($n = 3$) and the fourth ($n = 4$) shells are shown. The energy level of 4*s* is *below* that of 3*d*. This means that a subshell of the fourth shell will fill before the third shell is completely full. Also, subshells in the fifth energy level will fill before the fourth shell is complete. This means that the quantum mechanical order of filling does not simply complete each shell in turn, but fills part of two shells at the same time.

The order of filling of subshells is
1*s* > 2*s* > 2*p* > 3*s* > 3*p* > 4*s* > 3*d* > 4*p* > 5*s* > 4*d* > 5*p*

So, to add electrons to atoms with more than 20 electrons, you start with the shell and subshell that has the lowest energy. The first two electrons fill the 1*s*-subshell in the first shell. The next eight electrons fill the *s*- and *p*-subshells in the second shell. The next eight electrons fill the *s*- and *p*-subshells in the third shell. The 4*s*-subshell is then filled with two electrons because this subshell is lower in energy than the 3*d*-subshells. Once the 4*s*-subshell is filled, the next 10 electrons are placed in the 3*d*-subshells, followed by the 4*p*-subshell. Table 3.4.2 shows the order of filling of subshells for the first 36 elements and the geometric pattern shown in Figure 3.4.2 is a commonly used and convenient way of remembering the order in which the subshells are filled.

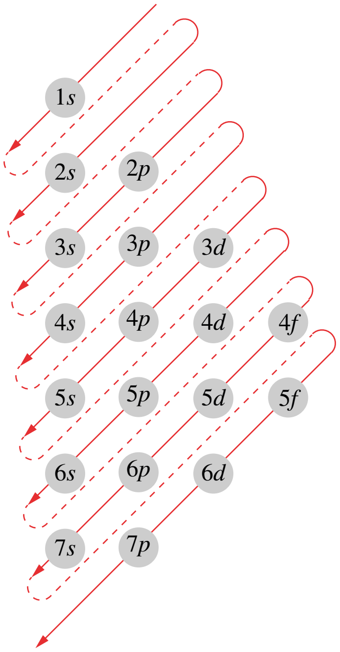


FIGURE 3.4.2 This geometric pattern shows the order in which the subshells are filled. Note that in this diagram the fourth shell starts filling before the third shell is completely filled. This is because the 4*s*-subshell is lower in energy than the 3*d*-subshell. As a result, the 4*s*-subshell accepts two electrons after the 3*s*- and 3*p*-subshells are filled but before the 3*d*-subshell begins filling.

TABLE 3.4.2 Order of filling of subshells of the first 36 elements

Z	Element	Order of filling subshells
1	hydrogen	1*s*
2	helium	1*s*
3	lithium	1*s*2*s*
4	beryllium	1*s*2*s*
5	boron	1*s*2*s*2*p*
6	carbon	1*s*2*s*2*p*
7	nitrogen	1*s*2*s*2*p*
8	oxygen	1*s*2*s*2*p*
9	fluorine	1*s*2*s*2*p*
10	neon	1*s*2*s*2*p*
11	sodium	1*s*2*s*2*p*3*s*
12	magnesium	1*s*2*s*2*p*3*s*
13	aluminium	1*s*2*s*2*p*3*s*3*p*
14	silicon	1*s*2*s*2*p*3*s*3*p*
15	phosphorus	1*s*2*s*2*p*3*s*3*p*
16	sulfur	1*s*2*s*2*p*3*s*3*p*
17	chlorine	1*s*2*s*2*p*3*s*3*p*
18	argon	1*s*2*s*2*p*3*s*3*p*
19	potassium	1*s*2*s*2*p*3*s*3*p*4*s*
20	calcium	1*s*2*s*2*p*3*s*3*p*4*s*
21	scandium	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
22	titanium	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
23	vanadium	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
24	chromium*	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
25	manganese	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
26	iron	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
27	cobalt	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
28	nickel	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
29	copper*	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
30	zinc	1*s*2*s*2*p*3*s*3*p*4*s*3*d*
31	gallium	1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*
32	germanium	1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*
33	arsenic	1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*
34	selenium	1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*
35	bromine	1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*
36	krypton	1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*

* Chromium and copper are exceptions to the rule and are explained on page 64.

WRITING ELECTRON CONFIGURATIONS USING SUBSHELLS

The electron configurations for the quantum mechanical model are similar to the electron configurations of the shell model. However, shell, subshell and number of electrons must be shown. The electron configuration of a sodium atom is shown in Figure 3.4.3.

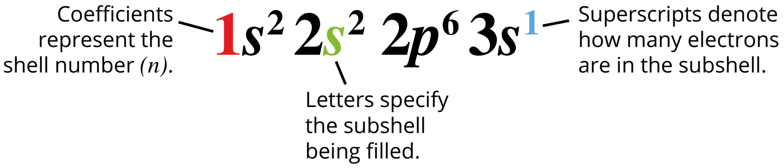


FIGURE 3.4.3 The electron configuration of a sodium atom. It shows that there are two electrons in the 1*s*-subshell, two electrons in the 2*s*-subshell, six electrons in the 2*p*-subshell and one electron in the 3*s*-subshell. Note that the total number of electrons is obtained by adding the superscripts.

Worked example 3.4.2

WRITING ELECTRON CONFIGURATIONS USING THE QUANTUM MECHANICAL MODEL

Apply the rules of the quantum mechanical model to write the electron configuration of a manganese atom with 25 electrons.	
Thinking	**Working**
Recall the maximum number of electrons that each *s*, *p*, *d* or *f* subshell can hold.	subshell — maximum number of electrons *s* — 2 *p* — 6 *d* — 10 *f* — 14
Recall the order of filling of subshells.	1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p* 5*s* etc.
Start at the first subshell and place all 25 electrons, putting the correct number of electrons in the subshells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed per subshell.	subshell — electrons 1*s* — 2 2*s* — 2 2*p* — 6 3*s* — 2 3*p* — 6 4*s* — 2 3*d* — 5
Write the electron configuration of manganese by placing the electron number as a superscript to each subshell.	$1s^22s^22p^63s^23p^64s^23d^5$
Rewrite the electron configuration, grouping subshells from the same shell together.	$1s^22s^22p^63s^23p^63d^54s^2$

➤ Try yourself 3.4.2

WRITING ELECTRON CONFIGURATIONS USING THE QUANTUM MECHANICAL MODEL

Apply the rules of the quantum mechanical model to write the electron configuration of a vanadium atom with 23 electrons.

Krypton has 36 electrons. According to the quantum mechanical order of subshell filling, its configuration is:

$$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$$

Although the 4*s*-subshell is filled before the 3*d*-subshell, the subshells of the third shell are usually grouped together. Therefore, the electron configuration for a krypton atom is written as:

$$1s^22s^22p^63s^23p^63d^{10}4s^24p^6$$

This means that after you have used the order of filling and written the electron configuration, it needs to be rewritten so that the subshells with the same *n* value are grouped together as shown in Figure 3.4.4.

$$\underbrace{1s^2}\underbrace{2s^22p^6}\underbrace{3s^23p^63d^{10}}\underbrace{4s^24p^6}$$

FIGURE 3.4.4 Electron configuration of krypton showing the grouping of subshells together

Worked example 3.4.3

WRITING ELECTRON CONFIGURATIONS USING THE QUANTUM MECHANICAL MODEL

Apply the rules of the quantum mechanical model to write the electron configuration of a strontium atom with 38 electrons.	
Thinking	**Working**
Recall the maximum number of electrons that each *s*, *p*, *d* or *f* subshell can hold.	subshell — maximum number of electrons *s* — 2 *p* — 6 *d* — 10 *f* — 14
Recall the order of filling of subshells.	1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p* 5*s* etc.
Start at the first subshell and place all 38 electrons, putting the correct number of electrons in the subshells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed per subshell.	subshell — electrons 1*s* — 2 2*s* — 2 2*p* — 6 3*s* — 2 3*p* — 6 4*s* — 2 3*d* — 10 4*p* — 6 5*s* — 2
Write the electron configuration of strontium by placing the electron number as a superscript to each subshell.	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$
Rewrite the configuration, grouping subshells from the same shell together.	$1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^2$

➤ Try yourself 3.4.3

WS 1.1.2

WRITING ELECTRON CONFIGURATIONS USING THE QUANTUM MECHANICAL MODEL

Apply the rules of the quantum mechanical model to write the electron configuration of an arsenic atom with 33 electrons.

Condensed electron configuration

Condensed electron configuration is also called **noble gas notation**. When writing the electron configuration of an element, the symbol of the noble gas (with square brackets) in the period before the element is used. This means that the actual numbers for the inner shells are not written—they are implied by the use of the noble gas symbol.

This a meaningful abbreviation of the full electron configuration notation that is especially useful for elements with multiple shells containing electrons. So, for example, the condensed electron configuration of phosphorus (2,8,5) is [Ne]5. The [Ne] symbol signifies 2,8 because this is the electron configuration of neon. Some other examples of condensed electron configuration are shown in Table 3.4.3.

Both ways of writing electron configurations are correct. However, an advantage of the condensed notation is that the number of valence electrons is clear. Use of this number is fundamental to describing how the atom bonds with other atoms.

TABLE 3.4.3 Some examples of the use of condensed electron configuration

Element	Atomic number	Electron configuration	Condensed electron configuration
Si	14	2,8,4	[Ne]$3s^2 3p^2$
Zn	30	2,8,18,2	[Ar]$3d^{10} 4s^2$
Br	35	2,8,18,7	[Ar]$3d^{10} 4s^2 4p^5$

Worked example 3.4.4

WRITING CONDENSED ELECTRON CONFIGURATIONS FOR UP TO 36 ELECTRONS

Apply the rules of the shell model to write the condensed electron configuration of silicon (14 electrons).

Thinking	Working
Recall the maximum number of electrons that each shell can hold.	shell (*n*) maximum number of electrons 1 2 2 8 3 18 4 32
Place the 14 electrons in the shells from the lowest energy (*n* = 1) to the highest energy. Do not exceed the maximum number of electrons allowed in each shell.	shell (*n*) electrons in atom 1 2 2 8 3 4 4
Write the electron configuration by listing the number of electrons in each shell separated by commas.	The configuration is 2,8,4.
Identify the noble gas in the period before silicon (neon) and write this in square brackets. Write the number of valence electrons.	[Ne]4

➤ Try yourself 3.4.4

WRITING CONDENSED ELECTRON CONFIGURATIONS FOR UP TO 36 ELECTRONS

Apply the rules of the subshell model to write the condensed electron configuration of calcium (20 electrons).

Writing electron configurations using condensed notation shows clearly the number of valence electrons, as shown in Table 3.4.4.

TABLE 3.4.4 The electron configuration of the first 36 elements using condensed notation

Z	Element		Configuration	Z	Element		Configuration
1	H	hydrogen	$1s^1$	19	K	potassium	$[Ar]4s^1$
2	He	helium	$1s^2$	20	Ca	calcium	$[Ar]4s^2$
3	Li	lithium	$[He]2s^1$	21	Sc	scandium	$[Ar]3d^14s^2$
4	Be	beryllium	$[He]2s^2$	22	Ti	titanium	$[Ar]3d^24s^2$
5	B	boron	$[He]2s^22p^1$	23	V	vanadium	$[Ar]3d^34s^2$
6	C	carbon	$[He]2s^22p^2$	24	Cr	chromium	$[Ar]3d^54s^1$
7	N	nitrogen	$[He]2s^22p^3$	25	Mn	manganese	$[Ar]3d^54s^2$
8	O	oxygen	$[He]2s^22p^4$	26	Fe	iron	$[Ar]3d^64s^2$
9	F	fluorine	$[He]2s^22p^5$	27	Co	cobalt	$[Ar]3d^74s^2$
10	Ne	neon	$[He]2s^22p^6$	28	Ni	nickel	$[Ar]3d^84s^2$
11	Na	sodium	$[Ne]3s^1$	29	Cu	copper	$[Ar]3d^{10}4s^1$
12	Mg	magnesium	$[Ne]3s^2$	30	Zn	zinc	$[Ar]3d^{10}4s^2$
13	Al	aluminium	$[Ne]3s^23p^1$	31	Ga	gallium	$[Ar]3d^{10}4s^24p^1$
14	Si	silicon	$[Ne]3s^23p^2$	32	Ge	germanium	$[Ar]3d^{10}4s^24p^2$
15	P	phosphorus	$[Ne]3s^23p^3$	33	As	arsenic	$[Ar]3d^{10}4s^24p^3$
16	S	sulfur	$[Ne]3s^23p^4$	34	Se	selenium	$[Ar]\ 3d^{10}4s^24p^4$
17	Cl	chlorine	$[Ne]3s^23p^5$	35	Br	bromine	$[Ar]\ 3d^{10}4s^24p^5$
18	Ar	argon	$[Ne]3s^23p^6$	36	Kr	krypton	$[Ar]3d^{10}4s^24p^6$

Chromium and copper—exceptions

The electron configurations for most elements follow the rules described and shown in Figure 3.4.1 and Figure 3.4.2 on pages 58 and 59. There are two notable exceptions: element 24—chromium, and element 29—copper. Table 3.4.5 shows these exceptions.

TABLE 3.4.5 The electron configurations for chromium and copper

Element	Electron configuration predicted using the rules above	Actual electron configuration
chromium, Cr	$1s^22s^22p^63s^23p^63d^44s^2$	$1s^22s^22p^63s^23p^63d^54s^1$
copper, Cu	$1s^22s^22p^63s^23p^63d^94s^2$	$1s^22s^22p^63s^23p^63d^{10}4s^1$

Chemists calculate that there is very little difference in energy between 3*d*- and 4*s*-orbitals, and the $3d^54s^1$ configuration for chromium is slightly more stable than the $3d^44s^2$ configuration. This is because each of the five *d*-orbitals is exactly half-filled.

Similarly, for copper, the $3d^{10}4s^1$ arrangement with five completely filled *d*-orbitals and an unfilled *s*-subshell provides additional stability compared to the filled *s*-subshell and incompletely filled *d*-subshell.

ORBITAL DIAGRAMS

The quantum mechanical model's final rule for electron configuration is that each subshell is made up of smaller components known as orbitals. Orbitals hold two electrons and can be described as regions of space surrounding the nucleus of an atom in which electrons may be found. An *s*-subshell has just one orbital. A *p*-subshell has three orbitals. A *d*-subshell has five orbitals and an *f*-subshell has seven. In diagrammatic representations of configuration, the symbol ⇅ is used to indicate two electrons in an orbital.

An orbital is a section of a subshell that only ever holds two electrons. As such, it is the smallest grouping of electrons in the quantum mechanical model. It has, however, been the subject of many different explanations of its nature. The two electrons in the orbital seem to have an interaction with each other that lasts even after they have been separated.

An **orbital diagram** showing individual electrons helps to correctly determine the electron configuration of an atom. For example, neon has 10 electrons and so will have both its first and second shells full. The subshells that have been filled are 1*s*, 2*s* and 2*p*. The first two electrons fill the 1*s*-subshell, the second two electrons go into the next highest energy level, the 2*s*-subshell. The last six electrons then fill the next highest energy level, the 2*p*-subshell. The electron configuration is written as $1s^2 2s^2 2p^6$, and the orbital diagram is shown in Figure 3.4.5. In this diagram, each box represents an orbital that can hold two electrons. Electrons have a characteristic known as spin and the arrow indicates that electrons in a pair must have opposite spin. This is the essence of the **Pauli exclusion principle** in the quantum mechanical model of the atom, which states that no two electrons can occupy the same position or state. This means that electrons in the same orbital must have opposite spin.

Figure 3.4.6 shows how the energy levels are filled in a neon atom, which has 10 electrons.

The example orbital diagram for neon shown in Figures 3.4.5 and 3.4.6 has subshells 1*s*, 2*s* and 2*p* full. **Hund's rule** describes how to arrange electrons when a *p* or *d* subshell is not full. It states that every orbital in a subshell is singly occupied before any orbital is doubly occupied. Figure 3.4.7 shows this idea. Nitrogen has three electrons in the 2*p* subshell and these are placed in separate orbitals. As the atomic number increases (to Ne), the subshell is filled.

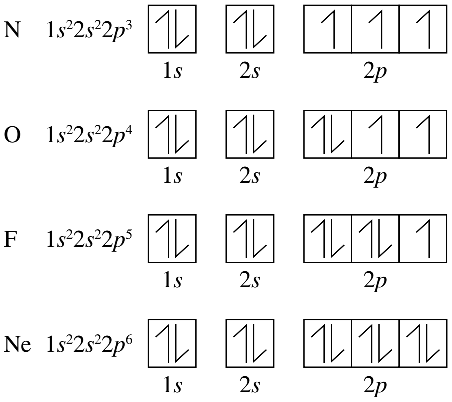


FIGURE 3.4.7 Electron configuration of nitrogen, oxygen, fluorine and neon showing the application of Hund's rule. Note that the electron configuration of nitrogen has three electrons in separate orbitals and with the same spin.

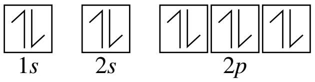


FIGURE 3.4.5 Orbital diagram showing the electron configuration of neon, also written as $1s^2 2s^2 2p^6$. Each box represents an orbital and each arrow is an electron.

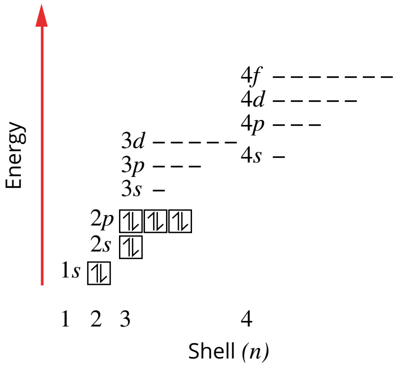


FIGURE 3.4.6 This orbital diagram shows how the orbitals in a neon atom are filled to give an electron configuration of $1s^2 2s^2 2p^6$.

SCIENCE AS A HUMAN ENDEAVOUR

Models of the atom: development of atomic theory

The 30-year period at the start of the twentieth century saw scientists hypothesise, investigate, infer and confirm new understandings relating to the nature of the most fundamental particle of matter, the atom. Before that, at least 30 elements had been characterised but little was known about the particles that comprised these substances. Laboratory equipment was basic and experimental data was used to make inferences about certain possibilities—this led to new experiments to unlock the mysteries of the atom. Between 1900 and 1935, one-third of the Nobel prizes in physics were awarded to researchers investigating the structure of the atom.

Dalton describes a basic particle of matter

Scientists investigating the structure of the atom used the insightful proposals made by John Dalton almost one hundred years earlier. In 1802, John Dalton presented the first **atomic theory of matter**. He meticulously measured the volumes of gases during reactions and observed them reacting in fixed ratios. Dalton proposed that all matter is made up of tiny spherical particles called atoms which are indivisible and indestructible.

Dalton also accurately described elements as materials containing just one type of atom and **compounds** as materials containing different types of atoms in fixed ratios. Dalton's atomic theory of matter was mostly correct. Figure 3.4.8 is a copy of a page from Dalton's notebook showing his proposed system of denoting different elements as spheres, and a representation of his idea that these would form compounds by combining in predetermined ratios. As more elements were described, scientists decided to use the system of letters in use today.

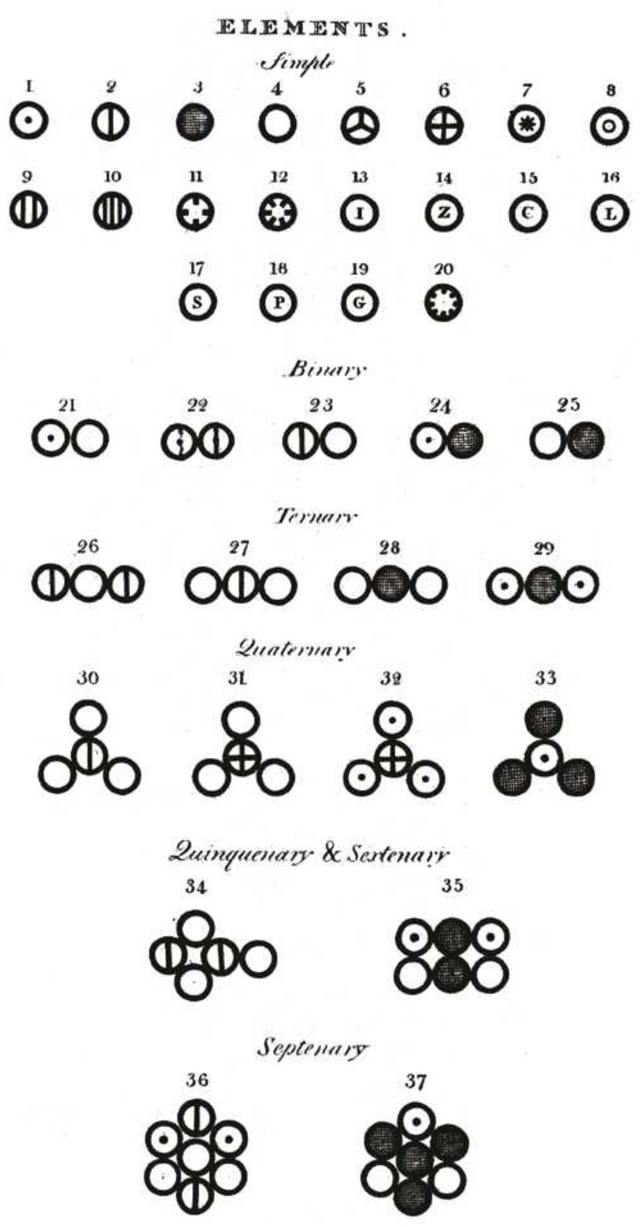


FIGURE 3.4.8 John Dalton's early system for identifying different elements

A negatively charged particle with very small mass

In the mid-nineteenth century scientists could produce electric current. A series of experiments tested an electric current passed through different gases in sealed tubes (**cathode ray tubes**) at very low pressures. At high voltage, the cathode ray tubes glowed with a coloured light and the glass on the wall opposite the negative electrode became fluorescent (Figure 3.4.9).

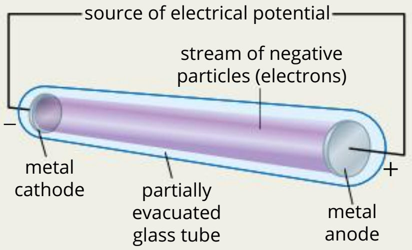


FIGURE 3.4.9 This glass cathode ray tube contains a gas at low pressure. When an electric potential is applied, a glow is observed on a phosphorescent screen. The particles responsible for the glow came from the cathode and travelled to the anode.

English physicist Joseph John Thomson repeated these experiments by using different metals as electrodes and different gases in the tube. In every case, the results were the same. The properties of the cathode rays were independent of both the gas and the metal. From his experiments in 1897, Thomson deduced that:

- the rays were a stream of particles with very low mass
- the particles came from the negative electrode (cathode) and were negatively charged
- the particles must be found in all matter and therefore were subatomic particles
- atoms must also contain positive subatomic particles since atoms have no overall charge.

Thomson proposed that the atom was a structure like a 'plum pudding' (Figure 3.4.10) with negatively charged particles throughout an area of positive charge. Thomson's research laid the foundation for other scientists to build the modern model of the atom. Importantly, excited physicists and chemists then began to investigate if there were smaller subatomic particles.

FIGURE 3.4.10 After discovering a negative particle, Thomson proposed a 'plum pudding' structure of the atom that had negative particles throughout a positive region.

Robert A. Millikan was also fascinated by the nature of this negative charge. In 1910, he devised an experiment where the movement of charged oil droplets was changed by the voltage applied. He found that the charge on various-sized oil droplets was always a multiple of 1.5924×10^{-19} coulomb. His conclusion was that this was the charge on an electron.

An extremely dense body surrounded by empty space

Between 1899 and 1911, Ernest Rutherford conducted experiments in which he fired a beam of alpha particles at a piece of extremely thin gold foil. Most of the alpha particles passed straight through the gold foil. From this observation, Rutherford deduced that the gold atoms were made up almost entirely of empty space. Figure 3.4.11 shows the apparatus and observations of Rutherford's experiments.

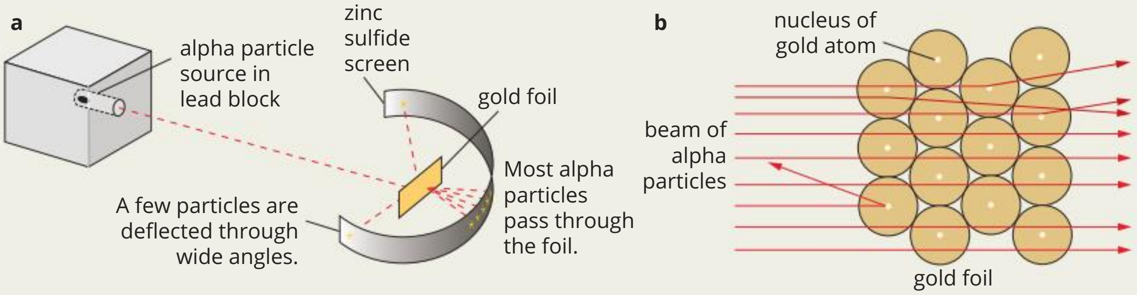


FIGURE 3.4.11 (a) Ernest Rutherford's apparatus that provided evidence for the discovery of nuclei in atoms. (b) Only those alpha particles that closely approach the nuclei in the gold foil were deflected significantly. Most particles passed almost directly through the foil.

However, Rutherford was surprised to discover that occasionally an alpha particle would bounce straight back. He stated: ‘It was about as incredible as if you had fired a 16-inch shell at a piece of tissue paper and it came back and hit you!’

This remarkable observation suggested that the centre of the atom is very small, positive and extremely dense. This central region was later named the nucleus.

Rutherford provided experimental evidence for at least two subatomic particles (the proton and the electron). Rutherford’s model (1911) proposed the following.

- Most of the mass of an atom and all of the positive charge is located in a tiny central region called the nucleus.
- Most of the volume of an atom is empty space, occupied only by electrons.
- The electrons move in circular orbits around the nucleus.
- The force of the attraction between the positive nucleus and the negative electrons is electrostatic.

Particles changing energy levels

Max Planck and Niels Bohr directed their research group to investigate energy-related phenomena. They wanted to explain why each element emits a characteristic spectrum of wavelengths of light. Bohr hypothesised that the best way to understand these emission spectra was to begin with the simplest element, hydrogen.

Bohr measured emission spectra of hydrogen first, followed by other elements. He found that the types of energy being emitted were specific to the element. In 1911, he proposed that this energy was the result of electrons moving between energy levels. This led to the shell model, which is the basis of the one in use today.

Later, in 1926, Erwin Schrödinger proposed that electrons behaved as waves around the nucleus. Using a mathematical approach and this wave theory, Schrödinger developed a model of the atom called **quantum mechanics** which describes how electrons exist in shells, subshells and orbitals.

Completion of a working model of the atom

As a working model for the atom began to emerge, scientists could still not explain why protons were not bursting out of the nucleus in every direction due to electrostatic repulsion.

In 1923, James Chadwick used a neutron chamber to collide alpha particles with boron or beryllium atoms. A type of radiation was produced that caused protons to be ejected from a sample of paraffin (Figure 3.4.12). Chadwick proposed that these heavy but uncharged particles, which he called neutrons, resided in the nucleus and that they bring stability to the nucleus by effectively shielding protons from each other.

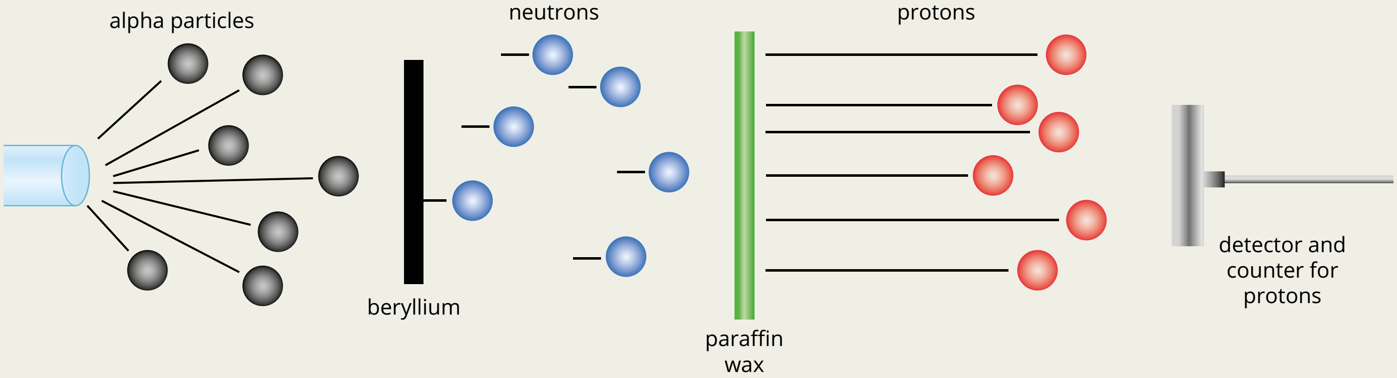


FIGURE 3.4.12 Schematic representation of the equipment used to identify the neutron. Alpha radiation pushes neutrons out of beryllium. The neutrons are heavy enough to force protons from a material such as paraffin wax.

Recent developments

Scientists of the twentieth century greatly enhanced our understanding of atomic structure (Figure 3.4.13).

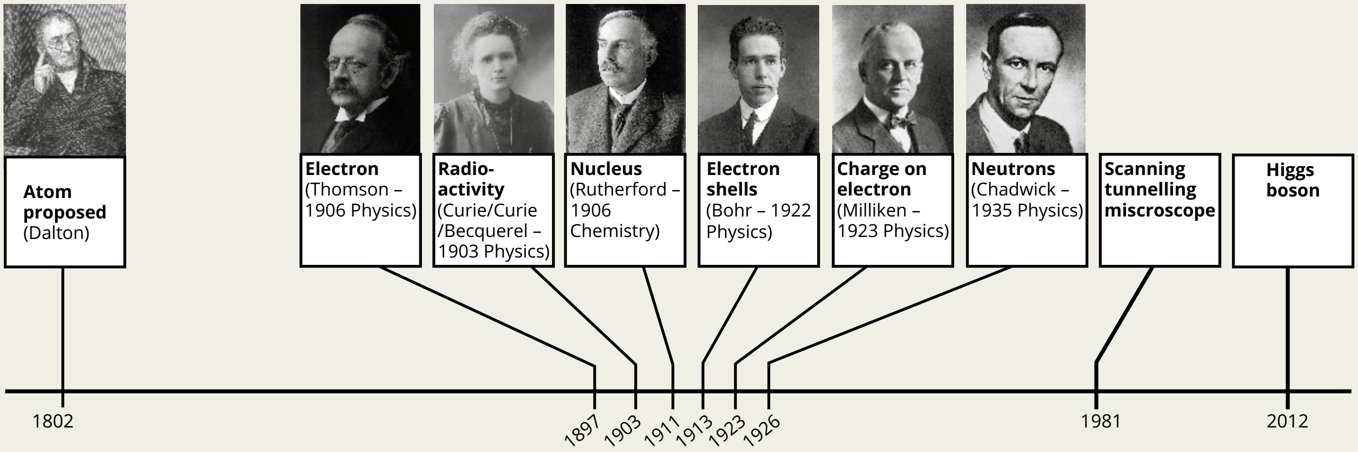


FIGURE 3.4.13 Timeline of key discoveries leading to our model of the atom. The scientist and year of their Nobel prize is shown in brackets.

It wasn't until 1981 that the scanning tunnelling microscope (STM), capable of viewing atoms, was developed by IBM researchers Gerd Binnig and Heinrich Rohrer. Using STMs, scientists confirmed that atoms are indeed spherical.

In the spirit of Rutherford and Chadwick, bombarding atoms with high speed particles and other atoms continues to this day. These collisions can create new elements and allow characterisation of other subatomic particles. The **Large Hadron Collider** (LHC) has enhanced our understanding of ever smaller and more fundamental parts of the atom. Data produced in its massive 27 km underground 'atom smasher' supports the physicist's 'standard model' of subatomic particles. The most recent addition to this model is what some call the 'God' particle, the **Higgs boson**, confirmed in 2012. It is proposed that this particle is responsible for an atom demonstrating the measurable property of mass.

Review

1 Review Figure 3.4.11 on page 67 showing Rutherford's gold foil experiment. It showed that all the mass of the atom was in one place and that it was heavy enough to reflect alpha particles. Sketch the expected results of his experiment if Thomson's plum pudding model was correct.

2 Describe the finding of Millikan's oil drop experiment.

3 Explain the cathode ray tube experiment and the conclusions that Thomson developed during his investigations using this equipment.

4 In investigating phenomena, a scientist uses careful observation or experiments to report regularities or trends, known as laws. Subsequently, a scientific theory (a structure suggested by these laws) is devised to systematically explain them in a scientifically rational manner. Using two of the scientists described in this module, identify the law and the theory related to their experiments.

PA
1.1.2

3.4 Review

SUMMARY

- The quantum mechanical model of the atom describes that subshells are energy levels within the major shells. The subshells consist of orbitals.
- A subshell can be regarded as a region of space of particular energy in which electrons are found. The character of a subshell can be *s*, *p*, *d* and *f*.
- An orbital is an area around the nucleus that can hold a pair of electrons that have opposite spin according to Hund's rule.
- An orbital diagram shows each orbital as a box containing up to two electrons (shown as arrows with opposite directions).

Shell	Subshells	Orbitals in subshell
1	1*s*	1
2	2*s* 2*p*	1 3
3	3*s* 3*p* 3*d*	1 3 5
4	4*s* 4*p* 4*d* 4*f*	1 3 5 7

- The 4*s*-subshell is lower in energy than the 3*d*-subshell, so the fourth shell accepts two electrons before the third shell is completely filled.
- The Pauli exclusion principle states that no two electrons can occupy the same position (quantum state) in an atom.
- The Aufbau principle states that subshells of lower energy will be filled before higher energy subshells. This provides maximum stability.
- The order of filling of subshells is 1*s*2*s*2*p*3*s*3*p*4*s*3*d*4*p*; however, the order of writing the electron configuration is 1*s*2*s*3*s*3*p*3*d*4*s*4*p*.
- Cr and Cu are exceptions to the order of filling because their 4*s* subshell is not full (has a single electron).
- Electron configurations specify the number of electrons in each subshell using a superscript ($1s^22s^22p^63s^23p^63d^{10}4s^24p^6$).
- Hund's rule states that *p*- or *d*-subshell electrons will exist in separate orbitals where possible.
- Condensed electron configuration uses the noble gas before the element (in square brackets) to shorten the writing of electron configuration.

KEY QUESTIONS

Retrieval

1 Recall the number of electrons that can be held in each of the subshell types *s*, *p*, *d* and *f*. Copy and complete the table by drawing lines to connect the shell to its number of electrons.

Subshell		Electron number
s		14
p		2
d		10
f		6

2 Use Figure 3.4.1 first and then Figure 3.4.2 on pages 58 and 59 to write out the order of filling of electron subshells ending with 5*s*.

3 Identify subshells from the following list that are not part of the quantum mechanical model.

1*s* 1*p* 2*s* 2*p* 3*s* 3*p* 1*d* 4*s* 4*f* 2*f*

4 Write the electron configuration of each of the atoms listed in the following table.

Element (atomic number)	Electron configuration using the shell model	Electron configuration using the subshell model
boron (5)	2,3	$1s^22s^22p^1$
lithium (3)		
chlorine (17)		
sodium (11)		
neon (10)		
potassium (19)		
scandium (21)		
iron (26)		
bromine (35)		

Comprehension

5 Draw the orbital diagrams for Cl, Be, C, Mn and P. State the five elements in order of greatest number of unpaired electrons to least number of unpaired electrons.

6 Use Figure 3.4.2 on page 59 to 'read' the order of filling of subshells and write the electron configuration of Fe, which has 26 electrons.

7 Show the condensed electron configurations that represent the following elements.

- **a** Ca
- **b** Na
- **c** Cl

8 Describe the essential difference between the shell model and the subshell model of the atom in terms of energy levels.

Analysis

9 Identify the elements shown by the following electron configurations.

- **a** $1s^22s^22p^63s^23p^63d^{10}4s^24p^1$
- **b** $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$
- **c** $1s^22s^22p^63s^23p^63d^{10}4s^1$
- **d** $1s^22s^22p^63s^23p^63d^54s^2$
- **e** $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^2$
- **f** $[\mathrm{Kr}]4d^{10}5s^25p^3$
- **g** $[\mathrm{Kr}]4d^{10}5s^25p^6$
- **h** $[\mathrm{Ar}]4s^1$

10 Identify an electron configuration of a sulfur atom that has been heated so that its electrons are no longer in the ground state.

Chapter review

03

KEY TERMS

absorption
adsorption
alpha particle
atom
atomic number
atomic theory of matter
Aufbau principle
beta radiation
Bohr model
carbon nanotube
catalyst
cathode ray tube
chemical property
chemical symbol
composite
compound
condensed electron configuration
electron
electron configuration
electron configuration notation
electron shell
electrostatic attraction
element
emission spectrum
energy level
excited state
gamma radiation
ground state
half-life
Higgs boson
Hund's rule
ionisation energy
isotope
Large Hadron Collider
mass number
matter
molecule
nanomaterial
nanoparticle
nanoscale
nanoscience
nanotechnology
neutron
noble gas notation
nuclear symbol notation
nucleon
nucleus
octet rule
orbital
orbital diagram
outermost shell
Pauli exclusion principle
periodic table (of elements)
physical property
porous
positron emission tomography (PET)
proton
quanta
quantum mechanical model
quantum mechanics
radioactive
radioactivity
radioisotope
scanning tunnelling microscope (STM)
scientific model
subatomic particle
subshell
transportation
valence electron
valence shell
weighted average

KEY QUESTIONS

Retrieval

1 State the maximum number of electrons in the first three shells of an atom.
A 8,8,18
B 2,8,12
C 2,8,8
D 2,8,18

2 Isotopes of an element contain different numbers of:
A protons
B neutrons
C nucleons
D electrons

3 Identify the atom or atoms *not* in the ground state from the electron configurations for the four atoms shown below.
A $1s^{2}2s^{2}2p^{6}3s^{1}$
B $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}$
C $1s^{2}2s^{2}2p^{6}3s^{2}3p^{5}4s^{2}$
D $1s^{2}2s^{2}2p^{3}$

4 Convert the following lengths into nanometres (express your solutions in scientific notation).
a 5 cm
b 12 mm
c 2 km

5 Describe where you would find 99.97% of the mass of an atom.

6 Recall how protons, neutrons and electrons are arranged in an atom.

7 An atom of chromium can be represented by the symbol $^{52}_{24}\text{Cr}$.
a Identify its atomic number and mass number.
b Identify the number of electrons, protons and neutrons in the chromium atom.

8 Show the nuclear symbol notation of each element.
a phosphorus
b bromine
c beryllium
d gallium

9 Write the name and symbol of the element with each of these electron configurations.
a 2,7
b 2,8,18,1
c 2,8,5
d 2,8,18,17

10 Write electron configurations, using subshell notation, for the following elements. The atomic number of each element is shown in brackets.
a helium (2)
b carbon (6)
c fluorine (9)
d aluminium (13)
e argon (18)
f nickel (28)
g bromine (35)

11 Identify the elements shown by the following electron configurations.

- **a** $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
- **b** $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$
- **c** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
- **d** $1s^2 2s^2 2p^6 3s^2 3p^5$

12 Identify the following atoms.

- **a** $[\mathrm{Ne}]3s^1$
- **b** $[\mathrm{Ne}]3s^2 3p^1$
- **c** $[\mathrm{Ar}]3d^3 4s^2$
- **d** $[\mathrm{Ar}]3d^{10} 4s^2 4p^1$

Comprehension

13 Zinc oxide powder and zinc oxide nanoparticles both absorb UV light. Describe what property of zinc oxide nanoparticles makes them more suitable than zinc oxide powder for use in sunscreen.

14 Determine the number of valence electrons for an atom with 10 electrons. Draw a diagram of the electron configuration of this atom.

15 The nucleus of an atom has a radius of the order of 10^{-13} cm and the atom has a radius of the order of 10^{-9} cm. If the nucleus could be scaled up to the size of an orange (radius 10 cm), determine what the radius of the atom on that same scale would be.

16 Explain the rule describing the maximum number of electrons in each shell.

17 Explain why an element is defined by its atomic number and not its mass number.

18 Determine what these atoms are and calculate the number of neutrons each contains.

- **a** ${}^{119}_{50}X$
- **b** ${}^{23}_{11}X$
- **c** ${}^{96}_{42}X$
- **d** ${}^{27}_{13}X$

19 Explain why the number of electrons in an atom equals the number of protons.

20 Explain why isotopes of the same element have the same chemical properties.

21 Determine the quantum mechanical model of electron configuration that corresponds to the shell model electron configuration 2,8,16.

Analysis

22 Compare the mass and charge of protons, neutrons and electrons.

23 Two atoms both have 10 neutrons in their nucleus. The first has 9 protons and the other has 10 protons. Determine if they are isotopes of the same element. Explain your answer.

24 Analyse the following list of atoms.

${}^{27}_{13}A$ ${}^{40}_{20}B$ ${}^{32}_{16}C$ ${}^{14}_{7}D$ ${}^{8}_{3}E$ ${}^{15}_{7}F$ ${}^{19}_{9}G$ ${}^{20}_{9}H$ ${}^{4}_{2}I$

- **a** Identify which pairs of atoms are isotopes of the same element.
- **b** Identify which atoms have equal numbers of protons and neutrons in the nucleus.
- **c** Identify which is an isotope of sulfur.
- **d** Determine which has one more electron than a magnesium atom.
- **e** Identify how many different elements are shown.

25 Differentiate between the terms ‘shell’, ‘subshell’ and ‘orbital’ using a fluorine atom as an example.

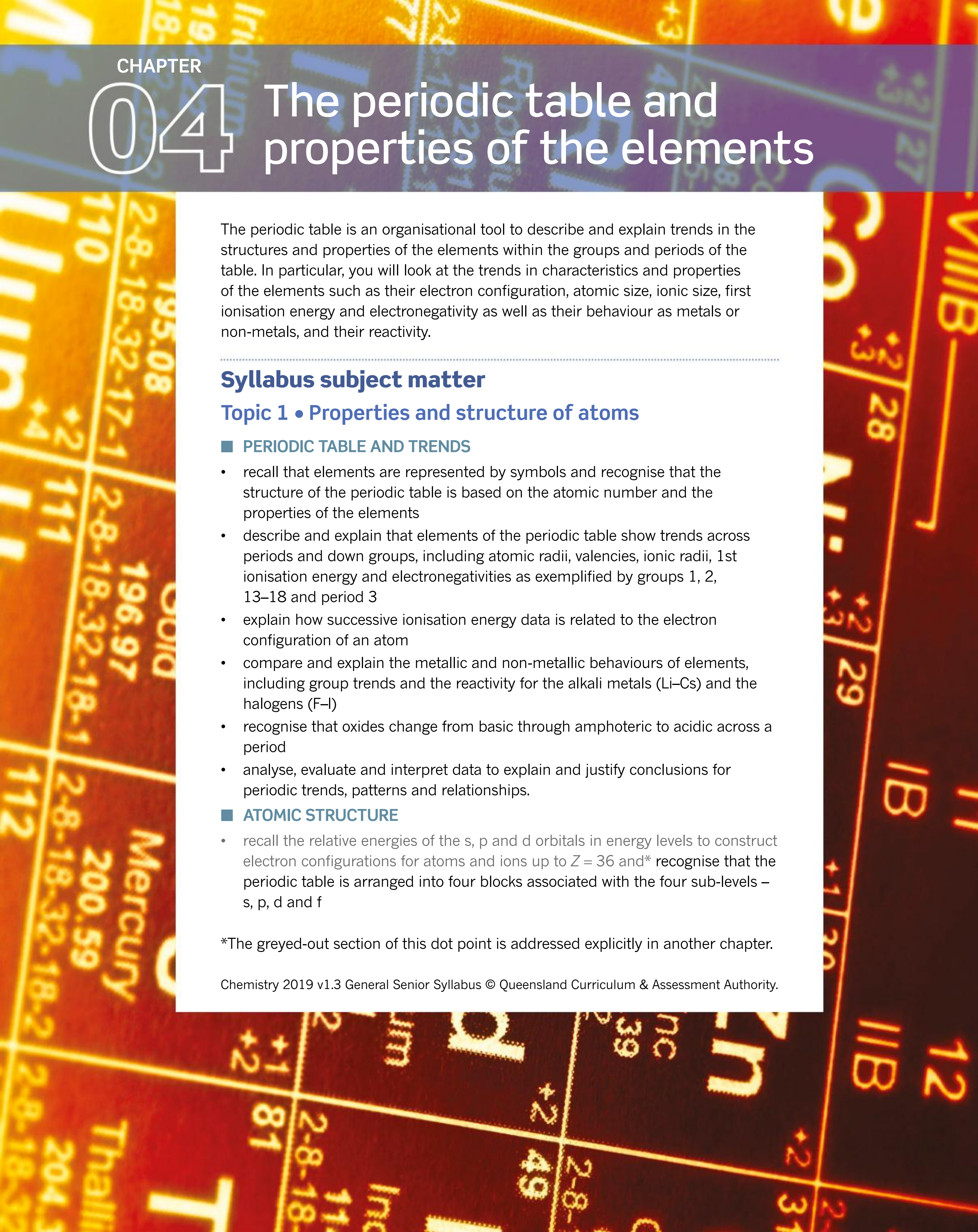
26 Explore why the fourth electron shell begins filling before the third shell is completely filled with reference to the quantum mechanical model of the atom.

27 Determine why the 4*s*-subshell in chromium and copper is only half-filled compared to the other transition elements in period 4.

Knowledge utilisation

28 The range of possible uses for carbon nanotubes is increasing dramatically because of their unique structure. Research and explain two recent advances in this area.

29 Quantum mechanics and the structure of the nucleus are fields where physics and chemistry are closely aligned. Research the momentous discovery of 2012 and describe how this important particle relates to fundamental concepts of both chemistry and physics.



CHAPTER

04 The periodic table and properties of the elements

The periodic table is an organisational tool to describe and explain trends in the structures and properties of the elements within the groups and periods of the table. In particular, you will look at the trends in characteristics and properties of the elements such as their electron configuration, atomic size, ionic size, first ionisation energy and electronegativity as well as their behaviour as metals or non-metals, and their reactivity.

Syllabus subject matter

Topic 1 • Properties and structure of atoms

PERIODIC TABLE AND TRENDS

- recall that elements are represented by symbols and recognise that the structure of the periodic table is based on the atomic number and the properties of the elements
- describe and explain that elements of the periodic table show trends across periods and down groups, including atomic radii, valencies, ionic radii, 1st ionisation energy and electronegativities as exemplified by groups 1, 2, 13–18 and period 3
- explain how successive ionisation energy data is related to the electron configuration of an atom
- compare and explain the metallic and non-metallic behaviours of elements, including group trends and the reactivity for the alkali metals (Li–Cs) and the halogens (F–I)
- recognise that oxides change from basic through amphoteric to acidic across a period
- analyse, evaluate and interpret data to explain and justify conclusions for periodic trends, patterns and relationships.

ATOMIC STRUCTURE

- recall the relative energies of the s, p and d orbitals in energy levels to construct electron configurations for atoms and ions up to $Z = 36$ and* recognise that the periodic table is arranged into four blocks associated with the four sub-levels – s, p, d and f

*The greyed-out section of this dot point is addressed explicitly in another chapter.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

4.1 The periodic table

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recognise that elements are categorised into groups, periods and blocks within the periodic table based on their configuration
- determine the number of valence electrons of an element based on its group
- determine the period of an element based on the number of occupied shells
- identify the period and group of an element based on its configuration
- identify an element from its group and period.

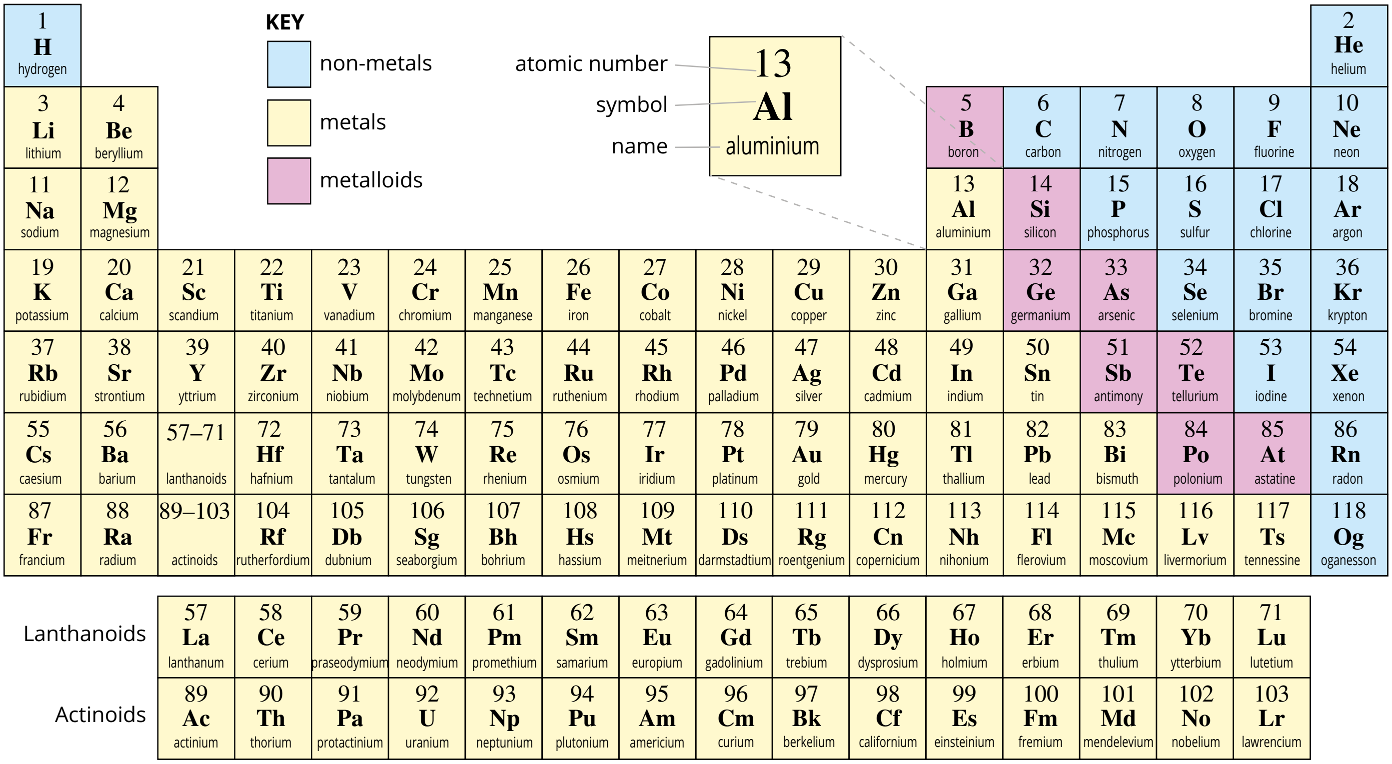
Chapter 3 looked at research into elements and the nature of atoms, the existence of subatomic particles, their charge and mass, the way particles are arranged in an atom and the way they behave. As scientists' understanding of the atom improved, and more elements were discovered, a way of organising this knowledge was needed.

The **periodic table** (Figure 4.1.1) is one of the most useful reference tools available to chemists. It minimises the need to memorise isolated facts about different elements, and provides a framework in which to organise our understanding. By knowing the properties of particular elements and trends within the table, chemists are able to organise what would otherwise be an overwhelming collection of disorganised information.

THE MODERN PERIODIC TABLE

The number of protons (the **atomic number**) is what makes one element fundamentally different from another element. The elements in the modern periodic table are arranged in rows in order of increasing atomic number.

Chemists use the number of electrons in the outer shell, called the **valence electrons** (see Chapter 3), to organise the elements into columns. The form of the periodic table in common use is shown in Figure 4.1.1.



KEY

🟦 non-metals

🟨 metals

🟪 metalloids

atomic number — 13
symbol — **Al**
name — aluminium

1 **H** hydrogen																	2 **He** helium
3 **Li** lithium	4 **Be** beryllium											5 **B** boron	6 **C** carbon	7 **N** nitrogen	8 **O** oxygen	9 **F** fluorine	10 **Ne** neon
11 **Na** sodium	12 **Mg** magnesium											13 **Al** aluminium	14 **Si** silicon	15 **P** phosphorus	16 **S** sulfur	17 **Cl** chlorine	18 **Ar** argon
19 **K** potassium	20 **Ca** calcium	21 **Sc** scandium	22 **Ti** titanium	23 **V** vanadium	24 **Cr** chromium	25 **Mn** manganese	26 **Fe** iron	27 **Co** cobalt	28 **Ni** nickel	29 **Cu** copper	30 **Zn** zinc	31 **Ga** gallium	32 **Ge** germanium	33 **As** arsenic	34 **Se** selenium	35 **Br** bromine	36 **Kr** krypton
37 **Rb** rubidium	38 **Sr** strontium	39 **Y** yttrium	40 **Zr** zirconium	41 **Nb** niobium	42 **Mo** molybdenum	43 **Tc** technetium	44 **Ru** ruthenium	45 **Rh** rhodium	46 **Pd** palladium	47 **Ag** silver	48 **Cd** cadmium	49 **In** indium	50 **Sn** tin	51 **Sb** antimony	52 **Te** tellurium	53 **I** iodine	54 **Xe** xenon
55 **Cs** caesium	56 **Ba** barium	57–71 lanthanoids	72 **Hf** hafnium	73 **Ta** tantalum	74 **W** tungsten	75 **Re** rhenium	76 **Os** osmium	77 **Ir** iridium	78 **Pt** platinum	79 **Au** gold	80 **Hg** mercury	81 **Tl** thallium	82 **Pb** lead	83 **Bi** bismuth	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89–103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **Nh** nihonium	114 **Fl** flerovium	115 **Mc** moscovium	116 **Lv** livermorium	117 **Ts** tennessine	118 **Og** oganesson

Lanthanoids	57 **La** lanthanum	58 **Ce** cerium	59 **Pr** praseodymium	60 **Nd** neodymium	61 **Pm** promethium	62 **Sm** samarium	63 **Eu** europium	64 **Gd** gadolinium	65 **Tb** trebium	66 **Dy** dysprosium	67 **Ho** holmium	68 **Er** erbium	69 **Tm** thulium	70 **Yb** ytterbium	71 **Lu** lutetium
Actinoids	89 **Ac** actinium	90 **Th** thorium	91 **Pa** protactinium	92 **U** uranium	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

FIGURE 4.1.1 The form of the periodic table in common use

The modern periodic table has several key features (numbering is based on the IUPAC convention).

- The periodic table is arranged in order of increasing atomic number.
- The horizontal rows are known as **periods** and are labelled 1–7.
- The vertical columns are known as **groups** and are labelled 1–18.
- **Main group elements** are elements in groups 1, 2 and 13–18.
- The elements in groups 3–12 are known as **transition metals**.

Some periodic tables you will see also indicate other properties of the elements such as boiling point or whether the element is a solid, liquid or gas at room temperature.

The International Union of Pure and Applied Chemistry (IUPAC) is an international organisation of chemists working to advance their science and create common standards and nomenclature. IUPAC work includes naming new elements and standardising methods of measuring such things as atomic weights. IUPAC encourages free sharing of scientific information.

Groups

Elements in the periodic table are arranged into vertical columns called groups. The group number of the main group's elements can be used to determine the number of valence electrons in an atom of the element.

In groups 1 and 2, the number of valence electrons is equal to the group number. For example, magnesium is in group 2 and therefore has two valence electrons.

In groups 13–18, the number of valence electrons is equal to the group number minus 10. For example, oxygen is in group 16 so oxygen has six outer-shell electrons. Similarly, neon is in group 18 so neon has eight valence electrons. Helium is an important exception. It is located in group 18 but only has two valence electrons. Helium is placed in group 18 because it is unreactive, like other group 18 elements. This information is summarised in Table 4.1.1.

TABLE 4.1.1 The number of valence electrons in elements belonging to each group

Group	Number of valence electrons
1	1
2	2
13	3
14	4
15	5
16	6
17	7
18	8*

*Helium has two valence electrons.

Groups of elements

The electrons in the outer shell of an atom (the valence electrons) are the electrons that are involved in chemical reactions. As a consequence, the number of valence electrons determines many of the chemical properties that an element exhibits.

Table 4.1.2 shows the names of some groups in the periodic table.

TABLE 4.1.2 Names of different groups in the periodic table

Group	Name
1	alkali metals
2	alkaline earth metals
17	halogens
18	noble gases

Elements in the same group have the same number of valence electrons, so elements in the same group have similar properties. For example, the **alkali metals** are elements in group 1 (with the exception of hydrogen). They are all relatively soft metals and are highly reactive with water and oxygen.

Metals

Consider the **electron configurations** of the atoms of the first three metals of this group in Table 4.1.3. The valence shell of each atom of each element in group 1 contains one electron in an *s*-subshell. This similarity in the valence shell structure gives these elements similar chemical properties.

TABLE 4.1.3 Electron configurations of selected metals in group 1

	Full electron configuration	Condensed electron configuration
Li	$1s^2 2s^1$	$[He]2s^1$
Na	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	$[Ar]4s^1$

Halogens

Fluorine, chlorine, bromine and iodine are **halogens** (group 17). They are all coloured and highly reactive (Figure 4.1.2). Their electron configurations are listed in Table 4.1.4. Notice how all these elements have a highest-energy subshell electron configuration of s^2p^5.

FIGURE 4.1.2 Three examples of halogens. These conical flasks contain, from left to right, chlorine (Cl, pale green), bromine (Br, red-brown) and iodine (I, purple).

TABLE 4.1.4 Electron configurations of selected halogens in group 17

	Full electron configuration	Condensed electron configuration
F	$1s^2 2s^2 2p^5$	$[He]2s^2 2p^5$
Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Br	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$	$[Ar]3d^{10} 4s^2 4p^5$
I	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$	$[Kr]4d^{10} 5s^2 5p^5$

Noble gases

The **noble gases** (group 18) are a particularly interesting group. The noble gases have a very stable electron arrangement: helium has a full outer shell and the other members of this group have a stable octet of valence electrons (eight electrons). Chemical reactions involve the rearrangement of valence electrons to achieve a stable outer shell. Noble gases have a stable electron configuration, so they do not tend to lose or gain electrons. This means that the noble gases have low **reactivity**.

The arrangement of electrons in atoms is responsible for the **periodicity** (periodic pattern) of element properties.

Periods

The horizontal rows in the periodic table are called periods. Periods are numbered 1–7. The number of a period gives information about the electron configuration of an element. The period an element is located within is equal to the number of occupied electron shells in the element's atoms. For example, the outer shell of magnesium and chlorine is the third shell and both of these elements are in period 3:

$$\begin{array}{ll} \text{Mg} & 1s^2 2s^2 2p^6 3s^2 \\ \text{Cl} & 1s^2 2s^2 2p^6 3s^2 3p^5 \end{array}$$

Similarly, the elements in period 5 all have outer-shell electrons in the fifth shell.

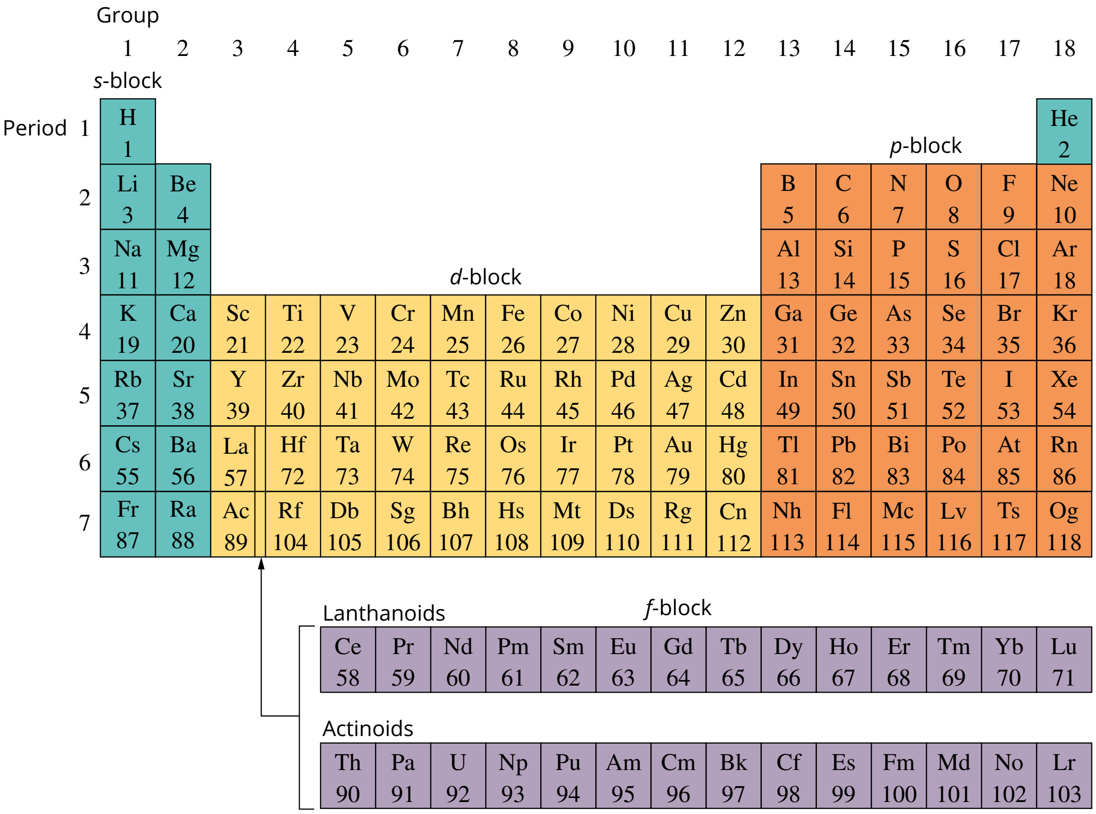
Blocks

The periodic table has four main **blocks**. The elements in each block have the same type of subshell (*s*, *p*, *d* or *f*) as their highest energy subshell. For example, the highest energy subshell of an element in the *s*-block is the *s*-subshell.

The *s*-block contains the elements in group 1 (alkali metals), group 2 (**alkaline earth metals**), hydrogen and helium. These elements have a half-filled or fully filled *s*-subshell, that is, s^1 or s^2, as the highest energy subshell configuration. Table 4.1.5 and Figure 4.1.3 show which groups of elements fall into the four different blocks.

TABLE 4.1.5 Elements in the different blocks of the periodic table

Block	Elements	Highest energy subshell configurations
s-block	groups 1 and 2 and helium	s^1 or s^2
p-block	groups 13–18 (except helium)	s^2p^1 to s^2p^6
d-block	groups 3–12	d^1s^2 to $d^{10}s^2$
f-block	lanthanoids and actinoids	4*f*-subshell progressively being filled in the lanthanoids 5*f*-subshell progressively being filled in the actinoids



Group

s-block

p-block

d-block

Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1																	He 2
2	Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
3	Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
4	K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
5	Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
6	Cs 55	Ba 56	La 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
7	Fr 87	Ra 88	Ac 89	Rf 104	Db 105	Sg 106	Bh 107	Hs 108	Mt 109	Ds 110	Rg 111	Cn 112	Nh 113	Fl 114	Mc 115	Lv 116	Ts 117	Og 118

f-block

Lanthanoids													
Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
Actinoids													
Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

FIGURE 4.1.3 Colour is used to distinguish between the *s*, *p*, *d* and *f* blocks of elements.

4.1 Review

SUMMARY

- The periodic table is a tool for organising elements according to their chemical and physical properties.
- The elements of the periodic table are arranged in order of increasing atomic number.
- Columns in the periodic table are known as groups and are numbered 1–18.
- The number of valence electrons in an atom of an element can be determined by the group in which it is located.
- The main group elements are in groups 1, 2 and 13–18 in the periodic table.
- Elements in between the main group elements (groups 3–12) are known as transition metals.
- Rows in the periodic table are known as periods and are numbered 1–7.
- The number of occupied electron shells of an atom of an element is equal to the number of the element's period.
- The periodic table has four main blocks of elements; the elements in each block have the same type of subshell (*s*, *p*, *d* or *f*) as their highest energy subshell.

KEY QUESTIONS

Retrieval

1 Define the terms 'period', 'group' and 'block'.

2 State the IUPAC name for the following groups with reference to the periodic table.

- **a** group 1
- **b** group 2
- **c** group 17
- **d** group 18

3 State the number of valence electrons in atoms of elements in these groups:

- **a** group 1
- **b** group 15
- **c** group 17
- **d** group 2

Comprehension

4 Describe the key features of the periodic table.

Analysis

5 Determine the name and symbol of the following elements. In addition, identify the shell configuration of each element.

- **a** second element in group 14
- **b** second element in period 2
- **c** element that is in group 18 and period 3

6 Identify the period of the periodic table in which each of the following elements belongs. Use the periodic table in Figure 4.1.1 on page 76 as a reference.

- **a** K
- **b** F
- **c** He
- **d** H
- **e** U
- **f** P

4.2 Periodic trends

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the full and condensed electron configuration of an element
- recognise that the periodic trends are related to the electron configuration of the elements
- explain the concept of effective nuclear charge and its effect on the electrostatic attraction of the valence electrons
- define electronegativity and explain how it is affected by the effective nuclear charge across the period and down the group
- define atomic radius and explain how it is affected by the effective nuclear charge across the period and down the group
- define ionic radius and explain how it is affected by the effective nuclear charge across the period and down the group
- define the first ionisation energy and how it is affected by the effective nuclear charge across the period and down the group
- recognise the exceptions to the trend in the first ionisation energy and how the exceptions are related to the filling of the subshells.

The periodic table does not just provide information about an element's electron configuration. It can also be used as a tool for summarising the relative properties of elements and explaining the trends observed in those properties.

You have already seen in Module 4.1 that the group number of an element identifies how many valence electrons an atom of that element has. The period indicates how many electron shells are occupied in an atom of an element. Properties such as atomic radii, ionic radii, electronegativity and ionisation energy show common trends in the periodic table.

Periodic trends were observed by Dmitri Mendeleev and formed the basis of the table of the elements that he first published in 1869. Mendeleev described the way the properties of the elements vary as the periodic law.

ELECTRON CONFIGURATION

To understand the reason for the periodicity of element properties, look at two groups of elements that Mendeleev recognised as being similar. The first group is the alkali metals (group 1). The elements in this group (lithium, sodium, potassium, rubidium and caesium) are all relatively soft metals and are highly reactive with water and oxygen. Consider their electron configurations, as shown in Table 4.2.1.

TABLE 4.2.1 Electron configuration (full and condensed) for the group 1 metals

	Full electron configuration	Condensed electron configuration
Li	$1s^22s^1$	$[He]2s^1$
Na	$1s^22s^22p^63s^1$	$[Ne]3s^1$
K	$1s^22s^22p^63s^23p^64s^1$	$[Ar]4s^1$
Rb	$1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^1$	$[Kr]5s^1$
Cs	$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^66s^1$	$[Xe]6s^1$

These elements have similar valence shell electron configurations—all have one electron in an *s*-subshell. This similarity in a group's arrangement of electrons gives elements similar properties and is responsible for the periodicity of element properties.

You can also see that the number of electron shells increases moving down the group. The increase in electron shells means that the valence electrons are in higher energy subshells and have a weaker attraction to the nucleus. The decrease in the attractive force between the valence electrons and the nucleus as you move down a group causes trends in properties to be observed within a group.

EFFECTIVE NUCLEAR CHARGE

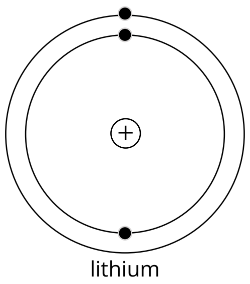


FIGURE 4.2.1 A lithium atom with one valence electron and two electrons in the inner shell. The atom has an effective nuclear charge of +1.

The **effective nuclear charge** of an atom is a measure of the attractive force felt by the valence shell electrons towards the nucleus. Effective nuclear charge can be used to predict the properties of elements and explain trends observed in the periodic table.

Consider an atom of lithium, which has an atomic number of three. It has three protons in its nucleus, two electrons in the first shell and one electron in the second shell (Figure 4.2.1).

The valence shell electron is attracted to the three positive charges in the nucleus. This electron is also repelled by the two electrons in the inner shell. The electrons in the inner shell shield the valence shell electron from the attraction of the nucleus. The valence shell electron is effectively attracted to the nucleus as if there were a +1 nuclear charge. This atom is therefore said to have an effective nuclear charge of +1.

In atoms with two or more shells filled with electrons, the attraction between the nucleus and valence shell electrons is reduced by repulsion between the inner-shell electrons and the valence shell electrons. This is commonly known as the **shielding effect**.

In general:

effective nuclear charge = number of protons in the nucleus − number of total inner-shell electrons.

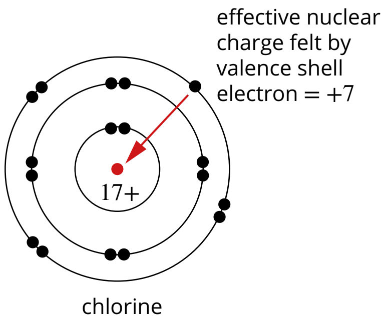


FIGURE 4.2.2 A Bohr diagram for the chlorine atom showing the valence shell electron shielded by 10 electrons in shells *n* = 1 and *n* = 2

For example, an atom of chlorine (Figure 4.2.2) has 17 protons and seven valence shell electrons; the number of electrons in the inner shells is 10. The effective nuclear charge of a chlorine atom is 17 − 10 = +7.

Worked example 4.2.1

EFFECTIVE NUCLEAR CHARGE

Determine the effective nuclear charge of an atom of aluminium.	
Thinking	**Working**
Determine the number of electrons in an atom of the element, using the periodic table as a reference.	The atomic number of aluminium is 13. Therefore, an atom of aluminium has 13 protons and 13 electrons.
Use the number of electrons to determine the electron configuration.	With 13 electrons, the electron configuration is $1s^22s^22p^63s^23p^1$.
Determine the effective nuclear charge. Effective nuclear charge = number of protons − number of inner-shell electrons	The third shell is the valence shell in this atom. There are 10 inner-shell electrons, which in this atom are electrons in the first and second shell. Effective nuclear charge = 13 − 10 = +3

➤ Try yourself 4.2.1

EFFECTIVE NUCLEAR CHARGE

Determine the effective nuclear charge of an atom of fluorine.

Consider the atoms of two different elements in group 1, lithium and sodium, shown in Figure 4.2.3.

As for all group 1 elements, the valence electron of a lithium atom and a sodium atom experiences an effective nuclear charge of +1.

Moving down a group, you can see the following changes.

- The effective nuclear charge remains constant but the number of electron shells between the nucleus and the valence electrons increases.
- The valence electrons are further from the nucleus. Consequently, they will be pulled less strongly towards the nucleus due to the reduced **electrostatic attraction** and increased repulsion from the inner-shell electrons (greater shielding effect).

Now consider sodium and chlorine. They are both in period 3 in the periodic table (Figure 4.2.4).

The effective nuclear charge experienced by the valence shell electrons in atoms of elements increases from left to right across a period, as you have seen for sodium and chlorine.

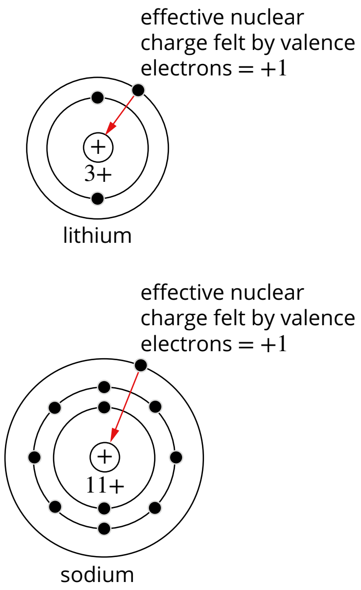


FIGURE 4.2.3 Lithium and sodium atoms both have an effective nuclear charge of +1.

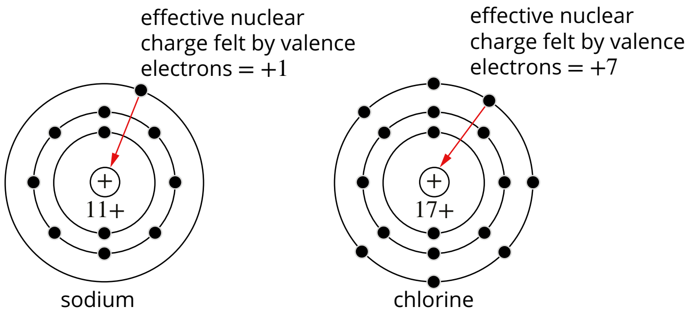


FIGURE 4.2.4 The effective nuclear charges of two period 3 elements, sodium and chlorine, are +1 and +7 respectively.

The effective nuclear charge of an atom of a main group element is equal to the number of valence electrons in the atom, as summarised in Table 4.2.2.

TABLE 4.2.2 Effective nuclear charges of main group elements

Group	Effective nuclear charge
1	+1
2	+2
13	+3
14	+4
15	+5
16	+6
17	+7
18*	+8

*Helium has an effective nuclear charge of +2.

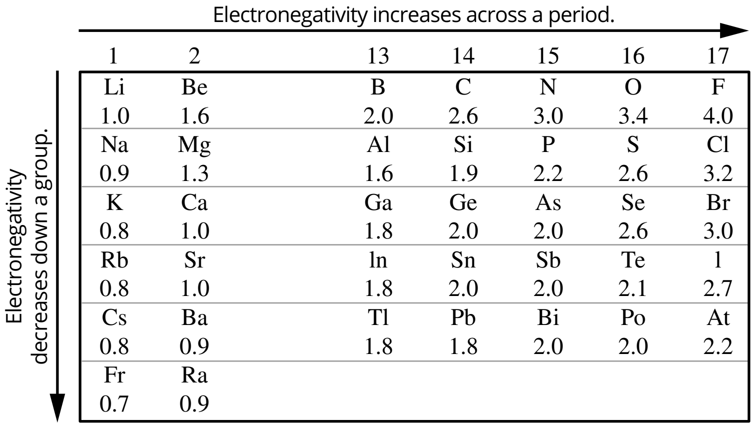
Table 4.2.3 summarises how the attraction between the nucleus and valence electrons changes in the periodic table.

TABLE 4.2.3 The changes in attraction between the nucleus and valence electrons within groups and periods of the periodic table

	Trend in effective nuclear charge	Trend in attraction between the nucleus and valence electrons	Explanation for the trend in attraction between the nucleus and valence electrons
down a group	remains constant	decreases	Effective nuclear charge stays constant down a group, but the valence electrons are held less strongly as they are further from the nucleus, as there are more shells, and there is an increased shielding effect as a result of the greater number of inner electron shells.
left to right across a period	increases	increases	The valence electrons are more attracted to the nucleus as the effective nuclear charge increases.

ELECTRONEGATIVITY

Electronegativity is the ability of an atom to attract electrons towards itself. The more strongly the valence electrons of an atom are attracted to the nucleus of the atom, the greater the electronegativity. Therefore, the greater the effective nuclear charge of an atom, the greater the electronegativity. Figure 4.2.5 shows the electronegativity of many of the main group elements.



1	2	13	14	15	16	17
Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Ga 1.8	Ge 2.0	As 2.0	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	In 1.8	Sn 2.0	Sb 2.0	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	Tl 1.8	Pb 1.8	Bi 2.0	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

FIGURE 4.2.5 The electronegativity of elements generally decreases down a group and increases across a period, from left to right.

The trends observed in the electronegativity of the elements are summarised in Table 4.2.4.

TABLE 4.2.4 Trends in electronegativity in groups and periods of the periodic table

	Trend in electronegativity	Explanation
down a group	decreases	The effective nuclear charge stays constant and the number of shells increases down a group. Therefore, valence electrons are less strongly attracted to the nucleus as they are further from the nucleus. As a result, electronegativity decreases.
left to right across a period	increases	The number of occupied shells in the atoms remains constant but the effective nuclear charge increases across a period. Therefore, the valence electrons become more strongly attracted to the nucleus. As a result, electronegativity increases.

ATOMIC RADIUS

Atomic radius is a measurement used for the size of atoms. It can be regarded as the distance from the nucleus to the valence shell electrons. It is usually measured by halving the distance between the nuclei of two atoms of the same element that are bonded together. Figure 4.2.6 depicts the atomic radii of many of the main group elements. For example, in a hydrogen molecule (H_2) the two nuclei are 74 picometres (pm) apart. The radius of each hydrogen atom is assumed to be half of that distance, i.e. 37 pm.

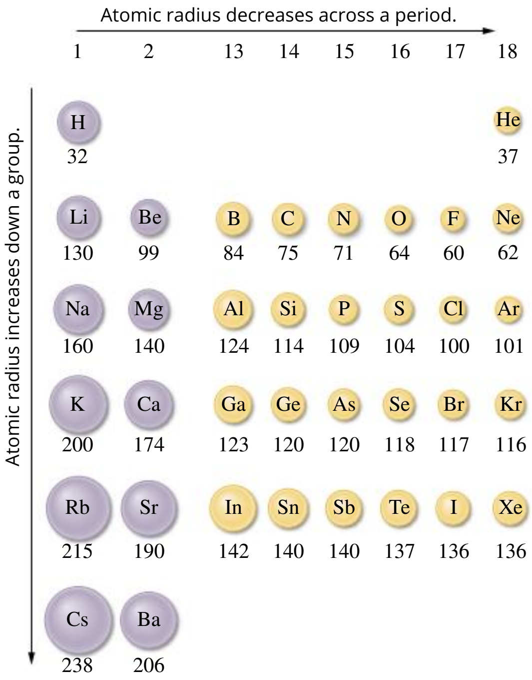


FIGURE 4.2.6 The relative sizes of atoms of selected main group elements. The atomic radii are given in picometres (pm). A picometre is 10^{-12} m.

Table 4.2.5 explains the trends in atomic radii in the periodic table.

TABLE 4.2.5 Trends in atomic radii in the periodic table

	Trend in atomic radii	Explanation
down a group	increases	Effective nuclear charge stays constant and the number of shells increases as you move down a group. As a result, atomic radii increase.
left to right across a period	decreases	As you move across a period, the number of occupied shells in the atoms remains constant but the effective nuclear charge increases. The valence electrons become more strongly attracted to the nucleus, so atomic radii decrease across a period.

IONIC RADIUS

An ion is any atom of an element with a different number of protons than electrons. If an atom loses electrons, it becomes positively charged as there are now more protons than electrons. A positively charged ion is called a **cation**. If an atom gains electrons, it becomes negatively charged and is called an **anion**. You will learn more about ions in Chapter 7.

Atomic theory proposes that ions gain additional stability when they attain the electron configuration of a noble gas (group 18). Metallic elements tend to lose one or more electrons to form positive ions (cations), whereas non-metallic elements tend to gain electrons to form negative ions (anions). The **ionic radius** is defined as the radius of an atom's ion.

Similarly to the atomic radii, it is impossible to precisely determine the ionic radius. While it is possible to make a fairly reasonable assumption that the atomic radius is half the distance between atoms, this is not possible for ions.

There are multiple ways to determine ionic radii, including experimental observations of the ion placement in the crystal lattice or refractive index to theoretical calculations. However, the ionic radius is dependent on the coordination of the ion (that is how many oppositely charged ions are touching it) and the nature of those oppositely charged ions.

While there are large uncertainties in the actual dimensions, it is important to understand the trends that exist across the period and down the group. Due to these uncertainties, comparison of ionic radii necessitates that the data comes from the same source.

Consider an atom of sodium, which has an atomic number of 11. It has a completely filled second shell and a single electron in the third shell. When it forms a cation, it loses the electron from the third shell to attain the electron configuration of neon. Since this electron comes from the outermost shell, its loss will make the cation much smaller (102 pm compared to 160 pm) (Table 4.2.6).

TABLE 4.2.6 Electron configurations of the sodium atom and sodium ion

	Full electron configuration
Na	$1s^22s^22p^63s^1$
Na^+	$1s^22s^22p^6$

Moving across the period from sodium to magnesium, the effective nuclear charge increases. When magnesium forms an ion it also attains the electron configuration of neon (Table 4.2.7).

TABLE 4.2.7 Electron configurations of the magnesium atom and magnesium ion

	Full electron configuration
Mg	$1s^22s^22p^63s^2$
Mg^{2+}	$1s^22s^22p^6$

Atoms and ions that have the same electron configuration are said to be isoelectronic.

For instance, Mg^{2+} and Na^+ are isoelectronic.

The effective nuclear charge of a Mg^{2+} ion is higher than the effective nuclear charge of a Na^+ ion; therefore, the attraction between the nucleus and the valence electrons increases from sodium to magnesium ions. Since the electron configurations of the ions formed are the same (**isoelectronic**) but the effective nuclear charge has increased, the size of the cations decreases, as shown in Table 4.2.8.

TABLE 4.2.8 Comparison of the effective nuclear charge and electron configuration of the cations in period 3

	Effective nuclear charge	Electron configuration of the stable cation	Ionic radius (pm)
Na^+	1	$1s^22s^22p^6$	102 pm
Mg^{2+}	2	$1s^22s^22p^6$	72 pm
Al^{3+}	3	$1s^22s^22p^6$	54 pm

Consider an atom of chlorine, which has an atomic number of 17 and five electrons in the third shell. When it forms an anion, it gains an electron to fill the third shell, attaining the electron configuration of argon (Table 4.2.9).

TABLE 4.2.9 Electron configurations of the chlorine atom and chlorine ion

	Full electron configuration
Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
Cl^-	$1s^2 2s^2 2p^6 3s^2 3p^6$

Although the effective nuclear charge of the chlorine atom and chloride ion are the same, the chloride ion has a greater number of valence electrons than the atom. This means that the electrons are held less strongly and as a result the ionic radius of the chloride ion is greater than the atomic radius of chlorine (181 pm compared to 100 pm).

While the effective nuclear charge of the element increases as you cross the period from left to right, from sulfur to chlorine, the electron configuration of the anions formed are isoelectronic with the argon atom. In an identical way to the ionic radii of the cations the valence electrons are more attracted to the nucleus as the effective nuclear charge increases. This means the ionic radius of the chloride ion is smaller than the ionic radius of the sulfide ions (181 pm compared to 184 pm).

Since the anions of period 3 gain electrons to complete the third shell as opposed to the cations which lose their electrons from the third shell, the cations are much smaller than the anions formed by elements in the same period.

The trend down the group is similar to the trend in atomic radii. As the number of electron shells between the nucleus and the valence electrons increases, the valence electrons are further from the nucleus. This means that the ionic radius increases down the group. Table 4.2.10 explains the trends in ionic radii in the periodic table.

TABLE 4.2.10 Trends in ionic radii in the periodic table

	Trend in ionic radii	Explanation
down a group	increases	Effective nuclear charge stays constant and the number of shells increases as you move down a group. As a result, ionic radii increase.
left to right across a period	decreases for species of the same charge	As you move across a period, the number of occupied shells in the ions remains constant for the species of the same charge but the effective nuclear charge increases. The valence electrons become more strongly attracted to the nucleus, so ionic radii decrease across a period for the species of the same charge. Anions are larger than cations in the same period.

Figure 4.2.7 depicts the ionic radii of many of the main group elements.

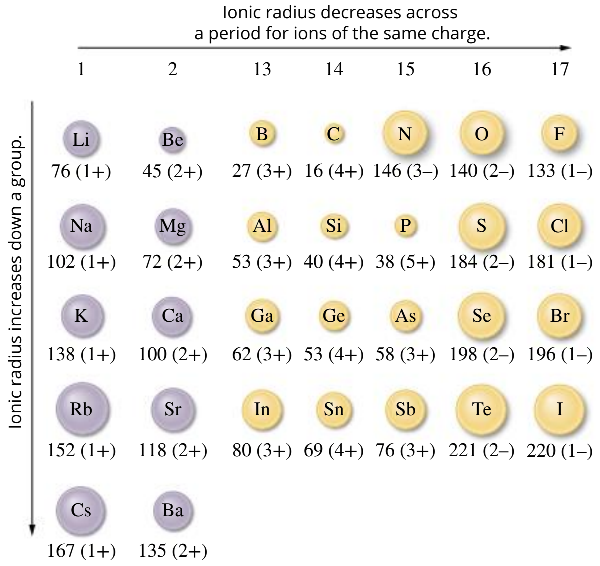


FIGURE 4.2.7 The relative sizes of common ions of selected main group elements. The ionic radii are given in picometres (pm). A picometre is 10^{-12} m. Where two ions are formed, the greater ionic radius of the more common ion is shown.

FIRST IONISATION ENERGY

When an element is heated, its electrons can move to higher energy shells. If an atom is given sufficient energy, an electron can be completely removed from the atom. If this occurs, the atom will now have one less electron than the number of protons in the nucleus, and becomes a cation as shown in the equation:

$$M(g) + \text{energy} \rightarrow M^+(g) + e^-$$

The process of removing an electron from an atom and forming an ion is called **ionisation**. The valence electrons are removed first because they are the furthest electrons from the nucleus and the least strongly held.

The energy required to remove one electron from an atom of an element in the gas phase is called the **first ionisation energy**. For example, the ionisation energy of sodium is 496 kJ per mole of sodium atoms.

Figure 4.2.8 shows the first ionisation energies of most main group elements.

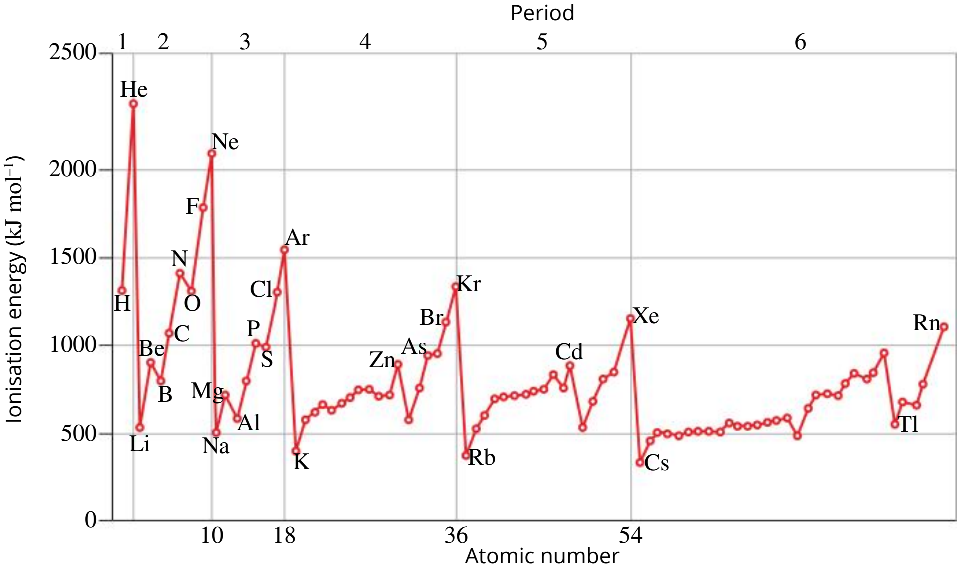


FIGURE 4.2.8 The first ionisation energy generally increases from left to right across a period.

The magnitude of the first ionisation energy reflects how strongly the valence electrons are attracted to the nucleus of the atoms. The more strongly the valence electrons are attracted to the nucleus, the more energy is required to remove them from the atom and the higher the first ionisation energy. Table 4.2.11 explains the typical trends in ionisation energy in the periodic table.

TABLE 4.2.11 The trend in ionisation properties in groups and periods of the periodic table

	Trend in ionisation energy	Explanation
down a group	decreases	Effective nuclear charge stays constant and the number of shells increases down a group. Therefore, the valence electrons are less attracted to the nucleus as they are further from the nucleus. As a result, the energy required to overcome the attraction between the nucleus and the valence electron is less, and the first ionisation energy decreases down a group.
across a period (left to right)	increases	Effective nuclear charge increases and the number of occupied shells remains constant across a period. As a result, the valence electrons become more strongly attracted to the nucleus, and more energy is required to remove an electron. Therefore, the first ionisation energy increases across a period.

The *s*- and *p*-block elements of the periodic table follow these patterns. As the effective nuclear charge increases across a period, so too does the ionisation energy. Down the group the effective nuclear charge stays the same but the increased atomic radius and greater shielding effect from the inner electron shells means that the valence electrons in elements lower in a group have a smaller electrostatic attraction to the nucleus. Therefore, they can be removed more easily, meaning the ionisation energy decreases.

WS 1.1.3

PA 1.1.3

Exceptions to the rule

As you have seen, across a period the first ionisation energy increases due to the increase in the effective nuclear charge; however, there are two exceptions to this trend in period 2. There is a decrease from 900 to 801 kJ mol^{-1} going from beryllium (Be) to boron (B) and a decrease from 1402 to 1314 kJ mol^{-1} going from nitrogen (N) to oxygen (O). To understand these exceptions better, you first need to determine the electron configuration of the elements involved. Table 4.2.12 shows the electron configurations of the beryllium and boron atoms.

TABLE 4.2.12 Electron configurations of the beryllium atom and boron atom

	Full electron configuration
Be	$1s^2 2s^2$
B	$1s^2 2s^2 2p^1$

The electron that is removed from beryllium comes from a full *s*-subshell, whereas the electron that is removed from boron comes from a *p*-subshell. In Module 3.4, you saw that the *p*-subshell is higher in energy than the *s*-subshell (Aufbau principle). In addition, the filled 2*s*-subshell shields the electron in the $2p^1$ orbital so the electrostatic attraction of the nucleus is reduced. These two factors mean that less energy is required to remove the electron from the $2p^1$ orbital and so the first ionisation energy of B is lower than that of Be, even though the atom has a greater effective nuclear charge.

You have seen in Module 3.4 that as you cross the periodic table the additional electrons enter unfilled orbitals until all the orbitals in the same subshell contain one electron (Hund's rule). In nitrogen, the three electrons are in the three different *p* orbitals; however, when an atom of oxygen is ionised the electron is removed from a filled *p* orbital. The electron configurations of the nitrogen atom and oxygen atom are shown in Table 4.2.13. Since like charges repel, when two electrons are in close proximity to one another, the electrons are destabilised. Even though the electron that is removed from oxygen is higher in energy than the electron in nitrogen, the first ionisation energy of oxygen is lower than the first ionisation energy of nitrogen.

TABLE 4.2.13 Electron configurations of the nitrogen atom and oxygen atom

	Full electron configuration
N	$1s^2 2s^2 2p^3$
O	$1s^2 2s^2 2p^4$

Competing effects

It is important to recognise that there are two **competing effects** here.

1 The effective nuclear charge increases across the period but the valence electron is in the same shell and so the electrostatic attraction increases across the period. Thus, the electron is harder to remove (consistent with the general trend in ionisation energy across a period).

2 The nature of the orbital from which the electron is removed means that the electron is in a higher energy subshell. Since the electron is in a higher energy subshell it is easier to remove.

It is important to note that since these effects act in opposite directions it is impossible to precisely predict which effect has a greater impact. Only by observing the result can you identify which effect is more significant. In this case, the higher energy of the electron in a *p*-subshell compared to the *s*-subshell or destabilising effect as a result of pairing is more significant than the increased effective nuclear charge. Therefore, the first ionisation energy decreases from beryllium to boron and nitrogen to oxygen.

This competing effect concept can be explored in greater detail by considering the difference in the trends in first ionisation energy for elements in groups 5 and 6 from period 2 to period 4 as shown in Table 4.2.8 on page 86. In each case the electron is being removed from a half-filled *p* orbital (group 5) or a filled *p* orbital (group 6) and so the ionisation energy results from competition between the stability of the electron and the increased effective nuclear charge.

In period 2, the difference between N and O is 88 kJ mol^{-1} but decreases to 12 and 4 kJ mol^{-1} in period 3 (P and S) and period 4 (As and Se) respectively. Thus, the impact of the electron repulsion decreases from period 2 to period 4.

While it is not important to understand the reasons why the effect of the electron repulsion decreases, it is important to interpret and evaluate from observations and data which effect is more significant.

4.2 Review

SUMMARY

- The effective nuclear charge of an atom is a measure of the attractive force felt by the valence electrons towards the nucleus.
- The effective nuclear charge is calculated by subtracting the total number of inner-shell electrons from the number of protons in the nucleus.
- Electronegativity is the ability of an element to attract electrons towards itself.
- Atomic radius is a measurement used for the size of atoms. It can be regarded as the distance from the nucleus to the outermost electrons.
- The first ionisation energy is the energy required to remove one electron from an atom of an element in the gas phase and is represented by the equation

  $$\mathrm{M(g) + energy \rightarrow M^{+}(g) + e^{-}}$$

- Table 4.2.14 summarises how properties of elements have specific trends within the groups and periods of the periodic table.

TABLE 4.2.14 Summary of changes in properties of elements in the periodic table

Property	Down a group	Across a period (left to right)
effective nuclear charge	no change	increases
atomic radius	increases	decreases
ionic radius	increases	decreases for species of the same charge; larger for anions than cations in the same period
electronegativity	decreases	increases
first ionisation energy	decreases	increases

KEY QUESTIONS

Retrieval

1 Define the term 'effective nuclear charge' of an atom and determine the effective nuclear charge of an atom of carbon.

2 Define the first ionisation energy of an atom.

Comprehension

3 Determine the electron configuration of the following atoms or ions.
  - **a** Ca
  - **b** Al^{3+}
  - **c** N^{3-}
  - **d** P

4 Explain the term 'shielding effect'.

5 Explain the relationship between electronegativity and effective nuclear charge.

6 Figure 4.2.5 on page 84 gives electronegativity values for the elements in groups 1, 2 and 13–17 of the periodic table.
  - **a** Determine the name and symbol of the element that has the:
    - **i** highest electronegativity
    - **ii** lowest electronegativity.
  - **b** Identify the group which has the following changes:
    - **i** greatest change in electronegativity as you go down the group
    - **ii** smallest change in electronegativity as you go down the group
  - **c** Explain why the elements of group 18 are usually omitted from tables that give electronegativity values.

7 Explain why ionisation energy increases from left to right across a period.

8 Explain why the size of the Al^{3+} cation is different from the size of the atom from which it was formed.

Analysis

9 Compare and contrast the trends in atomic and ionic radii in the periodic table using specific examples to illustrate your explanation.

10 Sort the following in order of increasing atomic radius: based on your understanding of the trends in the periodic table.

N, B, Ga, Al, Cl

11 Organise the following elements in order of increasing first ionisation energies: using the periodic table on page 76 (Figure 4.1.1).

Na, He, Al, K, S, Ca and P

12 Predict whether Mg^{2+} is larger than F^{-} using the periodic table on page 76 (Figure 4.1.1). Explain your choice based on the structure of the two ions.

13 Deduce why the number of subatomic particles in an atom increases across a period but the size of the atom decreases.

4.3 Impact of the periodic trends on atomic properties

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- use the periodic table to determine which elements are metals, metalloids and non-metals
- compare and explain differences in the reactivity of the alkali metals as you go down group 1
- compare and explain differences in the reactivity of the halogens as you go down group 17
- explain how the differences in the reactivity of the alkali metals are related to the trends in ionisation energy of the elements
- explain how the differences in the reactivity of the halogens are related to the trends in electron affinities of the elements
- recognise how the nature of the oxides changes from basic through amphoteric to acidic across period 3.

You have learnt that the effective nuclear charge and the number of electron shells in an atom of an element can be used to predict some properties of elements. You will remember that the effective nuclear charge of an atom is the attractive force felt by the valence electrons towards the nucleus. Effective nuclear charge increases as you move across a period. As you move down a group, the effective nuclear charge stays constant but the atomic radius increases due to the additional electron shell in each period. The changes in effective nuclear charge and atomic radii account for periodic trends in ionisation energy, reactivity and metallic character.

METALLIC CHARACTER

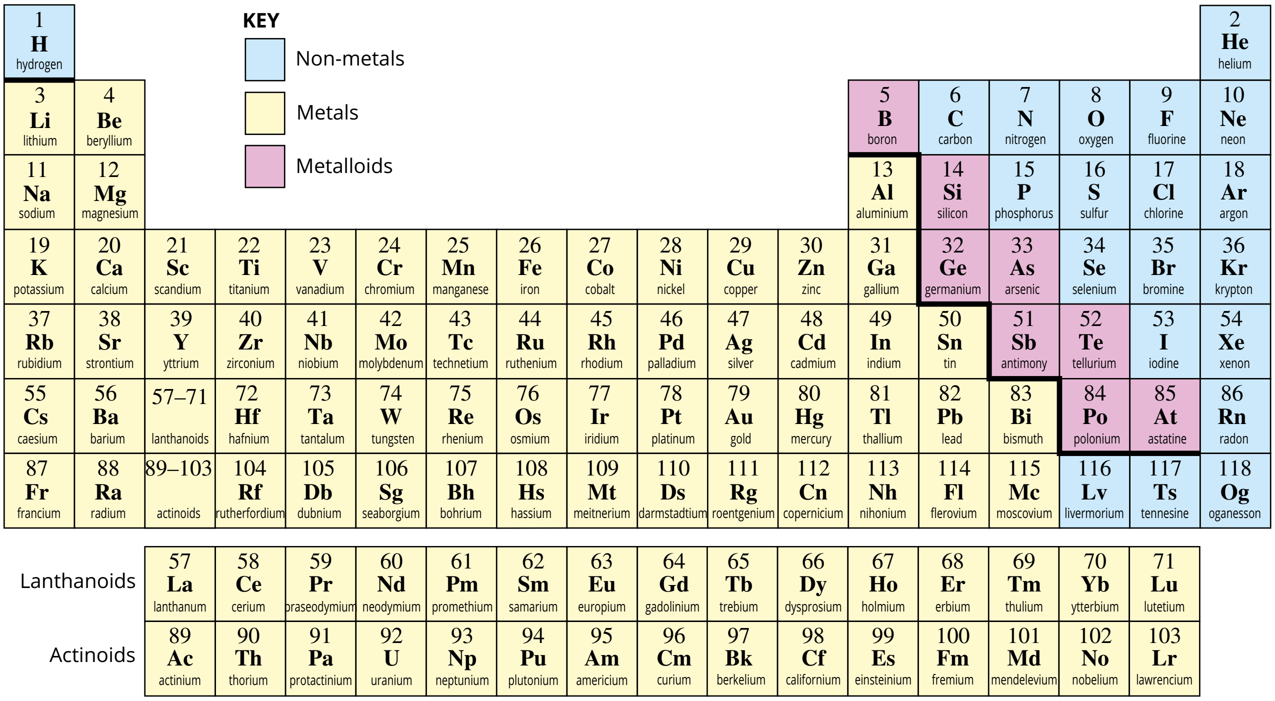
Metallic character of an element is defined as the ease with which its atoms lose electrons. Therefore, metallic character is related to the ionisation energy of the element. Metals conduct electricity and are usually solids at room temperature. Conversely, non-metallic elements usually do not conduct electricity and many are gases at room temperature.

The differences between the properties of metals and non-metals are related to the number of electrons in the outer shell of their atoms. In general, elements with atoms containing one, two or three valence electrons tend to behave as metals, whereas those with four or more valence electrons behave as non-metals. You will study this in more detail in Chapter 6.

Many periodic tables have a solid jagged line with metals located to the left side and non-metals to the right (Figure 4.3.1). Elements known as the **metalloids** are located adjacent to the line between the metals and non-metals. The metalloids exhibit both metallic and non-metallic properties. Silicon is one of the most abundant metalloids. It is a brittle solid, which is a common property of non-metals. However, it is also a semiconductor, meaning it conducts electricity, making it useful in many electronic devices such as computers and calculators.

REACTIVITY

The reactivity of an element is an indication of how easily an atom of that element loses or gains electrons.



KEY

- 🟦 Non-metals
- 🟨 Metals
- 🟪 Metalloids

1 **H** hydrogen																	2 **He** helium
3 **Li** lithium	4 **Be** beryllium											5 **B** boron	6 **C** carbon	7 **N** nitrogen	8 **O** oxygen	9 **F** fluorine	10 **Ne** neon
11 **Na** sodium	12 **Mg** magnesium											13 **Al** aluminium	14 **Si** silicon	15 **P** phosphorus	16 **S** sulfur	17 **Cl** chlorine	18 **Ar** argon
19 **K** potassium	20 **Ca** calcium	21 **Sc** scandium	22 **Ti** titanium	23 **V** vanadium	24 **Cr** chromium	25 **Mn** manganese	26 **Fe** iron	27 **Co** cobalt	28 **Ni** nickel	29 **Cu** copper	30 **Zn** zinc	31 **Ga** gallium	32 **Ge** germanium	33 **As** arsenic	34 **Se** selenium	35 **Br** bromine	36 **Kr** krypton
37 **Rb** rubidium	38 **Sr** strontium	39 **Y** yttrium	40 **Zr** zirconium	41 **Nb** niobium	42 **Mo** molybdenum	43 **Tc** technetium	44 **Ru** ruthenium	45 **Rh** rhodium	46 **Pd** palladium	47 **Ag** silver	48 **Cd** cadmium	49 **In** indium	50 **Sn** tin	51 **Sb** antimony	52 **Te** tellurium	53 **I** iodine	54 **Xe** xenon
55 **Cs** caesium	56 **Ba** barium	57–71 lanthanoids	72 **Hf** hafnium	73 **Ta** tantalum	74 **W** tungsten	75 **Re** rhenium	76 **Os** osmium	77 **Ir** iridium	78 **Pt** platinum	79 **Au** gold	80 **Hg** mercury	81 **Tl** thallium	82 **Pb** lead	83 **Bi** bismuth	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89–103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **Nh** nihonium	114 **Fl** flerovium	115 **Mc** moscovium	116 **Lv** livermorium	117 **Ts** tennesine	118 **Og** oganesson

Lanthanoids	57 **La** lanthanum	58 **Ce** cerium	59 **Pr** praseodymium	60 **Nd** neodymium	61 **Pm** promethium	62 **Sm** samarium	63 **Eu** europium	64 **Gd** gadolinium	65 **Tb** trebium	66 **Dy** dysprosium	67 **Ho** holmium	68 **Er** erbium	69 **Tm** thulium	70 **Yb** ytterbium	71 **Lu** lutetium
Actinoids	89 **Ac** actinium	90 **Th** thorium	91 **Pa** protactinium	92 **U** uranium	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fremium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

FIGURE 4.3.1 A periodic table showing the metals, metalloids and non-metals. The jagged black line separates the metals (left) from the non-metals (right). Note: hydrogen is a non-metal.

Reactivity of metals

When metals react, they lose electrons. Therefore, the reactivity of metals is a measure of how easily an atom of a metallic element can lose electrons. The weaker the attraction of the valence electrons to the nucleus, the more easily the electrons can be lost.

Moving down a group, the electrostatic attraction between the nucleus and the valence electrons decreases due to the increasing number of electron shells in the atoms. This means that reactivity of metals increases down a group. The rate at which metals react with water indicates their relative reactivity.

From left to right across the periodic table, the effective nuclear charge of the atoms increases and it becomes more difficult for the element to lose electrons. Within the metals, this means there is a decrease in reactivity. This agrees with the observations from Table 4.3.1, which describes the reaction of some of the group 1 and 2 metals with water. Table 4.3.1 shows that the group 1 metals are more reactive than the group 2 metals. The table also shows that the reactivity with water increases from sodium through potassium to rubidium, that is, as the period increases from 3 to 5.

TABLE 4.3.1 Reaction of some group 1 and 2 metals with water. In each case, a reaction results in the formation of hydrogen gas.

Period	Group	Element	Reaction with water
3	1	sodium	vigorous, producing enough energy to melt the sodium, which fizzes and skates on the water surface
4	1	potassium	violent, making crackling sounds as the heat evolved ignites the hydrogen produced by the reaction
5	1	rubidium	violent explosion
3	2	magnesium	no reaction at room temperature but will react with steam
4	2	calcium	slow reaction at room temperature

In summary, the reactivity of metals:

- increases down a group as it is easier for a metal with a greater number of shells to lose electrons
- decreases across the period as the increasing effective nuclear charge makes it more difficult for a metal to lose electrons.

Reactivity of non-metals

Atoms of non-metallic elements undergo chemical reactions to gain electrons and form a stable **octet arrangement**. The energy released when an atom gains an electron is known as **electron affinity** as shown in the equation:

$$X(g) + e^- \rightarrow X^-(g) + \text{energy}$$

The more easily a non-metal can attract or share electrons, the more reactive that non-metal is. Non-metallic elements that have fewer electron shells and higher effective nuclear charges will have a greater attractive force between the nucleus and valence electrons and thus be more reactive. This means that the higher the electron affinity, the greater the reactivity of the non-metal element; therefore, electron affinity should decrease down the group. However, in Table 4.3.2 the electron affinity of fluorine is less than that of chlorine.

TABLE 4.3.2 Electron affinities of the halogens

Element	Electron affinity, kJ mol^{-1}
fluorine	–328
chlorine	–349
bromine	–324
iodine	–295

The smaller size of the electron clouds for period 2 elements results in a greater electron repulsion between electrons in filled orbitals than the larger elements. This means that chlorine would be expected to be more reactive than fluorine because more energy is released when the chlorine atom gains an electron; however, fluorine is much more reactive than chlorine. Fluorine is more reactive than chlorine, despite the lower electron affinity, because the reaction involves a number of other transformations and the energy released in these steps compensates for the lower electron affinity. This is outside the scope of this course so it is important to recognise that both electron affinity and reactivity decrease for a non-metal down a group.

The trend in electron affinity across a period and down a group is similar to the trend in ionisation energy because both concepts are related to the effective nuclear charge and atomic size of an atom of an element.

Therefore, the reactivity of non-metals:

- decreases down a group as it is harder for a non-metal atom to attract electrons into its valence shell with a greater number of shells
- increases across the period as the increasing effective nuclear charge makes it easier for a non-metallic atom.

The energy needed to ionise metals is less than the energy required to form non-metal cations, whereas more energy is released when non-metals gain electrons compared to metals. This explains why metals tend to lose electrons while non-metals gain electrons.

Change in the nature of the oxides across a period

An **ionic bond** is formed between two atoms when there is a large difference in electronegativity (greater than 1.8) between two elements. Since oxygen is highly electronegative it forms ionic bonds with all metals.

Therefore, across a period (from left to right) the oxides go from ionic to covalent. This trend is most obvious for the elements in periods 2 and 3. Table 4.3.3 shows the trend in period 3.

TABLE 4.3.3 Relationship between the electronegativity difference, bonding and acidic nature of the oxide of elements in period 3

Element	Electronegativity	Electronegativity difference[1]	Bonding	Acidic nature of the oxide
sodium	0.9	2.5	ionic	strong base
magnesium	1.3	2.1	ionic	weak base
aluminium	1.6	1.8	ionic	amphoteric
silicon	1.9	1.5	covalent	weak acid
phosphorus	2.6	1.2	covalent	strong acid
sulfur	2.6	0.8	covalent	strong acid
chlorine	3.2	0.2	covalent	strong acid

[1] electronegativity of oxygen is 3.4

The acidic nature of the oxides changes across the period and this can be observed through the reaction of the oxide with water. Table 4.3.4 describes the reaction of some of the oxides of period 3 elements with water to highlight the acidic nature of the oxide.

TABLE 4.3.4 Reaction of some of the period 3 oxides with water

Group	Element	Reaction of the oxide with water
1	sodium	highly soluble; reacts exothermically (releases heat) to form a highly basic solution of sodium hydroxide
2	magnesium	slightly soluble; reacts with water to form a weakly basic solution of magnesium hydroxide
3	aluminium	insoluble in water; reacts with both acids and bases so is termed **amphoteric**
4	silicon	completely insoluble in water; reacts with very strong bases so it is a weakly acidic oxide
5	phosphorus	the two oxides of phosphorus, P_2O_3 and P_2O_5, react with water to form strongly acidic solutions
6	sulfur	the two oxides of sulfur, SO_2 and SO_3, react with water to form strong acids
7	chlorine	reacts with water to form a very strong acid

4.3 Review

SUMMARY

- Metals tend to lose electrons and non-metals tend to gain electrons.
- From left to right across a period the elements change from metals to metalloids to non-metals.
- Chemical reactivity is the ease with which an element undergoes a reaction.
- Chemical reactivity for metals increases down a group but decreases across a period and is related to the trend in ionisation energy of the atom of the element.
- Chemical reactivity of non-metals decreases down a group but increases across a period and is related to the trend in electron affinity of the atom of the element.
- The nature of the oxides changes from basic through amphoteric to acidic across period 3.
- Many trends in the physical properties of elements in the periodic table can be explained using two key ideas.
  - From left to right across a period, the effective nuclear charge of atoms increases, so the attractive force felt between the valence electrons and the nucleus increases.
  - Down a group, the number of shells in an atom increases so that the valence electrons are further from the nucleus and are held less strongly.

KEY QUESTIONS

Retrieval

1 Define the term ‘electron affinity’.

2 Identify whether the following elements are metals, metalloids or non-metals using the periodic table on page 93.

a K
b N
c Ge
d Cu
e Sn
f I

Comprehension

3 Determine whether the oxides of the following elements will be acidic, basic or amphoteric.

a C
b Li
c Ca
d Br

4 Explain the difference in reactivity with water between lithium and caesium.

5 Explain the relationship between electron affinity and effective nuclear charge.

Analysis

6 Predict whether sodium metal will react more violently with iodine atoms or chlorine atoms based on your knowledge of the trends in electron affinity.

7 Organise the following elements in order of increasing electron affinity:
O, Cl, S, Ar

8 Analyse the elements in period 2.

Describe the changes in the following properties across the period.

a the sizes of atoms
b metallic character
c electronegativity

9 **a** Identify the most reactive non-metal from the following list of atoms.
magnesium, sulfur, chlorine, fluorine, aluminium, oxygen

b Explain your reasoning.

4.4 Introduction to bonding

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recognise that bonding is caused by the electrostatic attraction between oppositely charged species
- recognise that the position of the elements in the periodic table determines the nature of the chemical bonds formed
- recognise that ionic bonds are formed by the combination of a metal and a non-metal
- recognise that covalent bonds are formed by the combination of two non-metals
- recognise that metallic bonds are formed by the combination of two metals.

When elements combine to form a compound, there are three major types of bonding that can result.

- When two non-metals combine the electrons are shared in a covalent bond.
- When two metals combine, a 'sea' of **delocalised** electrons is produced, forming a metallic bond.
- When a metal and a non-metal combine, the electrons are transferred from one atom to the other, producing charged ions, which create a strong electrostatic attraction in the form of an ionic bond.

WHEN ATOMS COMBINE

In Module 4.3, you learnt that metals were to the left of the solid line in the periodic table with non-metals to the right of the line (Figure 4.3.1 on page 93).

Covalent bonding happens when non-metals combine with other non-metals, resulting in the formation of groups of atoms known as molecules. While non-metals have the ability to remove electrons from metals, they cannot do this to other non-metals. Instead, they tend to share electrons in **covalent bonds**, resulting in the formation of neutral molecules. While electrons are in constant motion and not fixed in position, they tend to exist between the neighbouring nuclei. Covalent bonds happen when two non-metals share one or more pairs of valence electrons. This creates an electrostatic attraction between the two nuclei that hold the atoms together in a covalent bond (Figure 4.4.1). As hydrogen is a non-metal, when it is combined with another non-metal it will also produce a covalent bond. Non-metals only share enough electrons to fill their outer shell or to have eight electrons in it. For example, nitrogen has five valence electrons so must pair up three of its electrons with electrons from other non-metallic elements. Covalent bonding will be covered in greater depth in Chapter 8.

A covalent bond results from the sharing of electrons between two non-metallic atoms.

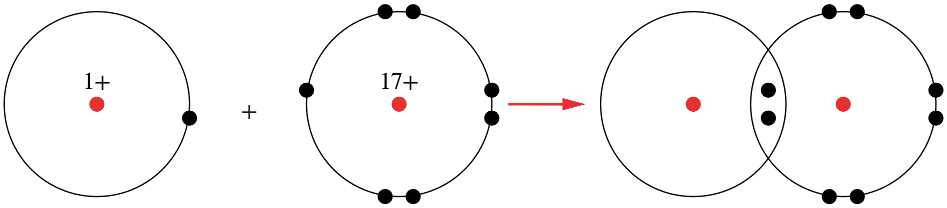
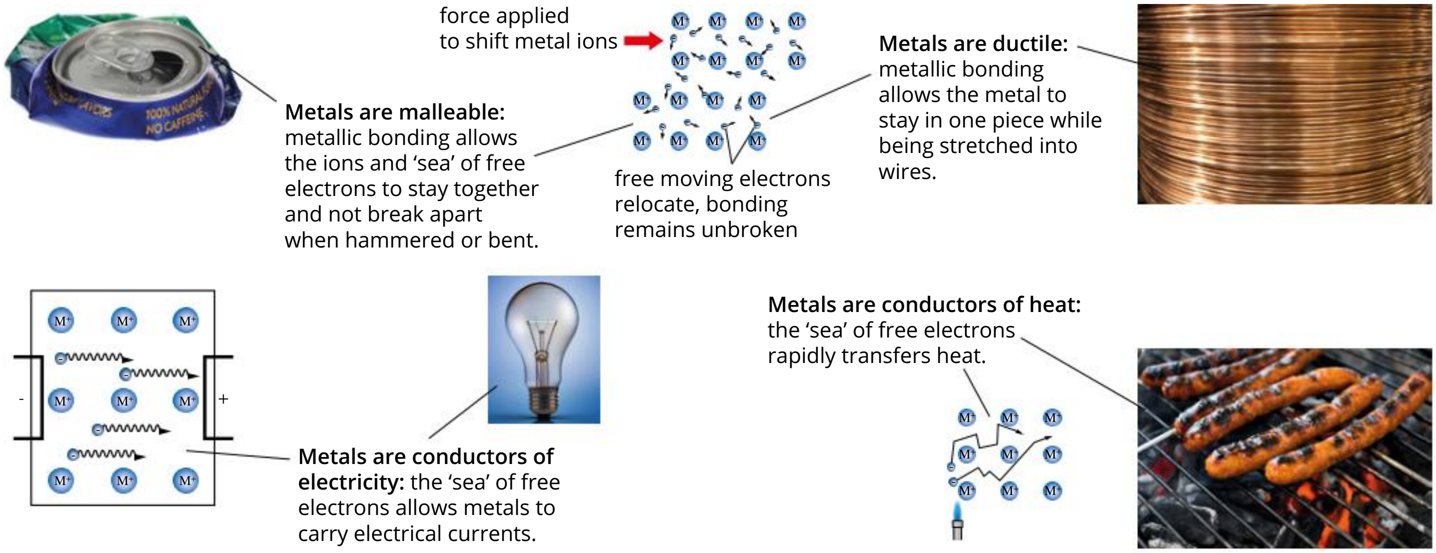


FIGURE 4.4.1 A pair of electrons is shared in the formation of a molecule of HCl.

When metal atoms combine with each other, they form **metallic bonds**. Metal atoms have a weak hold on their valence electrons. This gives the outer electrons the freedom to move throughout the metal without being bound to any one atom. In this case, each metal atom donates its valence electrons to create a mobile 'sea' of electrons that are not localised on a particular atom but are able to flow along the whole metal lattice structure. The free movement of electrons is the reason for the typical characteristics of metals such as conductivity, malleability and lustrousness, as shown in Figure 4.4.2. The attraction between the metal cations and the negatively charged electrons holds the metal together. A pure metal is where only one type of metal atom is present; however, if more than one type of metal atom is present then an alloy is formed. Metallic bonding will be covered in greater detail in Chapter 6.



Metals are malleable: metallic bonding allows the ions and 'sea' of free electrons to stay together and not break apart when hammered or bent.

free moving electrons relocate, bonding remains unbroken

Metals are ductile: metallic bonding allows the metal to stay in one piece while being stretched into wires.

Metals are conductors of electricity: the 'sea' of free electrons allows metals to carry electrical currents.

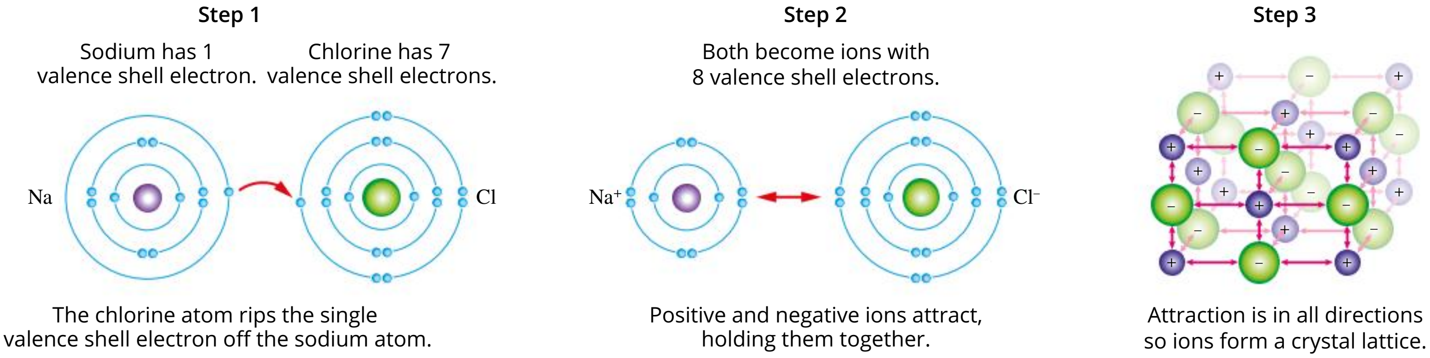
Metals are conductors of heat: the 'sea' of free electrons rapidly transfers heat.

FIGURE 4.4.2 Each metal cation is attracted to a 'sea' of delocalised valence electrons released from all metal atoms within the lattice. This mutual attraction bonds the metal together.

ℹ A metallic bond is the attraction of the stationary metal cations to the surrounding delocalised electrons.

In Module 4.2, you learnt that metals have low ionisation energies meaning that they lose electrons relatively easily to form positively charged cations. Non-metals have high ionisation energies and high electron affinities, so have a tendency to gain electrons relatively easily, forming negatively charged anions.

Therefore, when a metal and non-metal are combined the electrons are transferred from the metal to the non-metal to form a cation and anion. Each ion is surrounded by ions of the opposite charge, building up a three-dimensional structure called a lattice. The process is shown in Figure 4.4.3.



Step 1

Sodium has 1 valence shell electron. Chlorine has 7 valence shell electrons.

The chlorine atom rips the single valence shell electron off the sodium atom.

Step 2

Both become ions with 8 valence shell electrons.

Positive and negative ions attract, holding them together.

Step 3

Attraction is in all directions so ions form a crystal lattice.

FIGURE 4.4.3 Ionic bonding holds table salt (sodium chloride, NaCl) together.

Elements in the same group tend to lose or gain the same number of electrons when they react. This means they form ions of the same charge. For example, every element in group 1 forms an ion carrying a +1 charge when it bonds with a non-metal. Likewise, the elements of group 17 gain a single electron to get a −1 charge. The oppositely charged ions create an electrostatic attraction known as an ionic bond. Ionic bonding is covered in more detail in Chapter 7.

An ionic bond results from the electrostatic attraction between two oppositely charged ions.

This is a slightly simplistic description of the situation. Pure ionic bonding does not exist since all ionic compounds have a degree of covalent bonding. A bond is termed 'ionic' if the ionic character is greater than the covalent character.

Electronegativity was defined in Module 4.2 as the ability of an atom to attract electrons towards itself. The larger the difference in electronegativity between the two atoms involved in the bond, the more ionic the bond is. Bonds with partial ionic and covalent character are called polar covalent bonds (polarity will be discussed in Module 12.2). Bonds in which the electronegativity of the neighbouring atoms is less than 0.5 are termed 'covalent'. Table 4.4.1 shows the relationship between the difference in electronegativity and the resulting bond formed.

TABLE 4.4.1 Electronegativity difference and type of bond formed

Electronegativity difference	Type of bond formed	Examples (electronegativity difference)
>1.7	ionic bond	NaCl (2.3), MgO, (3.1)
1.7–0.5	polar covalent	O–H (1.2), ICl (0.5)
<0.5	covalent	C–H (0.4), Cl_2 (0)

In summary, you can use the periodic table to determine whether the substances are composed of metals or non-metals. By doing so, it is possible to predict the type of bonding present in the substance.

WS
1.1.5

Ionic bonding occurs between a metal and a non-metal.
Covalent bonding occurs between two non-metals.
Metallic bonding occurs between metals.

4.4 Review

SUMMARY

- Metals are located to the left in the periodic table.
- Non-metals are located to the right in the periodic table.
- Covalent bonds involve two non-metals.
- Ionic bonds involve a non-metal and a metal.
- Metallic bonds (including alloys) involve two metals.
- An ionic bond is formed through an electrostatic attraction between two oppositely charged ions.
- A covalent bond involves a pair of electrons being shared between atoms.
- Metals have a 'sea' of delocalised electrons.

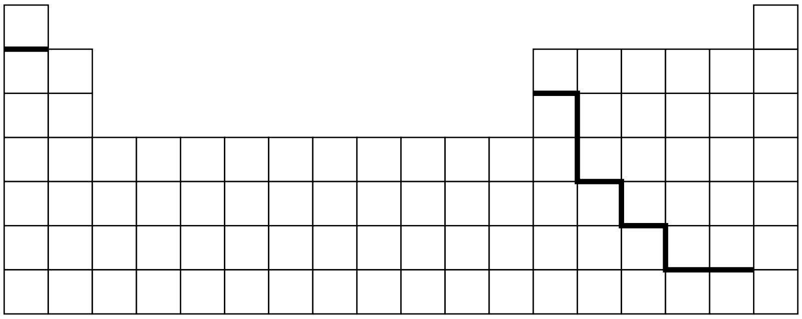
KEY QUESTIONS

Retrieval

1 Define the term 'metalloid'.

Comprehension

2 Use the blank periodic table below to identify and label the regions of metals, non-metals and metalloids.



Analysis

3 Classify the bond between the following pairs of atoms as covalent, ionic or metallic using the periodic table (Figure 4.3.1 on page 93).

- **a** sodium and chlorine atoms
- **b** magnesium and aluminium atoms
- **c** nitrogen and oxygen atoms
- **d** aluminium and sulfur atoms
- **e** hydrogen and lithium atoms
- **f** fluorine and chlorine atoms
- **g** two sulfur atoms

4 Classify the bonding as covalent or ionic for the following pairs of elements using the electronegativity values in Figure 4.2.5 on page 84.

- **a** carbon and nitrogen
- **b** aluminium and oxygen
- **c** silicon and iodine

5 Apply your understanding of electronegativity to distinguish between the three types of intramolecular bonds.

Chapter review

04

KEY TERMS

alkali metal
alkaline earth metal
amphoteric
anion
atomic number
atomic radius
block (periodic table)
cation
competing effects
covalent bond
delocalised
effective nuclear charge
electron affinity
electron configuration
electronegativity
electrostatic attraction
first ionisation energy
group (periodic table)
halogen
ionic bond
ionic radius
ionisation
isoelectronic
main group element
metallic bond
metallic character
metalloid
noble gas
octet arrangement
period (periodic table)
periodic table
periodic trend
periodicity
reactivity
shielding effect
transition metal
valence electron

KEY QUESTIONS

Retrieval

1 Select the response that best describes how the elements in the periodic table are arranged.
- **A** by the number of neutrons in their nucleus
- **B** by the number of protons in their nucleus
- **C** by increasing mass number
- **D** by increasing atomic mass

2 Identify the element whose properties are most similar to sodium.
- **A** magnesium
- **B** aluminium
- **C** potassium
- **D** hydrogen
- **E** silicon

3 Identify the response that correctly shows increasing radius of the following species.
- **A** Cl^-, Ar, K^+
- **B** K^+, Ar, Cl^-
- **C** Cl^-, K^+, Ar
- **D** Ar, Cl^-, K^+

4 Identify the most reactive non-metal from the following elements.
- **A** caesium
- **B** chlorine
- **C** fluorine
- **D** oxygen
- **E** sulfur

5 Use the periodic table in Figure 4.1.1 on page 76 to determine the period and block of the following elements.
- **a** hydrogen
- **b** carbon
- **c** phosphorus
- **d** copper
- **e** uranium

6 State the symbol for the element iron.

7 State the name of the element with the symbol K.

8 Identify the number of valence electrons in an atom of germanium.

9 Identify the effective nuclear charge of nitrogen.

10 Describe the changes that occur in the metallic character of the elements in period 2 of the periodic table.

11 Identify whether the following oxides are acidic, basic or amphoteric.
- **a** aluminium oxide
- **b** sodium oxide
- **c** sulfur dioxide

12 Select the appropriate word from the following options to complete the sentences below: increases, decreases
- **a** The force of attraction between the nucleus and valence electrons __________ in a period from left to right.
- **b** Atomic radii of elements __________ in a period from left to right.
- **c** Atomic radii of elements __________ in a group from top to bottom.
- **d** Metallic character of elements __________ from top to bottom in a group.

Comprehension

13 Nitrogen is found in period 2 and group 15 of the periodic table.
- **a** Explain the difference between the terms ‘period’ and ‘group’.
- **b** State the electron arrangement of nitrogen and explain why it is found in period 2 and group 15 of the periodic table.

14 Determine the period and group of the elements with the following electron configurations.

- **a** $1s^{2}2s^{2}$
- **b** $1s^{2}2s^{2}2p^{6}3s^{2}3p^{2}$
- **c** $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{10}4s^{2}4p^{3}$

15 Write an equation, including state symbols, for the first ionisation energy of aluminium.

16 Explain each of the following briefly.

- **a** The atomic radius of chlorine is smaller than that of sodium.
- **b** The first ionisation energy of fluorine is higher than that of lithium.
- **c** The reactivity of Be is less than that of Ba.
- **d** There are two groups in the *s*-block of the periodic table.

17 Explain why the ionic radius of an element across a period decreases.

18 Explain how you would expect each of the following properties to change based on the concepts of effective nuclear charge, shielding and size.

- **a** electronegativity down a group
- **b** atomic radius across a period
- **c** the radius of an anion compared to its parent atom
- **d** the first ionisation energy going down a group

19 Determine which element in each pair has the lower electron affinity.

- **a** Li or O
- **b** Cl or Na
- **c** Ca or K
- **d** Al or F

Analysis

20 Identify the correct chemical symbol for each of the following using the periodic table in Figure 4.1.1 on page 76.

- **a** the element that is in group 2 and period 4
- **b** a noble gas with exactly three occupied electron shells
- **c** an element from group 14 that is a non-metal
- **d** an element that has exactly three occupied electron shells and is in the *s*-block
- **e** the element in period 2 that has the largest atomic radius
- **f** the element in group 15 that has the highest first ionisation energy
- **g** the element in period 2 with the highest electronegativity

21 Since the chemical properties of the elements are determined by the number of valence electrons of the element, organise the following 23 elements into appropriate groups of the periodic table.

- third element in group 15; second element in period 5; sixth element in group 17
- elements with the following atomic symbols: C, Ca, Ga, Sb, Te, Rn
- elements with the following atomic numbers: 3, 12, 13, 19, 36, 53, 56
- the lightest noble gas; the third heaviest halogen; the heaviest alkali metal
- elements with the following electron configurations:
  - $1s^{2}2s^{2}2p^{6}3s^{2}3p^{4}$
  - $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{10}4s^{2}4p^{2}$
  - $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{10}4s^{2}4p^{6}4d^{10}5p^{2}5d^{10}6s^{1}$
  - $1s^{2}2s^{2}2p^{1}$

22 Determine the electron configuration of the following elements using the periodic table in Figure 4.1.1 on page 76.

- **a** phosphorus
- **b** Mg
- **c** element in period 4, group 14
- **d** element in period 2, group 16

23 Identify the group in the periodic table in which each of the following elements belongs. Use the periodic table in Figure 4.1.1 on page 76 as a reference.

- **a** B
- **b** Cl
- **c** Na
- **d** Ar
- **e** Si
- **f** Pb

24 Compare the following pairs of elements to predict which requires more energy to remove a valence electron.

- **a** phosphorus or magnesium
- **b** fluorine or iodine

25 Organise the following atoms in increasing atomic radius: Mg, Na, K, Ca.

26 Organise the following ions in decreasing ionic radius: Cl^{-}, K^{+}, Ca^{2+}, S^{2-}.

27 Sort the following elements in order of increasing first ionisation energy: Mg, F, Na, K, Ca.

28 Organise the following elements in order of increasing electron affinity: Br, S, K, Al.

29 Determine which element has the largest first ionisation energy from the following sets of elements.

- **a** phosphorus, arsenic, nitrogen
- **b** sodium, potassium, calcium
- **c** silicon, chlorine, sulfur

30 Identify which of the following pairs of elements would react together most vigorously.

Li and Cl_2; Li and Br_2; K and Cl_2; K and Br_2

31 **a** Sort the following elements from least reactive to most reactive: rubidium, sodium, lithium, potassium.

b Explain your reasoning.

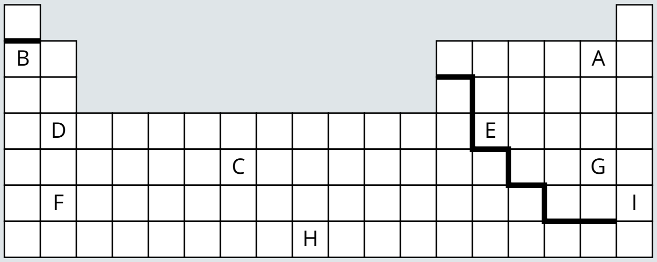
32 Predict the first ionisation energy of xenon using the information about first ionisation energies and atomic number of the noble gases in the following table.

Atomic number	Ionisation energy (kJ mol^{-1})
2	2372
10	2081
18	1521
36	1351
54	

33 Predict the electron affinity of oxygen using the information about electron affinities of group 16 and 17 elements listed in the following table.

Atomic number	Electron affinity (kJ mol^{-1})
8	
16	200
17	349
34	195
35	324
52	190
53	195
9	328

34 Identify the elements that meet the criteria below using the following hypothetical periodic table.



a have the same number of valence electrons

b are in the same period

c has smallest atomic radius

d has smallest ionic radius

e has lowest electronegativity

f is a noble gas

g is classified as a non-metal

h is classified as a metalloid

Knowledge utilisation

35 **a** Examine the data in the following table and explain the trend in electron affinity observed for the period 2 elements.

b Identify any exceptions to the trend and explain them based on the electron configuration of the elements.

Element	Electron affinity (kJ mol^{-1})
Li	60
Be	–50
B	27
C	122
N	0
O	141
F	328
Ne	–120

36 Investigate the relationship between the metallic character of an element and its ionisation energy.

Based on your findings, explain the variation of metallic character and position of the element in the periodic table.

CHAPTER

05 Analytical techniques

Think about what you had for breakfast today. The packaging on any processed food you ate, including cereal, milk, bread and yoghurt, has a label that provides you with nutritional information so you can make healthy diet choices. This information comes from chemical analysis.

Analytical chemists are employed in a range of industries, from food manufacturing to environmental monitoring and forensic sciences and many more. They use a range of techniques to determine the identity, quality and quantity of the chemicals that make up our world.

Analytical techniques allow us to characterise qualitatively or quantitatively the composition of any material and its chemical state. In this chapter, you will look at a number of spectroscopic techniques that are based on the interaction of radiation with a material.

The various spectroscopic techniques provide us with information about:

- the type of atom that is present (qualitative analysis)
- how much of a particular atom is present (quantitative analysis).

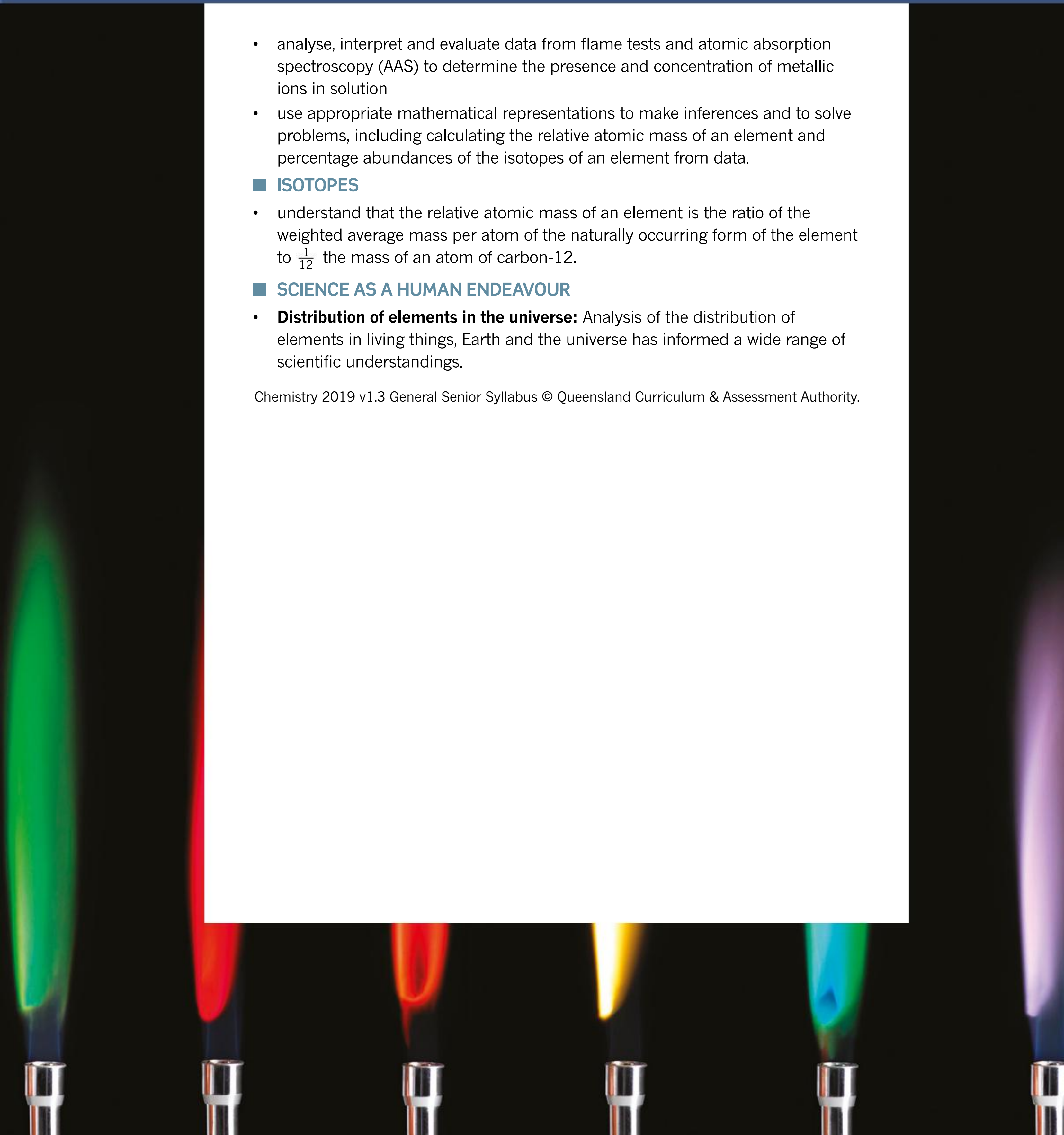
In each of the spectroscopic techniques we will look at in this chapter, the atom absorbs a specific quantum of energy, which causes electrons to move to higher energy levels. These energy levels are quantised; that is, they have certain fixed values. It is this characteristic that allows scientists to determine the identity of the atom present.

Syllabus subject matter

Topic 1 • Properties and structure of atoms

ANALYTICAL TECHNIQUES

- understand that mass spectrometry involves the ionisation of substances and the separation and detection of the resulting ions, and that the spectra generated can be analysed to determine the isotopic composition of elements, and interpreted to determine relative atomic mass (analysis to determine)
- understand that flame tests and atomic absorption spectroscopy (AAS) are analytical techniques that can be used to identify elements; these methods rely on electron transfer between atomic energy levels and are shown by line spectra
- distinguish between absorption and emission spectra and recognise that the emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels (Bohr model), which converge at higher energies. Explain that emission spectra are produced when photons are emitted from atoms when excited electrons return to a lower energy level



- analyse, interpret and evaluate data from flame tests and atomic absorption spectroscopy (AAS) to determine the presence and concentration of metallic ions in solution
- use appropriate mathematical representations to make inferences and to solve problems, including calculating the relative atomic mass of an element and percentage abundances of the isotopes of an element from data.

ISOTOPES

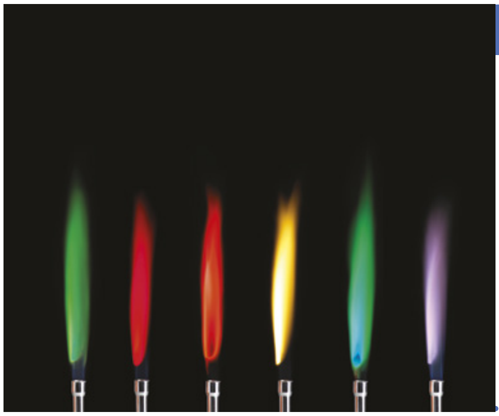
- understand that the relative atomic mass of an element is the ratio of the weighted average mass per atom of the naturally occurring form of the element to $\frac{1}{12}$ the mass of an atom of carbon-12.

SCIENCE AS A HUMAN ENDEAVOUR

- **Distribution of elements in the universe:** Analysis of the distribution of elements in living things, Earth and the universe has informed a wide range of scientific understandings.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

5.1 Mass spectrometry of atoms



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- ➤ understand that mass spectrometry involves the ionisation of substances and the separation and detection of the resulting ions, and that the spectra generated can be analysed to determine the isotopic composition of elements, and interpreted to determine the relative atomic mass (analysis to determine)
- ➤ use appropriate mathematical representations to make inferences and to solve problems, including calculating the relative atomic mass of an element and percentage abundances of the isotopes of an element from data.

Mass spectrometry is one of the most common and useful analytical tools, as it can be used on its own or in combination with other techniques.

The mass spectrometer can be used for:

- quantitative analysis—as a sophisticated and very sensitive detector that can measure how much of a substance is present
- qualitative analysis—to provide a unique fingerprint of a substance; this can be used to identify the nature of the substance from an online database or to give information about the nature of the elements present or structure of a new or unknown compound.

All of this information can be determined from the same analysis. This module will focus on the elemental analysis using mass spectrometry, as analysis of molecules is outside the scope of this unit.

MASSES OF PARTICLES

Since atoms are so small, it is not possible to determine their mass without specialised equipment; therefore, scientists categorise the mass of the atoms in relation to a mole of the particles. A mole represents a fixed number of particles, 6.02×10^{23} to be precise. This will be covered in more detail in Chapter 9, but in this module you will become familiar with the concepts of atomic mass and isotopic mass.

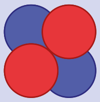
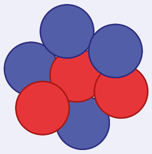
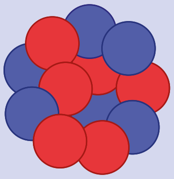
Relative masses

From Chapter 3, you will remember that isotopes are atoms of the same element that have different numbers of neutrons in their nucleus. So isotopes have the same atomic number but a different mass number.

Relative isotopic masses

The mass of the individual isotopes of each element is known as the **relative isotopic mass**. Table 5.1.1 shows the approximate mass of various isotopes relative to the carbon-12 isotope taken as 12 units exactly.

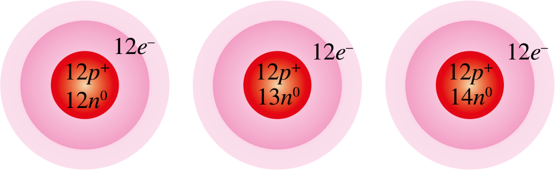
TABLE 5.1.1 Approximate mass of various isotopes relative to the ^{12}C isotope, which is taken as exactly 12 units

Isotope	Diagram of nucleus	Number of protons in nucleus	Number of neutrons in nucleus	Total number of protons and neutrons	Approximate mass of atom relative to ^{12}C isotope
hydrogen $^{1}_{1}H$		1	0	1	$\frac{1}{12}$ of 12 = 1
helium $^{4}_{2}He$		2	2	4	$\frac{4}{12}$ of 12 = 4
lithium $^{7}_{3}Li$		3	4	7	$\frac{7}{12}$ of 12 = 7
carbon $^{12}_{6}C$		6	6	12	12

As shown in Figure 5.1.1, there are three different isotopes of the element magnesium:

- $^{24}_{12}Mg$, which contains 12 protons and 12 neutrons
- $^{25}_{12}Mg$, which contains 12 protons and 13 neutrons
- $^{26}_{12}Mg$, which contains 12 protons and 14 neutrons.

Isotopes of magnesium



atomic symbol	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{26}_{12}Mg$
number of protons	12	12	12
number of electrons	12	12	12
mass number	24	25	26
number of neutrons	12	13	14
isotope notation	Mg-24	Mg-25	Mg-26

FIGURE 5.1.1 The three isotopes of the element magnesium. Different numbers of neutrons in the nuclei of these atoms give the atoms different masses.

The isotopes have different masses because they have different numbers of neutrons. Since the mass of an electron is relatively small, the mass of an atom is mainly determined by the mass of the protons and neutrons in its nucleus.

The relative isotopic masses of the three magnesium isotopes are experimentally determined to be 78.99% (^{24}Mg), 10% (^{25}Mg) and 11.01% (^{26}Mg). Since the masses of a proton and a neutron are similar and close to 1 on the $^{12}C = 12$ scale, the relative isotopic mass of an isotope is almost, but not exactly, equal to the number of protons plus neutrons in the nucleus.

The relative isotopic mass of an isotope is the mass of an atom of that isotope relative to the mass of an atom of carbon-12 (^{12}C) taken as 12 units exactly.

Relative isotopic abundance

Naturally occurring magnesium is made up of the three isotopes, shown in Figure 5.1.1, 78.99% of the lighter isotope and 10% and 11.01% of the two heavier isotopes. This composition is virtually the same no matter the source of the magnesium. The percentage abundance of an isotope in the natural environment is called its **relative isotopic abundance**. Most elements, like magnesium, are a mixture of two or more isotopes. Details of the isotopes of some common elements are shown in Table 5.1.2.

TABLE 5.1.2 Isotopic composition of some common elements

Element	Isotopes	Relative isotopic mass	Relative isotopic abundance (%)
hydrogen	^{1}H	1.008	99.986
	^{2}H	2.014	0.014
	^{3}H	3.016	0.0001
oxygen	^{16}O	15.995	99.76
	^{17}O	16.999	0.04
	^{18}O	17.999	0.20
silver	^{107}Ag	106.9	51.8
	^{109}Ag	108.9	48.2

THE MASS SPECTROMETER

Relative isotopic masses of elements and their isotopic abundances are determined by using an instrument called a **mass spectrometer**, which was invented by Francis Aston in 1919.

A mass spectrometer separates the individual isotopes in a sample of an element and determines the mass of each isotope, relative to the carbon-12 isotope, and the relative abundances of the isotopes.

Operation of a mass spectrometer

There are a number of different types of mass spectrometers; however, they work on similar basic principles. A typical mass spectrometer is shown in Figure 5.1.2 and comprises several stages. While you are not required to learn how the mass spectrometer operates, it is important to understand these basic steps.

1. The sample is vaporised and the resultant gas is ionised to form positive ions.
2. The ions are then accelerated by an electric field.
3. These fast-moving ions are then separated by a magnetic field based on their mass-to-charge ratio.
4. The ions of a certain mass and charge reach a detector, which measures the current produced by a particular ion.

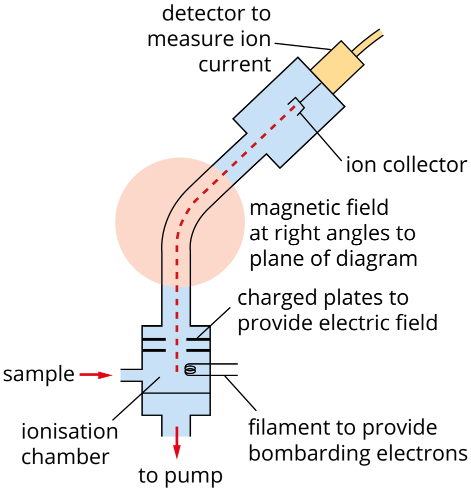


FIGURE 5.1.2 The construction of a mass spectrometer. The sample is ionised and the ions pass through an electric field and a magnetic field before reaching the detector.

Mass spectra

The output obtained from a mass spectrometer is called a **mass spectrum**. Figure 5.1.3 shows the mass spectrum of magnesium.

- The number of peaks indicates the number of isotopes; in this case, three.
- The horizontal axis indicates the relative mass of each isotope present in an element according to the isotope's mass-to-charge ratio, which is given the symbol *m*/*z* or *m*/*e*. (The charge on most ions reaching the detector is +1, so the mass of an isotope can be read directly from the horizontal axis.) The three isotopes of magnesium have relative masses of 24, 25 and 26.
- The vertical axis indicates the abundance of each isotope in the sample: 79%, 10% and 11%.

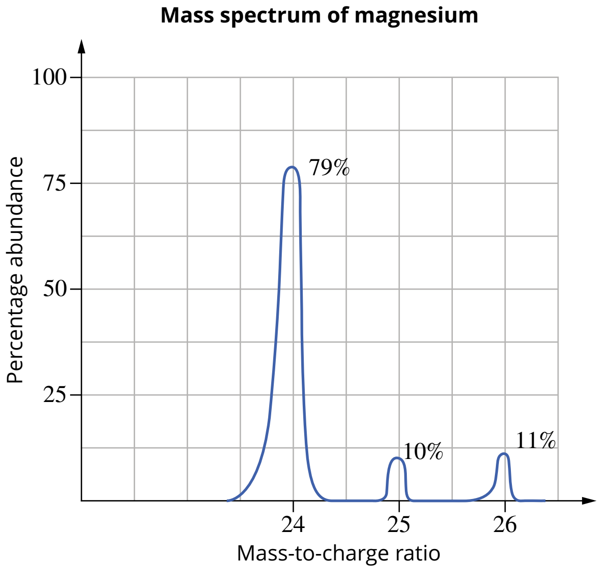


FIGURE 5.1.3 The mass spectrum of magnesium. The three peaks indicate there are three isotopes. The most abundant isotope has a relative isotopic mass of 24 and an abundance of 79%. The other isotopes have relative isotopic masses of approximately 25 and 26 and abundances of approximately 10% and 11% respectively.

WS 1.1.6

In summary, in the mass spectrum of an element, the:

- number of peaks indicates the number of isotopes
- position of each peak on the horizontal axis indicates the relative isotopic mass
- relative heights of the peaks correspond to the relative abundance of the isotopes.

RELATIVE ATOMIC MASS

Most elements consist of a mixture of isotopes. For the purpose of the calculations you will be doing later in this chapter, it is convenient to know the average relative mass of an atom in this mixture. This average is called the **relative atomic mass** of an element, and given the symbol A_r.

In order to calculate the average of the relative masses of the isotopes that exist in a naturally occurring mixture of an element, you must consider the relative abundances of each isotope. Data obtained from the mass spectrum of magnesium shown in Figure 5.1.3 is summarised in Table 5.1.3.

The relative atomic mass of an element is the weighted average of the relative masses of the isotopes of the element on the ^{12}C scale.

TABLE 5.1.3 The isotopic composition of magnesium

Isotope	Relative isotopic mass	% abundance
^{24}Mg	23.985	78.99
^{25}Mg	24.986	10.00
^{26}Mg	25.983	11.01

The data about the three isotopes is used to calculate the relative atomic mass of magnesium. A weighted average mass is calculated by using the relative isotopic masses and abundances to find the total mass of 100 atoms. This mass is then divided by 100 to find the average mass of one atom.

The relative atomic mass (A_r) of an element with three isotopes can be calculated using the formula:

$$A_r = \frac{(\text{relative isotopic mass} \times \%\ \text{isotopic abundance})}{100} + \frac{(\text{relative isotopic mass} \times \%\ \text{isotopic abundance})}{100} + \frac{(\text{relative isotopic mass} \times \%\ \text{isotopic abundance})}{100}$$

The periodic table on the inside cover at the back of the book provides relative atomic masses, calculated by taking into account the relative abundances of all the natural isotopes of each element.

Worked example 5.1.1

CALCULATING RELATIVE ATOMIC MASS FROM ISOTOPIC MASSES AND PERCENTAGE ABUNDANCES

Determine the relative atomic mass of magnesium from the data obtained from Table 5.1.3.

Thinking	**Working**
Determine the relative isotopic mass and abundance of each isotope.	The table shows three isotopes. first isotope: relative isotopic mass 23.985; abundance 78.99% second isotope: relative isotopic mass 24.986; abundance 10.00% third isotope: relative isotopic mass 25.983; abundance 11.01%
Substitute the relative isotopic masses and abundances into the formula for calculating relative atomic mass: $A_r = \frac{(\text{relative isotopic mass} \times \% \text{ isotopic abundance})}{100} + \frac{(\text{relative isotopic mass} \times \% \text{ isotopic abundance})}{100} + \frac{(\text{relative isotopic mass} \times \% \text{ isotopic abundance})}{100}$	$A_r = \frac{(23.985 \times 78.99) + (24.986 \times 10.00) + (25.983 \times 11.01)}{100}$
Calculate the relative atomic mass.	$A_r = \frac{1894.58 + 249.86 + 286.07}{100}$ $= 24.3051$
Express the answer to 2 decimal places.	A_r(Mg) = 24.31

➤ Try yourself 5.1.1

CALCULATING RELATIVE ATOMIC MASS FROM ISOTOPIC MASSES AND PERCENTAGE ABUNDANCES

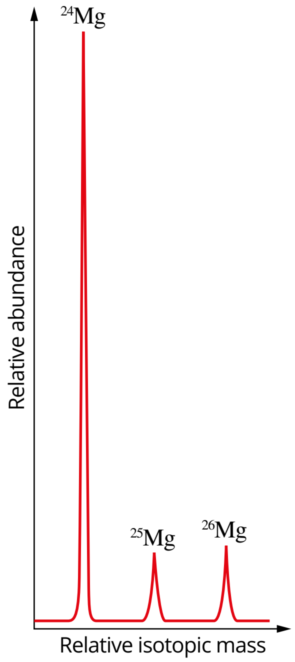
Boron has two isotopes. Their relative isotopic masses and percentage abundances are provided. Calculate the relative atomic mass of boron.

Isotope	**Relative isotopic mass**	**% abundance**
^{10}B	10.013	19.91
^{11}B	11.009	80.09

Worked example 5.1.2

CALCULATING PERCENTAGE ABUNDANCE OF EACH ISOTOPE FROM THE MASS SPECTRUM

Using the simplified mass spectrum of magnesium shown, calculate the percentage abundance of each of its three isotopes.

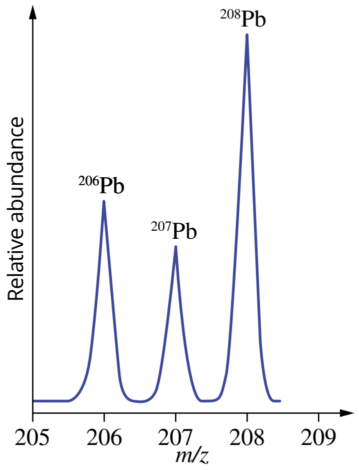


Thinking	**Working**
Measure the peak height for each isotope using a ruler.	From the spectrum, the height of each peak is: $^{24}\text{Mg} = 7.9\,\text{cm}$ $^{25}\text{Mg} = 1.0\,\text{cm}$ $^{26}\text{Mg} = 1.1\,\text{cm}$
Calculate the total peak height for the three isotopes by adding the individual peak heights.	$\text{Total peak height} = 7.9 + 1.0 + 1.1$ $= 10\,\text{cm}$
Substitute the peak height for each isotope into the formula.	$\%\ \text{abundance} = \frac{7.9}{10} \times 100 = 79\%$ $\%\ \text{abundance} = \frac{1.0}{10} \times 100 = 10\%$ $\%\ \text{abundance} = \frac{1.1}{10} \times 100 = 11\%$

➤ Try yourself 5.1.2

CALCULATING PERCENTAGE ABUNDANCE OF EACH ISOTOPE FROM THE MASS SPECTRUM

Using the simplified mass spectrum of lead shown, calculate the percentage abundance of each of its three isotopes.



Worked example 5.1.3

CALCULATING PERCENTAGE ABUNDANCES FROM RELATIVE ATOMIC MASS AND THE ISOTOPIC MASSES

The relative atomic mass of rubidium is 85.47. The isotopic masses of its two isotopes are 84.95 and 86.94. Calculate the relative abundances of the isotopes in naturally occurring rubidium.

Thinking	**Working**
State the relative abundances of the isotopes in terms of *x*, where *x* is the abundance of the lighter isotope. Abundance of lighter isotope = *x*. The abundance of heavier isotope must equal 100 – *x*.	Abundance of 84.95 isotope = *x* Abundance of 86.94 isotope = 100 – *x*
Substitute the isotopic masses, relative abundances and relative atomic mass into the formula: $A_r = \frac{(\text{isotopic mass} \times \%\ \text{isotopic abundance})}{100} + \frac{(\text{isotopic mass} \times \%\ \text{isotopic abundance})}{100}$	$85.47 = \frac{84.95x + (86.94 \times (100 - x))}{100}$
Expand the top line of the equation.	$85.47 = \frac{84.95x + 8694 - 86.94x}{100}$
Solve the equation to find *x*, the relative abundance of the lightest isotope.	$8547 = 84.95x + 8694 - 86.94x$ $8547 - 8694 = 84.95x - 86.94x$ $-147 = -1.99x$ $x = 73.87\%$ $= 73.7\%$
Determine the abundance of the heavier isotope.	Abundance of 86.94 isotope $= 100 - x$ $= 100 - 73.87$ $= 26.13\%$

➤ Try yourself 5.1.3

CALCULATING PERCENTAGE ABUNDANCES FROM RELATIVE ATOMIC MASS AND THE ISOTOPIC MASSES

The relative atomic mass of copper is 63.54. The isotopic masses of its two isotopes are 62.95 and 64.95. Calculate the relative abundances of the isotopes in naturally occurring copper.

5.1 Review

SUMMARY

- Most elements consist of a mixture of isotopes.
- The most common isotope of carbon, carbon-12, is used as the reference standard to compare the masses of atoms.
- The carbon-12 isotope is assigned a mass of exactly 12 units.
- The relative isotopic mass of an isotope is the mass of an atom of the isotope relative to the mass of an atom of carbon-12 taken as 12 units exactly.
- The relative isotopic mass and relative abundance of isotopes can be measured using a mass spectrometer.
- The relative atomic mass, A_r, of an element is a weighted average of its isotopic masses.
- The relative molecular mass of molecules, or formula mass of ionic compounds, is calculated from the sum of the relative atomic masses of its constituent elements.
- Operation of mass spectrum:
  - Abundance of the isotopes is related to the relative heights of the peaks.
  - Number of isotopes is equal to the number of peaks in the mass spectra.

KEY QUESTIONS

Retrieval

1 **a** Recall what the atomic number of an element is.
b Recall whether the atomic number for atoms of the same element ever changes.

2 Define the term ‘isotope’.

3 Complete the following table for the following isotopes.

Isotope name	Isotope symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
		19	37			
oxygen-17			17			
uranium-235						
			238	92		
	$^{10}_{5}B$					
				5	6	

Comprehension

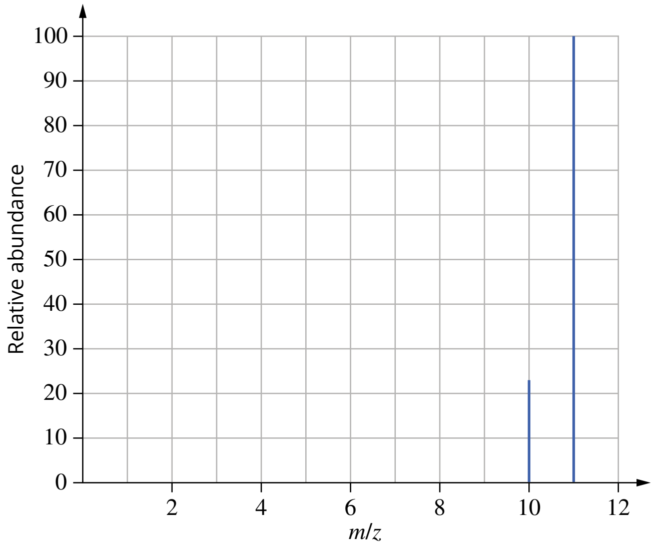
4 Explain why the relative atomic mass of an element shown in the periodic table is generally not a whole number.

5 Explain what is separated and detected by mass spectrometers.

Analysis

6 Identify the number of protons, electrons and neutrons in a neutral atom of ^{37}Cl.

7 Calculate the relative atomic mass of the element shown in the mass spectrum.



8 Calculate the number of protons, neutrons and electrons in each of the following ions. (Assume the most abundant isotope for each ion.)
a Cu^{2+} **b** P^{3-}

9 Determine which isotope of boron ($^{10}B^+$ or $^{11}B^+$) would travel furthest in a mass spectrometer.

10 Deduce the complete symbol for an atom with the following characteristics.
a contains 19 electrons and 20 neutrons
b nitrogen atom with 8 neutrons
c bromine atom with a mass number of 80
d mercury atom that contains 281 subatomic particles
e contains 20 protons and has a mass number of 42
f cobalt atom with 33 neutrons
g iodine atom with 181 subatomic particles
h sulfur atom with 52 subatomic particles
i uranium atom with 146 neutrons
j copper atom with 65 nucleons

5.1 Review *continued*

11 A mass spectrum identifies three isotopes of iodine with isotopic abundances of 80% ^{127}I, 17% ^{126}I and 3% ^{128}I. Calculate the average atomic mass of iodine.

12 Calculate the atomic mass of an element if 60.4% of the atoms have a mass of 68.9257 and the rest have a mass of 70.9249. Identify the element in the periodic table.

13 The table below includes some common elements and the relative abundance of a number of their isotopes.

Element	**Atomic formula**	**Isotopic mass**	**Relative isotopic abundance**
rubidium	^{85}Rb	84.911	72.17
	^{87}Rb	86.909	27.83
antimony	^{121}Sb	120.904	57.21
	^{123}Sb	122.904	42.79
strontium	^{84}Sr	83.913	0.56
	^{86}Sr	85.909	9.86
	^{87}Sr	86.909	7.00
	^{88}Sr	87.906	82.58
potassium	^{39}K	38.963	93.258
	^{40}K	39.964	0.0117
	^{41}K	40.962	6.7302
argon	^{36}Ar	35.968	0.3365
	^{38}Ar	37.963	0.0632
	^{40}Ar	39.962	99.6003

Calculate the relative atomic mass of the following elements using the information in the table.

a rubidium

b antimony

c strontium

14 Compare the relative atomic masses of argon and potassium using the isotopic data in the table in Question **13**. Determine if their relative positions in the periodic table would change if it were organised by mass instead of atomic number.

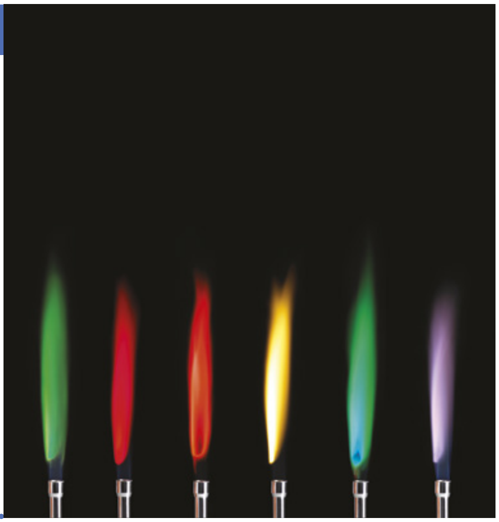
15 The relative atomic mass of gallium is 69.723. The relative isotopic masses of its two isotopes are 68.926 and 70.925. Calculate the relative abundances of the isotopes in naturally occurring gallium.

16 The relative atomic mass of bromine is 79.903. The relative isotopic masses of its two isotopes are 80.916 and 78.918. Calculate the relative abundances of the isotopes in naturally occurring bromine.

5.2 Atomic absorption spectroscopy

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that flame tests and atomic absorption spectroscopy (AAS) are analytical techniques that can be used to identify elements; these methods rely on electron transfer between atomic energy levels and are shown by line spectra
- distinguish between absorption and emission spectra and recognise that the emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels (Bohr model), which converge at higher energies. Explain that emission spectra are produced when photons are emitted from atoms when excited electrons return to a lower energy level
- analyse, interpret and evaluate data from flame tests and atomic absorption spectroscopy (AAS) to determine the presence and concentration of metallic ions in solution.



You will be familiar with the colourful effects that are created in fireworks displays. Small quantities of different metal compounds are incorporated in fireworks to produce different colours. For example, in Figure 5.2.1, strontium compounds have produced the eye-catching scarlet colour and sodium compounds have produced the yellow. If you have ever spilt salt into the flame of a gas stove, you might have noticed this same yellow colour in the flame.

In Chapter 3, you explored the atomic structure of atoms. You learnt that electrons are arranged in shells at different distances from the nucleus. This was related to the energy possessed by each electron. In this module, you will apply your understanding to the identification of the elemental composition based on the colour of the light emitted from the excited species. In addition, you will examine how the emission and absorption of light by metals can be used to determine the concentration of salts in solution.

FIGURE 5.2.1 Metal compounds incorporated into fireworks are responsible for the colours in this display.

SPECTROSCOPY

Light is a form of energy and is a type of **electromagnetic radiation**. Other forms of electromagnetic radiation are radio waves and X-rays. Visible light is only a small part of the range of different forms of electromagnetic radiation. The spread of the different types of radiation arranged according to their relative energies and wavelengths is referred to as the **electromagnetic spectrum** (Figure 5.2.2).

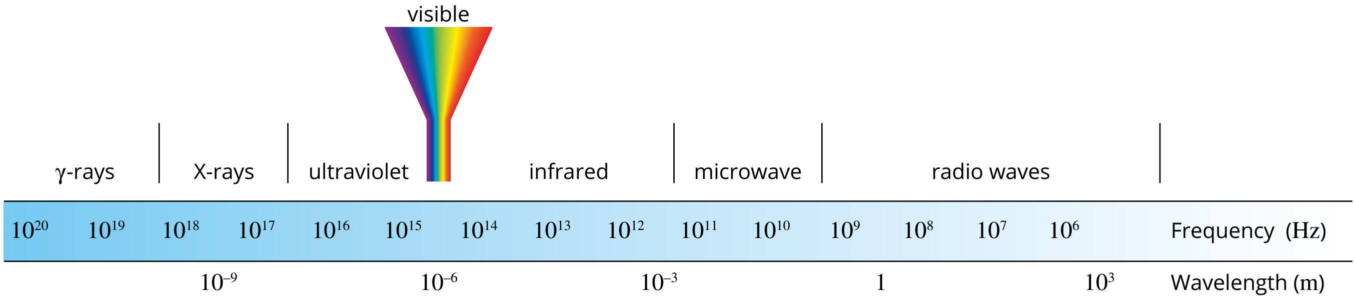


FIGURE 5.2.2 Visible light is only a small part of the electromagnetic spectrum. The analytical techniques detailed in this module use radiation within the visible spectrum.

The visible spectrum of light corresponds to the radiation with wavelengths between 400 and 700 nm on the electromagnetic spectrum.

Electromagnetic radiation, such as light, can interact with atoms, and the nature of this interaction depends upon the energy of the electromagnetic radiation. In this module, you will learn about an analytical technique called **spectroscopy**, which uses light and other radiation of the electromagnetic spectrum to give us information about the materials around us. The spectroscopic techniques that you will look at in this module will specifically deal with light within the visible region of the electromagnetic spectrum.

When a substance absorbs visible light, it appears coloured. The colour observed is not the same as the colour of the light absorbed. The colour you see is actually due to reflected or transmitted light. For example, plant leaves contain chlorophyll. When you look at the leaves, they appear green because chlorophyll absorbs light in the purple and red ranges of the spectrum. Chlorophyll does not absorb light in the green region of the spectrum, so this is reflected back into your eyes.

Spectroscopic techniques utilise the characteristics of atoms.

- Atoms absorb and emit electromagnetic radiation of specific energies.
- Atoms undergo a change when they absorb electromagnetic radiation.
- Different parts of the electromagnetic spectrum affect different parts of the atoms.

Radiation from each portion of the electromagnetic spectrum has a specific frequency, wavelength and energy associated with it. Ultraviolet light is very short wavelength radiation with high energy, while radio waves have long wavelengths and low energy.

The energy of the radiation determines which part of an atom is affected.

FLAME TEST

Chemists use the fact that some metals produce particular colours when they are heated as a convenient and simple method of analysis. The metallic elements present in a compound can often be determined just by inserting a sample of the compound into a non-luminous Bunsen burner flame, as shown in Figure 5.2.3.

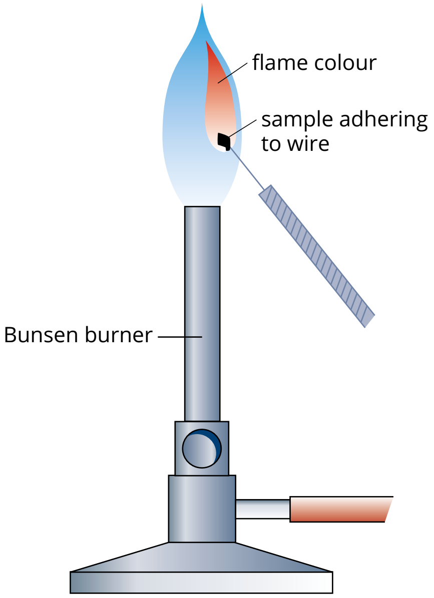


FIGURE 5.2.3 Performing a flame test. A moist wire has been dipped in the sample and then placed in the flame. A fine spray of solution from a spray bottle could be used instead.

Each metal ion produces a characteristic colour. This means that the metal in an unknown sample can be identified by comparing the flame colour with the known characteristic colours produced by metals. Some examples of the flame colours produced by metals can be seen in Figure 5.2.4.

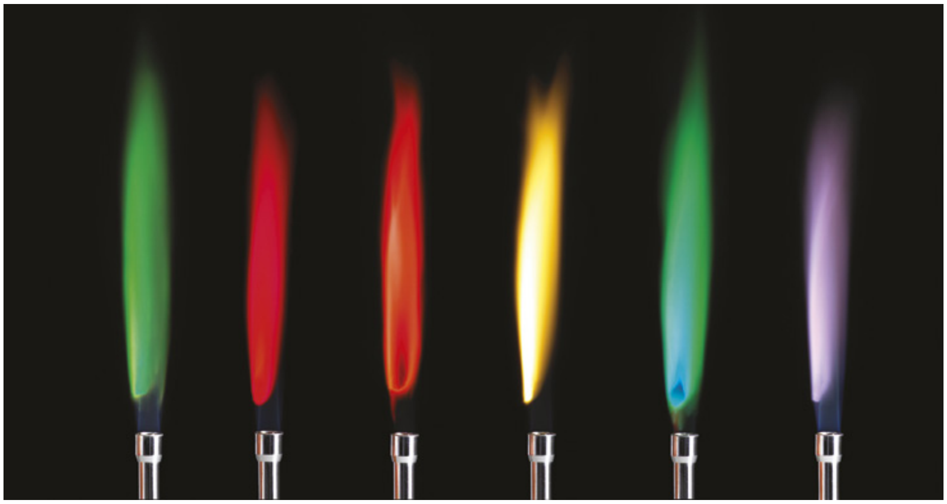


FIGURE 5.2.4 The colour of the flame is determined by the different metal compounds present and can be used to identify these metals. The flame colours shown here are for (from left to right) barium (yellow–green), lithium (crimson), strontium (scarlet), sodium (yellow), copper (green) and potassium (lilac).

WS 1.1.7

PA 1.1.1

Because each flame colour is unique, you can use the flame colour produced to qualitatively identify a metal ion in an unknown sample. Table 5.2.1 shows some of the flame colours that can be used to identify the presence of metals in a flame test.

TABLE 5.2.1 Characteristic flame colours of some metal ions

Metal	Flame colour
sodium	yellow
strontium	scarlet
copper	green
barium	yellow–green
lithium	crimson
calcium	yellow–red
potassium	lilac

It is important to note that only a small number of metals produce flame colours by this process. Although the exact colour produced by each metal is unique, simple flame tests can result in a level of uncertainty when trying to decide between different shades of similar colours such as scarlet (strontium), crimson (lithium) and yellow–red (calcium).

You will recall from Chapter 3 that electrons are arranged in atoms in different energy levels. Electrons in energy levels close to the nucleus have the lowest energies and experience the strongest attraction to the nucleus. When placed in a flame, an electron can jump up to a higher energy level if it absorbs energy that corresponds exactly to the difference in energy between the lower energy level and the higher energy level (Figure 5.2.5).

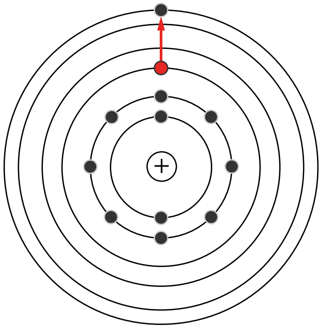


FIGURE 5.2.5 Energy from a flame can promote an electron in an atom to a higher energy level.

Atoms with electrons in higher energy levels are unstable, so the **excited electrons** quickly return to lower energy levels. Figure 5.2.6 shows that the energy absorbed by the electrons is emitted as visible light as they return to a lower energy level.

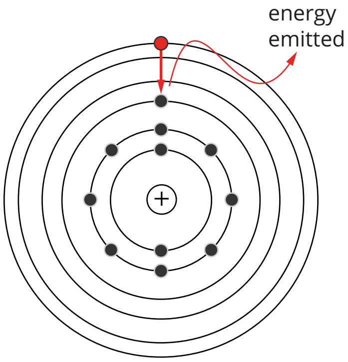


FIGURE 5.2.6 An excited electron quickly returns to a lower energy level, emitting electromagnetic energy in the form of visible light of a particular wavelength.

When electrons are located in the lowest energy shells possible, the electrons are in the **ground state**.

If the energy emitted when electrons move from the higher energy level back to their original level falls within the region of the visible light spectrum, light of a specific colour is emitted (Figure 5.2.7 on page 118). This emitted light is what you observe as the coloured flame in a flame test. The different colours observed correspond to specific wavelengths of light.

ℹ The colour of the light emitted from a sample is dependent on its energy. Red light has a lower energy than blue light.

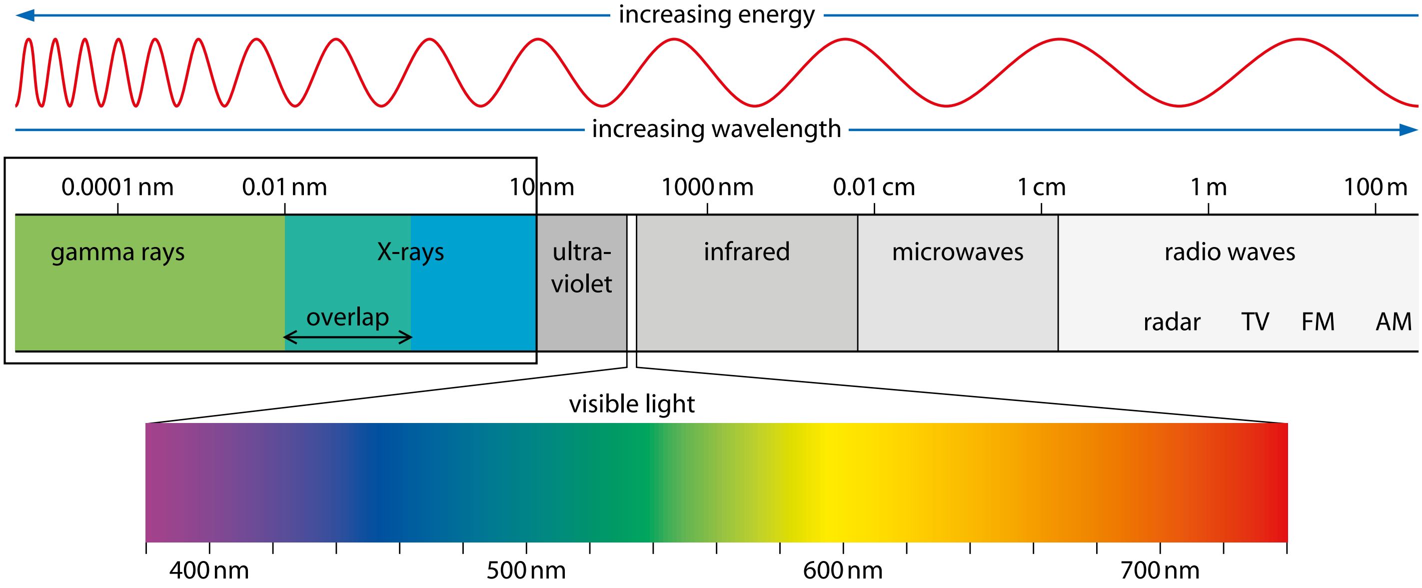


FIGURE 5.2.7 When excited electrons return to a ground state from a flame test, the emitted energy falls within the region of the electromagnetic spectrum that corresponds to visible light.

ATOMIC EMISSION SPECTROSCOPY

Flame tests provide only limited information about the likely elements present in a sample. As mentioned previously, only a few elements give a coloured flame in a Bunsen burner, and the colours of some are similar. In impure samples, a faint colour may be masked by a stronger one.

An atomic emission spectrometer is an apparatus that can be used to determine the metal ions present in a sample (Figure 5.2.8). By making two changes to the flame test technique, the reliability and usefulness of flame colour identification is greatly improved.

- Using a hotter flame ensures that sufficient energy is available to excite electrons in a wider range of elements.
- Passing the light through a prism separates the different energies in the light emitted by a heated sample into a series of coloured lines called a **line spectrum** or an **emission spectrum**.

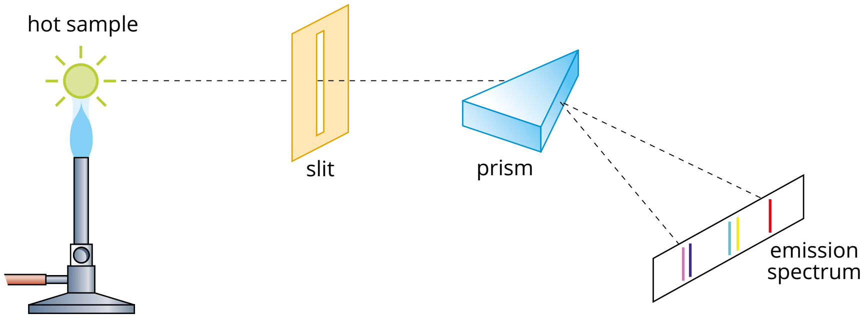


FIGURE 5.2.8 Essential components of an atomic emission spectrometer. The prism splits the emitted light into each of its component wavelengths, creating a line spectrum that is unique to the sample being tested.

When the light emitted from an excited atom is viewed through a spectrometer, it is much easier to see how the colour produced is unique. Because each element has a different number of protons in the nucleus and a unique electron configuration, the energy transitions that the electrons undergo as they move from the ground state to an excited state are also unique. Therefore, no two elements have energy levels of exactly the same energy, so a spectrum is characteristic of a particular element.

It may be used as a 'fingerprint' to identify the elements present in a substance. The emission spectra of calcium, sodium, mercury and cadmium are shown in Figure 5.2.9 as examples. You can see how the spectra consist of distinct lines corresponding to the different colours emitted by the atoms.

FIGURE 5.2.9 The emission spectrum of (a) calcium, (b) sodium, (c) mercury and (d) cadmium. The spectral lines are indicative of the different energies of light emitted in the visible region.

Formation of different spectral lines

Earlier in this module, you saw that when unstable excited electrons in higher energy levels fell back to lower energy levels they emitted energy as light. When falling back to the same shell, an electron from a higher energy level emits more energy, which corresponds to light of a lower wavelength, compared to an electron that falls from a lower energy shell. Consider the Lyman series of hydrogen, where the electrons fall back to $n = 1$ ground state. When an electron falls from $n = 3$ shell, it emits a photon with 12.09 eV (102.57 nm) compared to an electron that falls from $n = 2$ shell, which emits a photon with 10.20 eV (121.57 nm).

In the early twentieth century, a series of spectral lines was observed in the hydrogen spectrum. The lines are grouped together in a series. There are several series of spectral lines, showing specific energy level spectrums. You can see in Figure 5.2.10 in the Balmer series of hydrogen that these lines get closer together until a **continuum** occurs. This indicates that the differences between each transition becomes less and less until it forms a continuum.

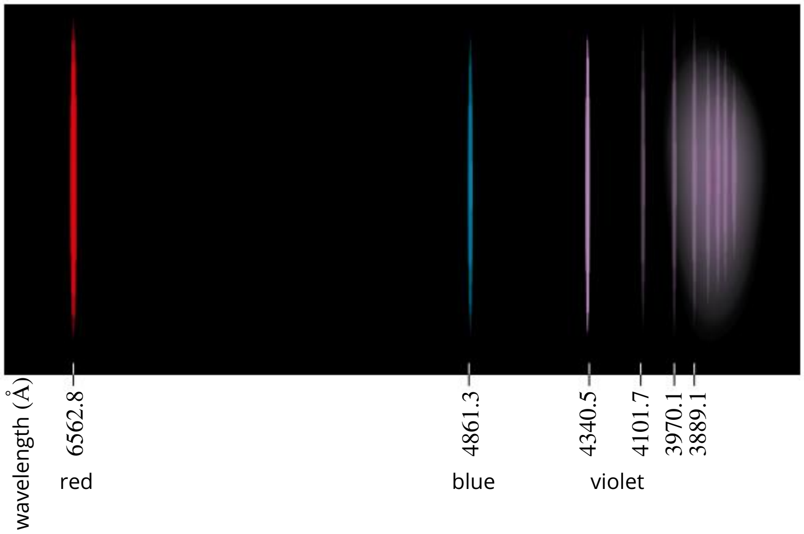


FIGURE 5.2.10 The electron transitions that create the Balmer series of spectral lines (in the visible spectrum) for the hydrogen atom

The series of lines observed corresponded to the energy levels to which the electrons returned. Electrons that fall back to the shell $n = 1$ are part of the Lyman series (ultraviolet). Emission lines produced from electrons that fall back to the $n = 2$ shell are known as the Balmer series (visible) and the Paschen series (infrared) corresponds to the spectral series in which electrons fall back to the $n = 3$ shell. So all transitions from higher energy levels to $n = 1$ shell result in lines in the ultraviolet region, all transitions to $n = 2$ shell result in visible lines and all transitions to $n = 3$ shell result in infrared lines. The formation of the Lyman, Balmer and Paschen series of lines for the hydrogen atom is shown in Figure 5.2.11. There are additional spectral lines corresponding to higher energy shells ($n = 4$ and $n = 5$) but these are outside the scope of this course.

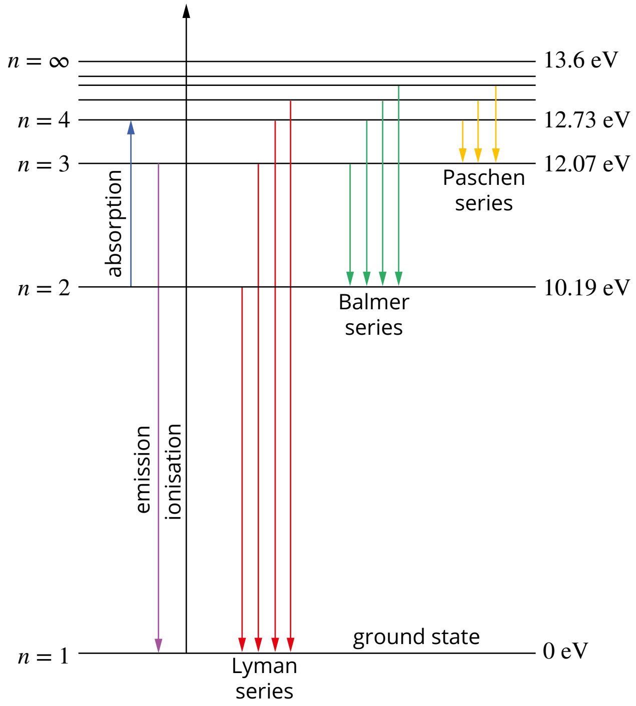


FIGURE 5.2.11 The electron transitions that create the Lyman, Balmer and Paschen series of spectral lines for the hydrogen atom

ATOMIC ABSORPTION SPECTROSCOPY

In the early 1950s, Australian scientist Alan Walsh was working on the measurement of small concentrations of metals at the CSIRO. During this time he developed the technique of **atomic absorption spectroscopy (AAS)**, which is now used widely for detecting the presence of most metals and determining their concentration.

How AAS works

Alan Walsh's breakthrough was to recognise that atoms will absorb light if the energy of the light is exactly equal to the energy required to promote an electron from its ground state energy level to a higher energy level. As every element absorbs light of different energies (and hence different wavelengths), the amount of light absorbed by a sample at a specific wavelength can be used to determine the concentration of that element.

Unlike emission spectroscopy, AAS measures the amount of light *absorbed* by the sample.

Figure 5.2.12 gives a simplified depiction of how an atomic absorption spectrometer works.

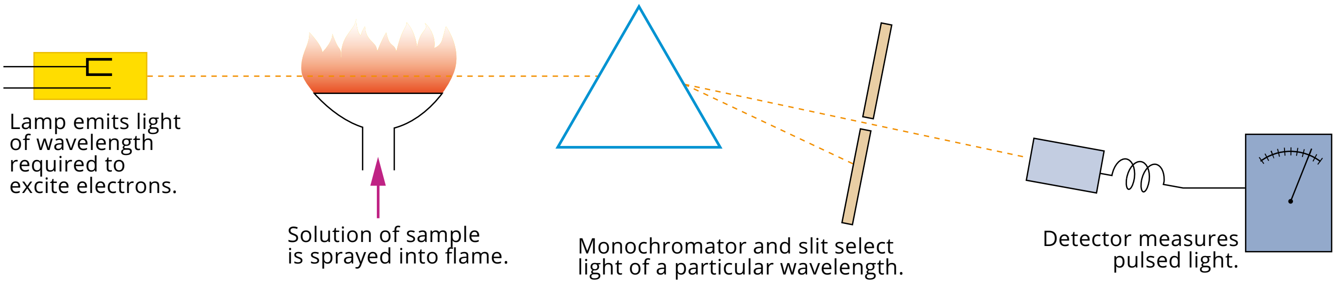


FIGURE 5.2.12 Schematic diagram of the inner workings of an atomic absorption spectrometer

The light absorbed at a specific wavelength by a sample can be determined by the following process.

1. A hollow source cathode lamp emits different wavelengths of light that are absorbed by the metal being analysed. The lamp must be made with a filament of the metal being analysed.
2. A solution of a sample is sprayed into the flame to create an atomic vapour.
3. A monochromator is used to select a wavelength of light for analysis.
4. A detector measures the amount of light that reaches it, and a computer determines the amount of light that has been absorbed by the sample.

Calibration curves

Atomic absorption spectroscopy can be used to simply detect the presence of most metals; however, it is more often used to determine the concentration of a metal in a sample. The absorbance measured for a sample can be related to the concentration of the metal being analysed by using a calibration curve.

To construct a calibration curve, you must first create a series of standard solutions of the metal ion, and then measure their absorbance by AAS. A calibration curve is then constructed by accurately plotting the concentrations of the standard solutions against the absorbance of each solution, as shown in Figure 5.2.13.

Because the amount of light that is absorbed by the sample is proportional to the amount of metal present, the relationship between concentration and absorbance is a linear one.

Once constructed, the calibration curve can be used to determine the concentration of the metal being analysed in the unknown sample. The absorbance of the unknown sample is measured and the corresponding concentration value can be read from the graph.

There are many different units for concentration. For simplicity, concentration will be measured in mg L^{-1} in the following examples, but it is important to always check the concentration unit specified on the *x*-axis of the calibration curve when answering questions involving AAS. In the example in Figure 5.2.13, the unit mg L^{-1} refers to the mass of metal (in mg) in every litre of the solution being analysed.

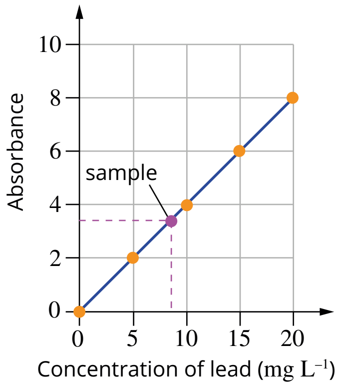


FIGURE 5.2.13 The AAS calibration curve shows the relationship between the absorbance of light and the concentration of lead in a sample.

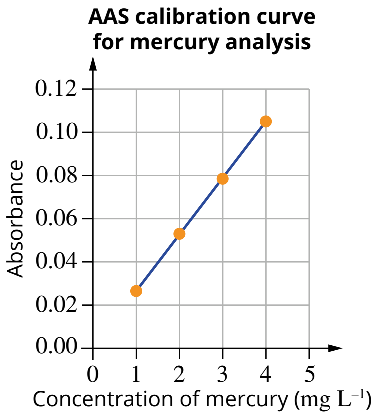
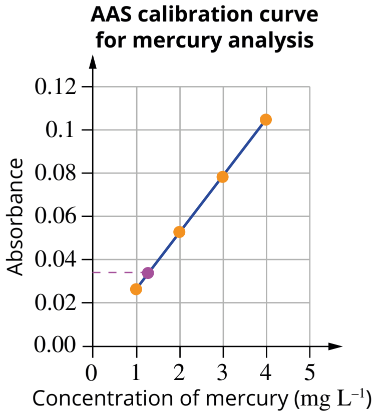
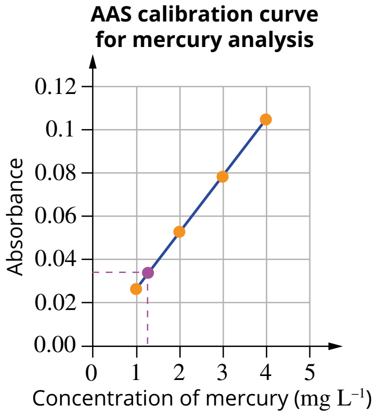
ℹ As the concentration of the metal in the sample increases, the amount of light absorbed by the sample increases.

Worked example 5.2.1

USING A CALIBRATION CURVE TO DETERMINE CONCENTRATION

Determine the concentration of mercury in a sample, given the data in the following table.

Concentration of mercury (mg L^{-1})	**Absorbance**
1.0	0.026
2.0	0.053
3.0	0.078
4.0	0.105
sample	0.036

<table>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Construct a calibration curve using concentrations of the standard solutions and their absorbance.</td><td>
</td></tr>
<tr><td>Mark where the absorption of the sample lies on the calibration curve. This can be done by finding the absorption of the sample on the y-axis and moving horizontally right until you reach the calibration curve and marking that point.</td><td>
</td></tr>
<tr><td>Determine the concentration of the sample by reading off the graph on the x-axis at the point where the sample's absorption lies on the calibration curve.</td><td>

The sample's absorption is 0.036.
The concentration of mercury for the sample can be read off the graph as 1.2 mg L^{-1}.</td></tr>
</table>

➤ Try yourself 5.2.1

USING A CALIBRATION CURVE TO DETERMINE CONCENTRATION

Determine the concentration of sodium in a sample, given the data in the following table.

Concentration of sodium (mg L^{-1})	Absorbance
20	0.041
40	0.080
60	0.121
80	0.159
sample	0.104

AAS TODAY

Approximately 60 years after Alan Walsh developed AAS, the technique is still being used by many modern-day chemists. It is a very good tool for analysing the concentration of metals in a wide variety of samples, including water, urine, blood, soil, fish and other foods. AAS is capable of detecting 68 different metallic elements, and is sensitive enough to detect metals present in concentrations of less than $1\ \text{mg L}^{-1}$.

Figure 5.2.14 shows an analytical chemist using an atomic absorption spectrometer to analyse the concentration of potassium in a sample.

FIGURE 5.2.14 An analytical chemist uses AAS to determine the concentration of metals in food samples.

ICP–AES analysis

Even with the use of very hot flames (a mixture of oxygen and ethyne gas gives the hottest flame, up to 3100°C), there are still many elements that cannot be analysed by AAS or give a poor response.

Modern instruments, such as that seen in Figure 5.2.15, can analyse most elements by using an energy source called **inductively coupled plasma (ICP)** in combination with atomic emission spectroscopy (AES). Instead of a flame to excite the atoms, the ICP generates very high temperatures of up to 10 000°C to create a **plasma**, a state of matter containing charged particles. At these temperatures, virtually all the atoms in the sample are excited and are able to emit electromagnetic radiation as an emission spectrum as they return to their ground state.

There are several advantages of using ICP–AES over AAS.

- It can be used to identify most elements.
- It is suitable for almost all concentrations.
- It can rapidly identify many elements present in a sample at the same time, as the emission spectrum is resolved by comparison with a computerised database of spectral lines. Analysis of 70 elements together takes just two minutes, whereas AAS can determine only one element at a time.

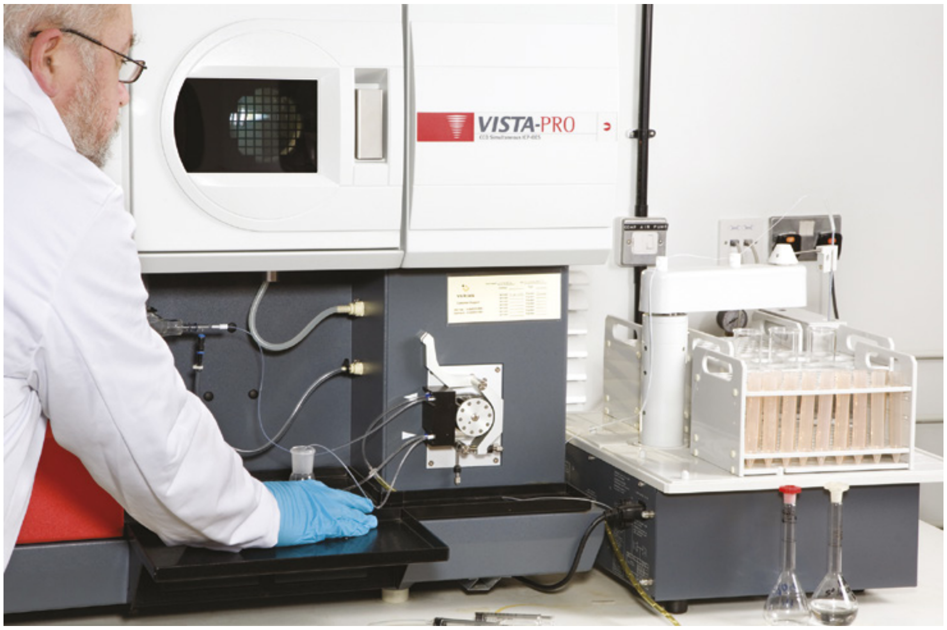


FIGURE 5.2.15 An ICP atomic emission spectrometer

5.2 Review

SUMMARY

- An atom in which the electrons are in their lowest possible energy state is said to be in its ground state.
- When an element is heated, electrons may absorb energy and move into higher energy levels (shells). The atom is described as being excited.
- When the electrons move back to lower energy levels, they emit electromagnetic radiation, such as light, with specific energies; each element emits light with a unique set of energies called an emission spectrum.
- The shell to which the electron returns determines the spectral series associated with the light emission. For the hydrogen atom, if $n = 1$ the emission is part of the Lyman series (ultraviolet), $n = 2$ Balmer series (visible), $n = 3$ Paschen series (infrared).
- Flame tests are performed by inserting a sample in a non-luminous Bunsen burner flame.
- Flame tests can be used to detect the presence of a small number of metal elements.
- The accuracy of flame tests is hampered by our ability to detect small differences in colours.
- Emission spectra are unique for each element and represent each of the energy transitions within an element as a series of different-coloured lines.
- The technique of atomic absorption spectroscopy (AAS) is based on the ability of electrons to absorb energy as they move between energy levels.
- AAS can be used to accurately determine the concentration of most metals in samples of water and other substances.
- The AAS technique involves the construction of a calibration curve to relate the concentration of the metal to the absorbance measured.

KEY QUESTIONS

Retrieval

1 State the wavelengths of light in the visible region of the electromagnetic spectrum.

2 Select words from the following list to complete the sentences below. Not all of the words provided are required.

protons, higher, transition, electrons, lower, let out, emit, excited

When a sample containing copper is heated in the flame of a Bunsen burner, the flame turns a green colour. This is because the ________ in the copper atoms absorb energy and move to ________ energy levels and then ________ light that corresponds to a green colour as they return to ________ energy levels.

3 Name the first three spectral series for the hydrogen atom and the region in the emission spectrum where the light is observed.

Comprehension

4 Explain what an electromagnetic spectrum is and how the electromagnetic waves are classified.

Analysis

5 A student observed the following results from his flame test experiment, shown below. Using the characteristic flame colours shown in Table 5.2.1 on page 117, analyse these results, identify any anomalies and suggest a reason for the differences.

Metal	Flame colour
sodium	yellow
strontium	yellow
copper	green
lithium	crimson

6 Figure 5.2.3 on page 116 shows a flame test being performed.

- **a** Identify what colour the flame would be if copper were present in the sample.
- **b** Deduce why it is necessary to heat the wire strongly for several minutes before performing the test.
- **c** Explain why copper wire would be unsuitable for use in flame tests of different metals.
- **d** Explain why flame tests are not used for qualitative analysis by modern chemists.

5.2 Review *continued*

7 Compare and contrast the Lyman and Balmer series of spectral lines for the hydrogen atom based on your understanding of atomic structure.

8 Determine whether a photon of light is emitted or absorbed when an electron goes from the level $n = 5$ to $n = 2$.

9 Evaluate why atomic emission spectroscopy is regarded as a superior method of analysis to flame tests.

10 **a** Plot a calibration curve using the absorbance readings of the standard solutions containing potassium ions given below.

Concentration of potassium (mg L^{-1})	Absorbance
0.0	0.01
2.0	0.08
4.0	0.15
6.0	0.21
8.0	0.27

b Use the calibration curve you created in part **a** to determine the concentration of potassium, in mg L^{-1}, in a sample solution that gives an absorbance of 0.17.

11 Iron is essential to our health. To determine the iron content in a Milo milk drink, a 5.0 mL sample was diluted to 50.0 mL. The absorption of the diluted solution and of several standard solutions was measured using AAS. The results are shown in the following table.

Solution concentration (ppm)	Absorbance
0.00	0.010
1.00	0.080
2.00	0.150
3.00	0.220
4.00	0.290
sample	0.190

a Plot a graph of absorption against concentration of iron.

b Calculate the concentration of iron, in ppm, in the diluted Milo.

c Calculate the concentration of iron, in ppm, in the undiluted Milo.

d Determine the mass of iron you would consume by drinking a 250 mL glass of Milo.

e The recommended daily allowance (RDA) of iron for people over the age of 11 years is 18 mg. Calculate the percentage of your RDA that a 250 mL glass of Milo provides.

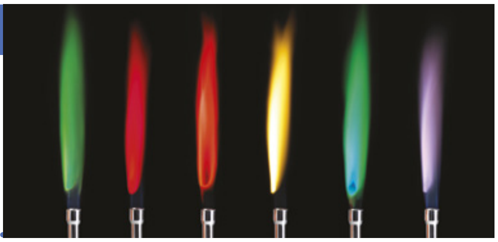
f The 0.00 ppm standard, which contained no added iron, gave a small absorption reading. Suggest an explanation for this.

g Suggest why the sample of Milo was diluted in order to measure its absorption.

5.3 How analytical chemistry has expanded our understanding of the universe

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain how atomic absorption spectroscopy has contributed to our understanding of the birth of our universe
- understand how mass spectroscopy and radioactive dating can be used to determine the age of objects on Earth.



For centuries, scientists have wondered about the chemical composition and beginning of the universe. Analytical techniques have contributed to our understanding of the processes in the very early universe and age of objects in the world around us. In this module, you will be introduced to how some analytical techniques have helped scientists determine the processes occurring in the very early universe.

HOW DO WE KNOW WHAT STARS ARE MADE OF

The light produced by stars can be viewed through a spectroscope to produce a emission spectrum just like that seen for the light from a flame test. The **Hubble Space Telescope**, which is responsible for many of the well-known images of our universe, is equipped with a spectrograph. This allows astronomers to analyse the spectra of stars and determine what they are made of.

The brightest star in our night sky is Sirius, a binary white star in the constellation Canis Major, approximately 6.8 light-years away from Earth. Figure 5.3.1 shows Sirius as seen in the southern sky and its corresponding spectrum. The light emitted from stars passes through the atmosphere. Each element within the star's atmosphere creates an absorption spectrum resulting in a series of dark lines. The dark lines correspond to the light absorbed by the element to promote an electron to a higher energy shell.

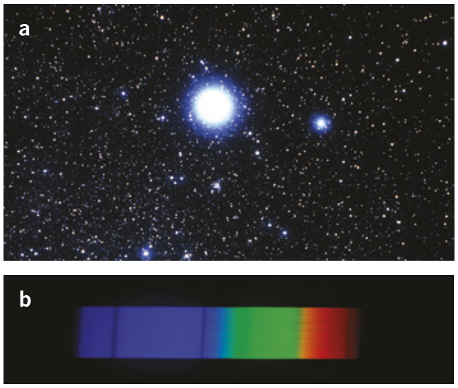


FIGURE 5.3.1 (a) The star Sirius, in the constellation Canis Major, as seen in the southern sky. (b) The spectrum obtained from the light emitted by Sirius can be used to work out the elements the star is made up of.

In Module 5.2, you saw that atoms had particular electron shells and elements absorb light of certain frequencies associated with the energy gaps between these shells. The position of the absorption lines in the spectrum matches the position of lines observed for the colours emitted by different elements. Comparison of the spectrum obtained for a star with the spectra of known elements allows astronomers to identify the elements in the star.

Figure 5.3.2 shows the emission spectrum of 13 different stars. Each of these contains the characteristic lines representative of the absorption spectrum of hydrogen (the major component of most stars). The remaining dark lines correlate to other elements that are present in the outer layers of the star; these include calcium, sodium and iron, to name a few. Astronomers have created a classification system for the stars in our universe based on their spectra. Spectral information is one of the most powerful tools we have to investigate the stars that exist light-years away from Earth.

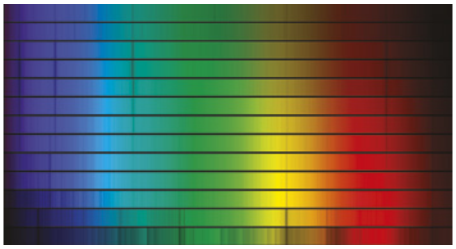


FIGURE 5.3.2 The spectra of 13 different stars. Astronomers collated the spectral data from the observable stars in the universe and classified them according to 13 types. Each of these spectra shows distinct absorption lines corresponding to the absorption pattern for hydrogen.

Our expanding universe

Spectroscopy can also determine how fast stars are moving towards or away from you by the change in wavelength of the light emitted. When an object is moving away, the wavelength of the light emitted appears to increase and is therefore said to be **redshifted**. In contrast, wavelength of light emitted by an object travelling towards the source gets smaller and the light is said to be **blueshifted**. This phenomenon is known as the **Doppler effect**.

The Doppler effect is the apparent change in frequency of electromagnetic waves caused by the relative motions of the source and the observer.

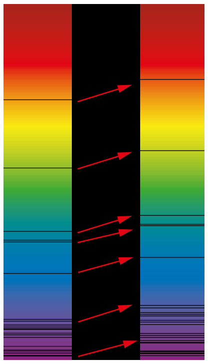


FIGURE 5.3.3 The dark absorption lines corresponding to the Balmer series of a star at rest (left) get shifted towards red if the star is moving away from Earth (right).

The atmosphere around all stars includes hydrogen gas and so when the light emitted from these stars passes through their atmosphere the hydrogen gas absorbs certain frequencies from the emitted radiation. This means that the spectral lines corresponding to the **Lyman**, **Balmer** and **Paschen series** are missing from the continuous spectrum, shown in Figure 5.3.3. The size of the redshift allows scientists to calculate how fast the star is moving away from Earth and based on the speeds of different stars, they have been able to calculate that the universe is approximately 14.3 billion years old.

Spectroscopy also allows us to determine the age of a star. When stars were formed after the Big Bang, they were composed of hydrogen and helium; however, as the star ran out of hydrogen, the helium began to fuse into heavier and heavier elements producing elements up to iron. Therefore, analysis of the spectral lines from stars allows scientists to identify the elements present and percentage composition of the star, enabling scientists to predict the age of the star.

However, stars cannot fuse iron into anything heavier and eventually gravity takes over and the star collapses, resulting in a supernova. The supernova creates such high temperatures that allow iron to fuse into heavier elements and the resulting explosion distributes these elements across the universe from which the next generation of stars are born.

CARBON DATING

Scientists use a number of techniques to determine the age of objects on Earth such as fossils. The techniques include **radioactive dating** and **accelerator mass spectrometry**.

Radioactive dating can be used to identify the age of material that was once living. Radioactive dating relies on organisms being able to uptake carbon while they are alive. Since the **isotopic abundance** of carbon-14 is fairly constant in our atmosphere the percentage of carbon-14 contained by all living organisms is also constant. However, when an animal dies it stops taking up carbon-14.

At this point, the carbon-14 atoms decay by emitting a beta particle with a **half-life** of about 5600 years, meaning that the ratio of ^{14}C to ^{12}C decreases. By counting the radioactive emissions of ^{14}C from the sample the scientists are able to estimate how long ago the animal died. This technique has been used to determine the age of artefacts as old as 50 000 years.

Recently, a technique has been developed using a more advanced type of mass spectrometer called an accelerator mass spectrometer (AMS). The AMS has a much greater sensitivity and can detect differences in the ratio of ^{14}C to ^{12}C of 10^{-15}, which has allowed samples as small as 1 mg to be analysed and has extended the dating range back to about 100 000 years. In contrast to radioactive dating, which works by measuring the radioactive decay of ^{14}C, AMS simply determines the isotopic abundance of the ^{14}C and ^{12}C **isotopes** in the sample.

The major advantages of AMS over radioactive dating result from the opportunities to date materials that were not alive and by reducing the amount of material that is required for analysis. This means that it is possible to date an object without significantly damaging it.

SCIENCE AS A HUMAN ENDEAVOUR

Distribution of elements in the universe

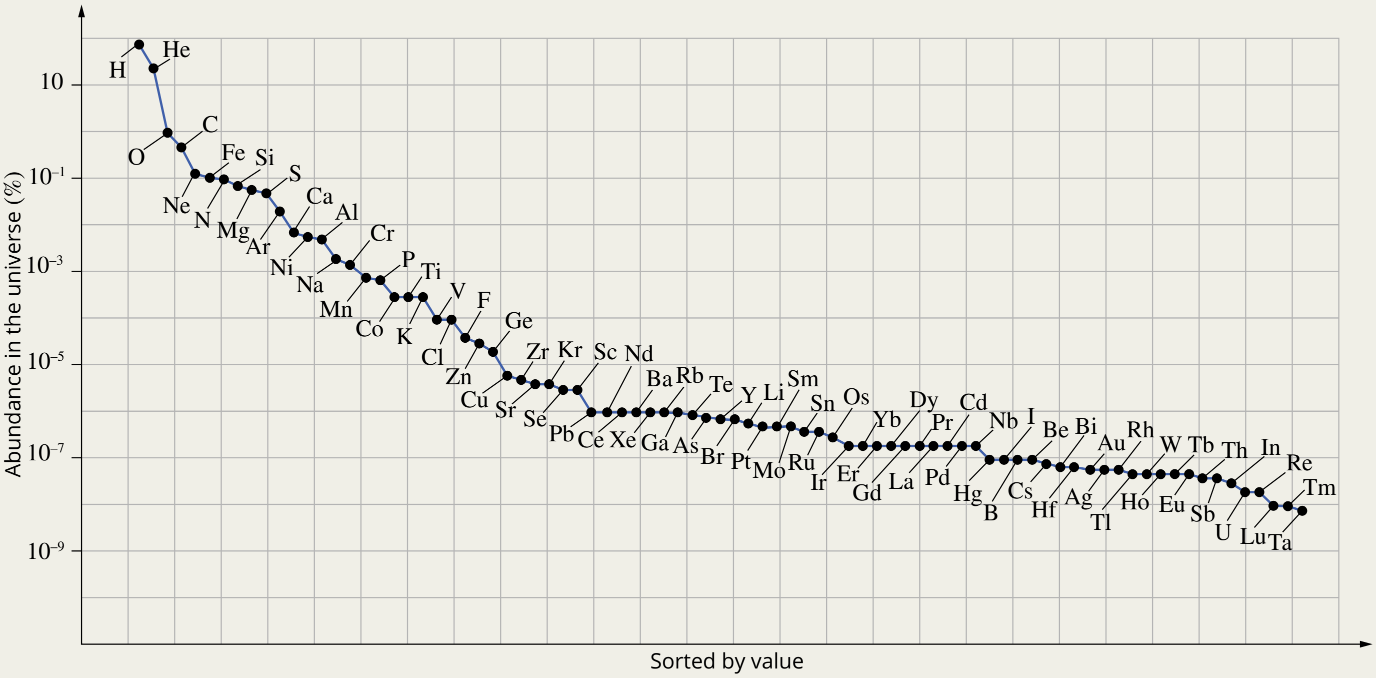


FIGURE 5.3.4 Abundance of the elements in the universe

Scientists have discovered that matter in the universe is composed of about 100 elements. They propose that these elements have been formed by different nuclear reactions and that the distribution of elements within the universe reflects these processes and the relative nuclear stability of different nuclides.

Furthermore, cosmochemists have discovered that the distribution of elements within the universe (Figure 5.3.4) is very different from the composition of the Earth. They conclude that this reflects the way the Earth was formed and then how organisms originated and evolved to what we have today.

There are several theories that have been proposed based on the analysis of the distribution of elements within living things, on Earth and in the universe.

Big Bang Nucleosynthesis (BBN)

While there is still much work to be done, analysis of the abundance of the elements indicates that the elements were formed following the Big Bang. The story of the BBN began in 1940 when Ralph Alpher performed theoretical calculations, which showed that the early universe contained neutrons. Some of these neutrons decayed into protons and electrons almost immediately as the universe expanded.

Alpher, Gamow and Bethe proposed that the nuclei produce heavier elements through the combination of additional nucleons. Their theory failed to predict the production of elements with more than five nucleons, since it relied on a stepping stone approach to the formation of elements.

It was 380 000 years after the Big Bang that the universe was sufficiently cool for the nuclei to capture electrons to form neutral atoms. Alpher and Robert Herman (Figure 5.3.5) predicted that this would emit cosmic microwave background (CMB) radiation, which was measured by Penzias and Wilson in 1964, confirming Alpher and Herman's theory.

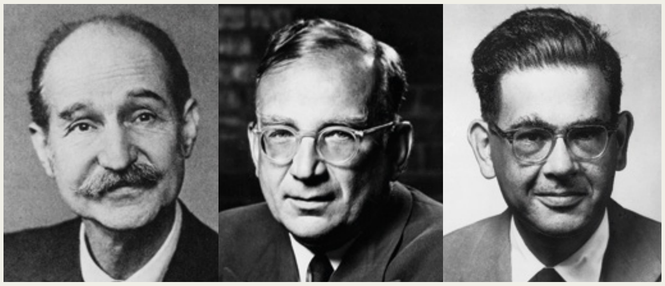


FIGURE 5.3.5 From left to right, Robert Herman, George Gamow and Ralph Alpher

As the universe continued to expand and cool, the protons and neutrons captured the electrons to form the light elements, hydrogen (H) and helium (He), which make up to 99% of all elements in the universe. Fred Hoyle proposed that the fusion within stars is capable of forming the heavier elements between carbon and iron.

Comparison of animal composition

Elemental analysis shows that humans and bacteria are very similar; however, the one element that is very different is calcium (Ca). Humans contain approximately six times more calcium than bacteria; bacteria actually excrete calcium. Calcium makes up bones, allowing greater mobility and eventually leading to the ability of animals to live on land.

Stellar evolution

Our universe is continuing to grow and expand. As you have learnt in this module, the spectral analysis of a star's light reveals its composition, temperature and motion. This analysis has revealed that stars are composed of elements up to iron (Fe) and that the universe is expanding.

Despite its abundance in the universe, hydrogen makes up only 9% of our body. Ninety per cent of the elemental composition of our body comes from carbon (C), oxygen (O), nitrogen (N), sulfur (S) and phosphorus (P) with the remaining 1% from various trace elements.

While the elements up to Fe can be made in nuclear fusion reactions within stars, meaning that we are all small pieces of stars, elements heavier than iron can only be formed by a supernova, which is a massive explosion of a star (Figure 5.3.6).

The fusion reactions that produce elements heavier than iron are endothermic and cause the star to lose energy. During the collapse of a star, the outer layers collide with the incompressible stellar core to produce a shockwave that expands outward through the unfused material of the outer shell. This induces fusion in that material, and the energy released leads to the star's explosion, dispersing material from the star into interstellar space.

Review

1 Compare the methods by which scientists analyse the elemental composition of stars and our body.

2 Aditya Chopra and Charles H. Lineweaver, in their paper entitled 'The major elemental abundance differences between life, the oceans and the Sun', investigated the elemental composition of life, the Sun and sea. They discuss how their findings could further our understanding of our origin and the evolution of life on Earth. Analyse their findings and evaluate whether their claims are accurate.

FIGURE 5.3.6 The supernova remnant W49B. By tracing the distribution and amounts of different elements in the stellar debris field, researchers were able to compare the data gathered by the *Chandra* spacecraft to theoretical models of how a star explodes.

5.3 Review

SUMMARY

- Analysis of spectral lines from stars allows scientists to identify the elements present and percentage composition of the star, the age of the star and how fast and in which direction the star is moving.
- Stars produce light in a similar way to that from a flame test.
- Elements absorb light of certain frequencies associated with the energy gaps between their electron shells.
- The position of the absorption lines in the emission spectrum of stars matches the position of lines observed for the colours emitted by different elements.
- Hydrogen is the major component of most stars and so characteristic lines representative of the absorption spectrum of hydrogen are visible in the emission spectrum of all stars.
- Other elements that are present in the outer layers of the star include helium, calcium, sodium and iron.
- Spectroscopy can determine how fast stars are moving towards or away from you by the change in wavelength of the light emitted.
- Redshifted light means that the wavelength of the light emitted has increased compared to that of a stationary star.
- Blueshifted light means that the wavelength of the light emitted has decreased compared to that of a stationary star.
- The size of the shift is related to the speeds of the star and Earth and is known as the Doppler effect.
- Radioactive dating can be used to identify the age of material that was once living by measuring the ratio of carbon-14 to carbon-12.
- Radioactive dating relies on organisms being able to take up carbon while they are alive and stop their uptake of ^{14}C once they die.
- The carbon-14 atoms decay by emitting a beta particle with a half-life of about 5600 years.
- By counting the radioactive emissions from ^{14}C from the sample, the scientists are able to estimate how long ago the animal died.
- Recent development of AMS has improved the sensitivity and range of carbon dating.

KEY QUESTIONS

Retrieval

1 Define spectroscopy.

2 State the principles behind carbon dating of a tree.

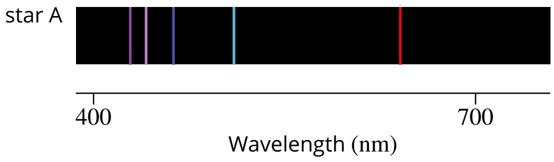
Comprehension

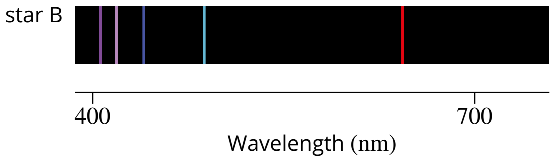
3 Explain how astronomers might use spectroscopy to determine the composition of a star.

4 Explain how astronomers use the Doppler effect to determine the velocities of astronomical objects.

Analysis

5 Compare the absorption spectra of star A and star B, shown below. Explain what information can be deduced about their motion from the spectra.





5.3 Review *continued*

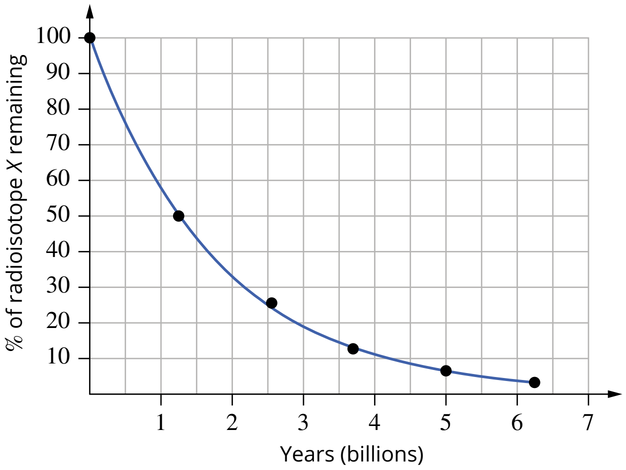
6 Examine the radioactive decay curve of element *X*, shown in the following graph.

a Determine the half-life of element *X*.

b Compare the time taken for the isotope to change from 100% to 50% with the time taken for the isotope to decay from 50% to 25%.

c Determine the age of a rock containing only 40% of its original amount of element *X*.

d Based on the information in the table below, determine the identity of element *X*.



Radioisotope	**Half-life (billion years)**
samarium-146	0.104
uranium-235	0.704
potassium-40	1.28
uranium-238	4.47
thorium-232	14.06

7 If the half-life of carbon-14 is 5700 years, apply this information to determine the ages of these artefacts.

a a skull with 50% of its original carbon-14 remaining

b a fossilised shark's tooth with about 12% of the original carbon-14 remaining

c a small sample of paper taken from the Dead Sea Scrolls that contains 75% of the original carbon

Chapter review

05

KEY TERMS

accelerator mass spectrometry
atomic absorption spectroscopy (AAS)
Balmer series
blueshifted
continuum
Doppler effect
electromagnetic radiation
electromagnetic spectrum
emission spectrum
excited electron
ground state
half-life
Hubble Space Telescope
inductively coupled plasma (ICP)
isotope
isotopic abundance
line spectrum
Lyman series
mass spectrometer
mass spectrum
Paschen series
plasma
radioactive dating
redshifted
relative atomic mass
relative isotopic abundance
relative isotopic mass
spectroscopic technique
spectroscopy

KEY QUESTIONS

Retrieval

1 An element has the following three isotopes: $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$.

a Identify the element.

b State what the numbers 12, 13 and 14 refer to.

c State what the number 6 in each isotope means.

d Identify the number of protons, neutrons and electrons in each isotope.

2 Define the term ‘relative isotopic abundance’ of isotopes.

3 For each component of an AAS, identify the corresponding description.

Component	Description
flame	selects a specific wavelength of light
hollow source cathode lamp	measures the amount of light
computer	produces light with wavelengths that are absorbed by the metal being analysed
monochromator	where the sample is sprayed and light is absorbed
detector	converts the amount of light detected into the amount of light absorbed by the sample

4 Select the option that best identifies the reason for the distinctive yellow colour in a flame test of sodium nitrate.

A The colour is due to sodium atoms losing an electron to form sodium ions.

B The colour is due to electrons near the nucleus of the sodium atom absorbing energy.

C The colour is due to sodium ions being converted to sodium atoms in the reducing flame.

D The colour is due to excited electrons returning to lower energy levels and releasing energy.

5 An electron in a lithium atom gains enough energy to move from the *n* = 1 to the *n* = 2 shell. Select which option best describes the end state of this lithium atom.

A The lithium atom becomes a positively charged ion.

B The lithium atom becomes a negatively charged ion.

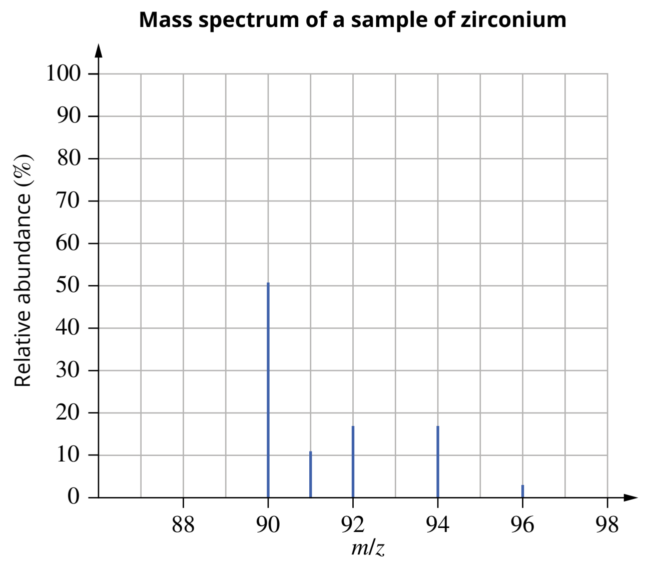
C The lithium atom becomes an atom in an excited state.

D The lithium atom becomes an atom in the ground state.

Comprehension

6 Determine the atomic number, mass number and symbol for the element that has 45 protons and 58 neutrons.

7 Explain how many isotopes are present in the mass spectrum of a sample of zirconium shown below.



Mass spectrum of a sample of zirconium

8 Potassium chloride can be used as a replacement for table salt by people suffering from high blood pressure. While cooking, some potassium chloride was spilt in the flame of a gas stove.

a Identify the colour of the flame that was observed.

b Describe what would be observed if the potassium chloride contained some sodium chloride.

c Explain why the chloride ions have no effect on the flame colour.

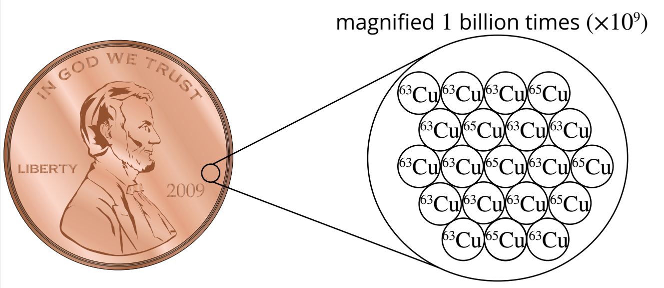
9 Explain why an emission spectrum contains a number of lines of different colours.

10 Explain how the absorption spectrum from a star is similar to a person's fingerprint.

Analysis

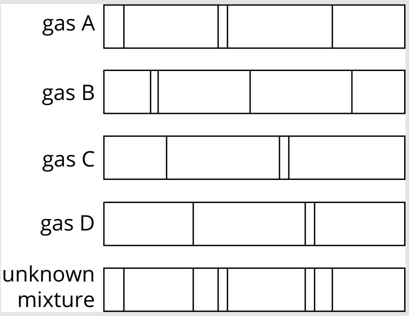
11 Determine whether all atoms of an element are identical. Explain how you would determine whether this is the case.

12 The elemental composition of a copper coin is shown in the diagram below. Calculate the relative atomic mass of copper using this information.

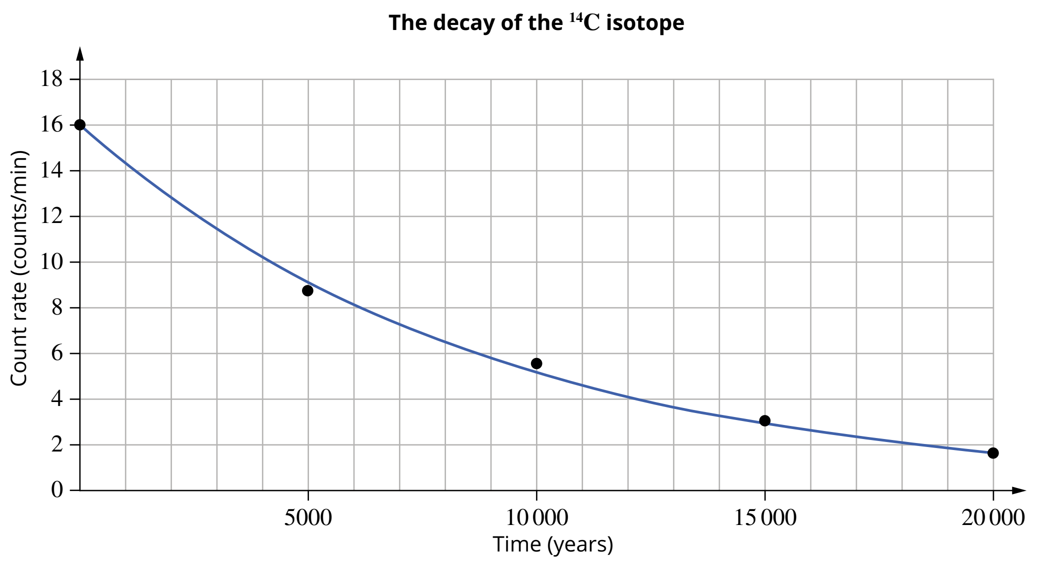


13 Barium and calcium are both group 2 elements. Deduce why barium and calcium compounds produce different colours when they are held in a flame.

14 Identify the two gases present in the mixture using the emission spectra shown in the diagram below.



15 The graph below shows the decay of the ^{14}C isotope. Trees absorb carbon from the atmosphere. After the tree dies, the uptake of carbon dioxide ceases and the proportion of carbon-14 decreases as it decays.



Archaeologists exploring a town in South America that was destroyed by a giant volcanic eruption uncovered a tree killed by the blast. The radioactive analysis conducted found the count rate was 10.5 counts per minute. Determine the year of the eruption using the data from the graph.

16 A highly controversial scientific debate in recent years was the authenticity of the Shroud of Turin. In 1988, samples from the shroud were taken for carbon dating. These were analysed at three highly prestigious laboratories in the USA, England and Switzerland. The results showed that the shroud had lost about 8% of its carbon-14 atoms. Deduce the approximate date the Shroud of Turin was made using the graph in Question **15**.

17 The concentration of copper(II) ions in industrial wastewater was analysed by atomic absorption spectroscopy. The absorbance values from a series of standards and the wastewater are provided below.

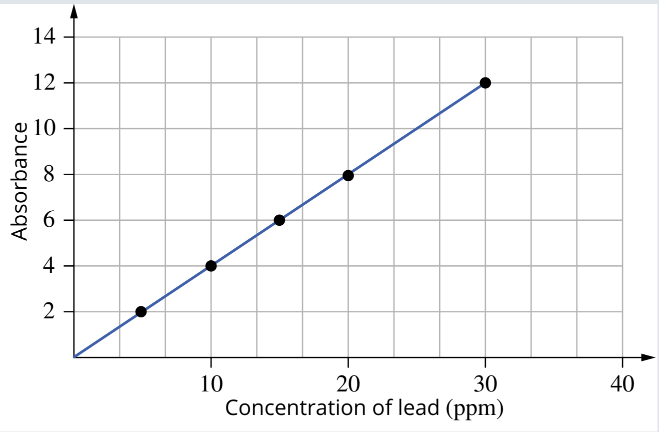
Concentration (mg L^{-1})	**Absorbance**
50	0.12
100	0.23
150	0.36
200	0.48
250	0.58
wastewater sample	0.42

Determine the concentration of copper(II) ions in the wastewater.

18 A batch of oysters was suspected of containing dangerous levels of lead. The following is an account of their analysis.

One of the oysters was removed from its shell and chopped finely in a food processor. A 1.50 g portion was heated on a hotplate with 10 mL of nitric acid. This mixture was filtered and then sprayed into the flame of an atomic absorption spectrometer. A hollow cathode lamp, which emitted light with a wavelength that is absorbed by lead atoms, was in place. The spectrometer measured an absorbance of 3.6 indicating that there was lead in the oyster.

To find the concentration of lead ions in the sample, the absorbance of different solutions containing known concentrations of lead ions was also measured. The results are graphed below.



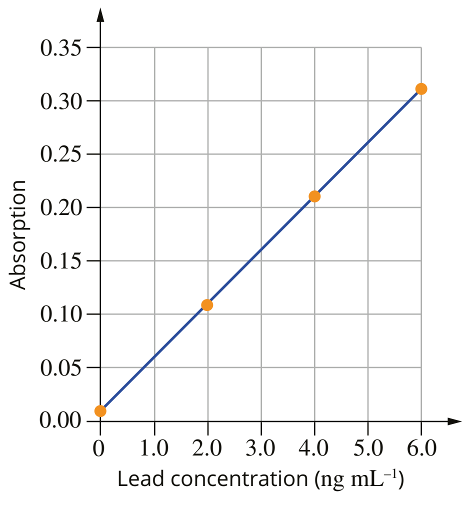
The recommended limit of lead in oysters for human consumption is 2.5 ppm. Determine whether these oysters are fit to eat.

19 Element A occurs naturally in the environment. Scientists have found 322 atoms of A-17, 569 atoms of A-18 and 2780 atoms of A-19. Calculate the relative atomic mass of the element A to 2 decimal places.

20 Use the data in Table 5.1.2 on page 108 to calculate the relative atomic mass of these elements.

- **a** oxygen
- **b** silver
- **c** hydrogen

21 Lead exposure can cause permanent brain damage in infants and young children, even at very low levels. The amount of lead in infant milk formula can be measured by atomic absorption spectroscopy. A 2.5 g sample of milk powder was dissolved in 50 mL of distilled water. A very small volume of this solution was analysed and gave an absorbance of 0.130. Four standard solutions were analysed in the same way and the following calibration graph was obtained.



- **a** Determine the concentration of lead (ng mL^{-1}) in the milk powder solution.
- **b** Calculate the concentration of lead (ng g^{-1}) in the dried milk powder.
- **c** Convert your answer to part **b** to ppm.
- **d** Calculate the percentage by mass of lead in the milk powder.
- **e** The zero standard, distilled water, gave an absorbance reading that was not zero. Suggest a reason for this.

22 The element lithium has two isotopes:

^{6}Li with a relative isotopic mass of 6.02

^{7}Li with a relative isotopic mass of 7.02.

The relative atomic mass of lithium is 6.94.

Calculate the percentage abundance of the lighter isotope.

23 Deduce how mass spectra can be used to identify the nature of the element and the masses of individual atoms of a specific element.

24 Determine which of the following properties in isotopes ^{12}C and ^{13}C is different.

A atomic number

B number of protons

C number of neutrons

D number of electrons

Knowledge utilisation

25 While sodium is an essential requirement in our diets, the amount of sodium consumed is often much higher than levels recommended by doctors. The sodium content of a sauce was determined using atomic absorption spectroscopy. A 25.00 mL sample of the sauce was diluted to 1.00 L with deionised water. Four standard solutions of NaCl were prepared and the absorbances of the four standard solutions and the diluted sauce solution were measured. The results are shown in the following table.

Concentration of Na^+ (aq) in mg L^{-1}	**Absorbance**
100	0.052
200	0.101
300	0.149
400	0.199
500	0.250
diluted sauce	0.184

a Create a calibration curve for Na^+.

b Use the calibration curve to determine the sodium ion concentration in the sauce.

c Determine the concentration of NaCl in the original (undiluted) sauce in g L^{-1}.

d Identify what important assumption you must make in order to calculate the NaCl content of the sauce from the Na^+ concentration.

e If the daily recommended NaCl intake for a healthy adult is 2.5 g, calculate the amount of sauce a person would have to consume to exceed the recommended intake.

f Explain why the AAS will only measure the sodium ion concentration in your sample and not the concentration of other ions as well.

CHAPTER 06 Metals

This chapter describes the properties and uses of metals and explains how the properties of metallic elements differ from those of non-metals.

You will learn how chemists have been able to relate these properties to the structure of metals, and to explain their structure in terms of a metallic bonding model and the utilisation of these properties in a range of everyday applications.

You will also see how the structures of metals can be modified to make them more useful, as well as how emerging nanotechnologies are finding new uses for very small metallic materials.

This chapter also looks at how the reactivities of metals determine the way they exist in the Earth's crust and how one particular metal, iron, is extracted from its ore.

Syllabus subject matter

Topic 1 • Properties and structure of atoms

INTRODUCTION TO BONDING

- recognise that the properties of atoms, including their ability to form chemical bonds, are explained by the arrangement of electrons in the atom and by the stability of the valence electron shell
- understand that the number of electrons lost, gained or shared is determined by the electron configuration of the atom and recall that transitional elements can form more than one ion*
- understand that chemical bonds are caused by electrostatic attractions that arise because of the sharing or transfer of electrons between participating atoms and the valency is a measure of the number of bonds that an atom can form*

Topic 2 • Properties and structure of materials

BONDING AND PROPERTIES

- understand that the type of bonding within ionic, metallic and covalent substances explains their physical properties, including melting and boiling point, thermal and electrical conductivity, strength and hardness*
- analyse and interpret given data to evaluate the properties, structure and bonding of ionic, covalent and metallic compounds.*

SCIENCE AS A HUMAN ENDEAVOUR

- **Nanomaterials:** Development of organic and inorganic nanomaterials is important to meet a range of contemporary needs, including consumer products, health care, transportation, energy and agriculture.

*The greyed-out section of this dot point is addressed explicitly in another chapter

6.1 Properties of metals

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the physical properties of metals
- recall the difference in properties between the main group of metals and the transition metals.

Metals constitute more than 80% of the elements in the periodic table. Metals have been important to human beings since early times. The development of civilisation can be measured by the way humans have used metals. The Copper Age (5000–3000 BCE) was followed by the Bronze Age (3000–1000 BCE) and the Iron Age (from 1000 BCE).

Gold, silver and copper can be found on Earth in an almost pure form. These metals were employed by prehistoric humans to make ornaments, tools and weapons. As humans' knowledge of metallurgy (the science of modifying metals) developed, metals have played a central role in fields as diverse as construction, agriculture, art, medicine and transport.

In this module, you will examine the properties of metals. Then you will learn about the bonding model that chemists have developed to explain these properties. This model has helped chemists and materials engineers to understand why metals behave the way they do and how metals can be modified to create useful new materials.

PROPERTIES AND USES OF METALS

The diverse properties of different metals make them suitable for many purposes. For example, titanium (Figure 6.1.1) is a very strong, relatively unreactive metal with a low density that is close to that of bone. Consequently, it is used in surgical implants that can last up to 20 years with little effect on the body. Titanium is also used in the aerospace industry, in art and architecture, and in sporting products such as golf clubs.

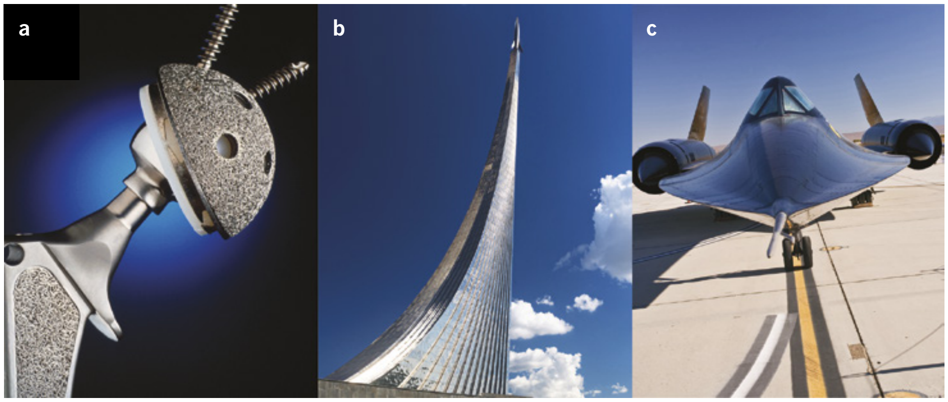


FIGURE 6.1.1 Titanium has many uses: (a) a replacement hip joint, (b) the spectacular curved space museum building in Moscow and (c) the SR-71 *Blackbird* reconnaissance aircraft. The SR-71 aircraft is the fastest manned aircraft that uses oxygen directly from the atmosphere.

Table 6.1.1 shows the uses of some metals.

TABLE 6.1.1 Properties and uses of some metals

Metal	Properties	Uses
iron	soft, malleable, magnetic, good thermal and electrical conductor, fairly reactive, readily forms alloys	can corrode and is usually converted to more stable steel, which is used in buildings and bridges, automobiles, machinery and appliances
aluminium	low density, relatively soft when pure, excellent thermal and electrical conductor, malleable and ductile, good reflector of heat and light, readily forms alloys	saucepans, frying pans, drink cans, cooking foil, food packaging, roofing, window frames, appliance trim, decorative furniture, electrical cables, aircraft and boat construction
titanium	very strong, high melting point, low density, low reactivity, readily forms alloys	medical devices within the body, wheelchairs, computer cases; lightweight alloys are used in high-temperature environments such as spacecraft and aircraft
gold	shiny gold appearance, excellent thermal and electrical conductor, unreactive, readily forms alloys	electrical connections, jewellery, monetary standard, dentistry

Table 6.1.2 gives the properties of some metals and non-metals. Despite the different properties of metals, most metals:

- exhibit a range of melting points and relatively high boiling points
- are good **conductors** of electricity
- are good conductors of heat
- generally have high **densities**.

TABLE 6.1.2 Properties of some metallic and non-metallic elements

Element	Melting point (°C)	Boiling point (°C)	Electrical conductivity ($MS\,m^{-1}$)*	Thermal conductivity ($J\,s^{-1}\,m^{-1}\,K^{-1}$)†	Density ($g\,mL^{-1}$)
Metals					
gold	1063	2970	45	310	19.3
iron	1540	3000	9.6	78	7.86
copper	1083	2567	6	385	8.96
mercury	–39	357	1	8.4	13.5
potassium	64	760	14	100	0.86
silver	961	2210	60	418	10.5
sodium	98	892	21	135	0.97
Non-metals					
carbon (diamond)	3550	‡	10^{-17}	–	3.51
oxygen	–219	183	–	0.026	1.15 (liquid)

*$MS\,m^{-1}$ = megasiemens per metre.
†Thermal conductivity measures the conductance of heat.
‡Diamond sublimes (changes straight from a solid to a gas) when heated.

FIGURE 6.1.2 This power transmission tower relies on the strength of iron in steel for its structural integrity. The electricity cables are made from aluminium, utilising its ductility and electrical conductivity.

Not all metals have all of these properties. Mercury is a liquid at room temperature. It has an unusually low melting point. The group 1 elements (the **alkali metals**) have some properties that make them different from most other metals. They are all soft enough to be cut with a knife and they react vigorously with water to produce hydrogen gas. Both mercury and the group 1 elements exhibit most of the other properties listed on page 139 and are classified as metals.

Metals also generally have the following characteristics in common.

- They are **malleable**—they can be shaped by beating or rolling.
- They are **ductile**—they can be drawn into a wire.
- They are **lustrous** or reflective when freshly cut or polished.
- They are often hard, with high **tensile strength**.
- They have low **ionisation energies** and electronegativities.

These properties can allow different metals to be used together in order to solve many engineering problems. The power transmission tower in Figure 6.1.2 is made of a few metals to take advantage of their different properties.

Generally, metals are shaped for use in different applications by hammering, exploiting their malleability. Some metals, such as gold, copper and aluminium, are very malleable at room temperature. Other metals, such as iron, must be heated before they can be shaped.

TRANSITION METALS

Between group 2 and group 13 in the periodic table is a block of elements known as the **transition metals**. These elements generally have unfilled *d*-subshells and are often referred to as the *d*-block. They include metals such as iron and nickel that are used to build cities, bridges, cars and railway lines, and precious metals such as silver and gold that have ornamental and economic uses. Most transition metals are silver-coloured and are similar in appearance, as can be seen in Figure 6.1.3.

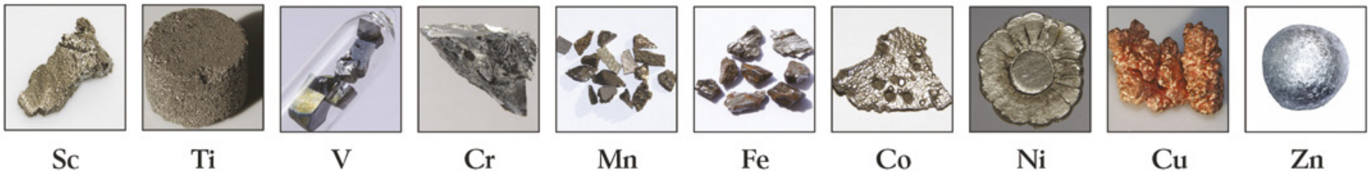


FIGURE 6.1.3 The first row of transition metals

The transition metals are very important to Australian industry. All of the metals in the first row of the *d*-block are found in Australia and many are being mined today.

Iron is by far the most important metal to us. Nearly ten times more iron is mined than all other metals combined. Iron obtained directly from a blast furnace is relatively **brittle** and corrodes easily. Carbon and other transition metals are combined with iron to produce mixtures or **alloys**, such as **steel**. Steel has more desirable characteristics than pure iron. (Alloys are covered in more detail in Module 6.4.)

Copper is one of the few transition metals that is mainly used in its pure form. Its high electrical conductivity makes it especially useful for most of the millions of kilometres of electrical wires that enable the transmission of electric energy for heating, lighting, telephones, radio and television (Figure 6.1.4).

FIGURE 6.1.4 Electrical cable used in domestic appliances. The three copper wires inside the cable are coated with plastic insulation.

Transition metals are not only important for industry, but are also important for life. All the transition metals in the first row, except scandium and titanium, are essential for animal life. Your body relies on the presence of trace elements to carry out certain biochemical reactions. For example, chromium, which you get from meat and bread, assists in the production of energy from glucose.

Properties of transition metals

Compared to the main group metals, transition metals have the following properties.

- They tend to be harder.
- They have higher densities.
- They have higher melting points.
- Some of them have strong magnetic properties.

The hardness, higher densities and higher melting points are due to the atoms of transition metals generally being a smaller size because of their greater effective nuclear charge. This allows them to pack together more tightly with stronger bonds. The decreased interatomic distance means that the electrostatic attraction between neighbouring atoms is greater and results in stronger bonding.

The high tensile strength of transition metals makes them suitable for use in the construction of buildings, cars, bridges and numerous other objects.

Transition metal compounds

Transition metal compounds display a wide range of different colours. They are extensively used as pigments in paints, and to colour glass, ceramics and enamel. In Figure 6.1.5, the colours used by the artists are caused by the different transition metals present and the colours remain vivid.

FIGURE 6.1.5 (a) *Homeward Laddie* (c. 1895) by Queensland artist Anthony Alder. (b) Indigenous Australian artwork (photograph by Dominique Landau).

Ochre is a type of hard clay that contains iron oxides and hydroxides, which can be found naturally in many colours, including red, pink, white and yellow. Ground into a powder and mixed with liquids, ochre forms a paste that has been used for millennia by Aboriginal and Torres Strait Islander peoples in Australia for body decoration, cave painting, bark painting and other artwork.

The colours of many gemstones are also due to the presence of transition metals. For example, sapphires (Figure 6.1.6) contain traces of titanium and iron in a crystal **lattice** of aluminium oxide.

The colours arise when electrons within the metal ions in the compounds absorb light of particular wavelengths and move to higher energy levels. Absorbance of light with some wavelengths and transmission of light with other wavelengths results in the compounds appearing coloured. In contrast, compounds of group 1 and group 2 metals are usually colourless.

FIGURE 6.1.6 Blue sapphires get their colour from impurities of titanium and iron.

6.1 Review

SUMMARY

- Metals have the following characteristic properties:
  - high boiling points
  - good conductors of electricity in solid and liquid states
  - malleable and ductile
  - high densities
  - good conductors of heat
  - lustrous
  - low electronegativities
  - low ionisation energies.
- The main differences between the properties of main group and transition metals are:
  - transition metals are harder
  - transition metals are more dense
  - transition metals have higher melting points
  - some transition metals have strong magnetic properties
  - transition metal compounds tend to be brightly coloured.

KEY QUESTIONS

Retrieval

1 Describe what is meant by the term 'malleable'.

2 Identify the main differences between the properties of the main group metals and the transition metals.

3 Name a metal that is a liquid at room temperature.

Comprehension

4 Sodium and iron have very different physical properties. Explain why this is so, based on where these metals are found in the periodic table.

5 Explain the meaning of the term 'ductile' when referring to metals. Identify how this property is different from malleability.

6 Explain how the metallic properties of metals and the first ionisation energy are related.

Analysis

7 Identify some physical and chemical properties *not* included in Table 6.1.2 on page 139 that you would need to consider before choosing between aluminium and iron for building a bridge.

8 Potassium is classed as a main group metal. Gold is a transition metal.

a Using the information provided in Table 6.1.2 on page 139, identify another metal that has similar physical properties to potassium.

b Using the information provided in Table 6.1.2 on page 139, identify another metal that has similar physical properties to gold.

c Identify the position of these four metals in the periodic table and give three differences between them.

9 Analyse the information in Table 6.1.2 on page 139.

a Identify the properties of copper that make it a good choice for use in electrical wiring.

b If another metal was to be chosen for electrical wiring, explain what other physical or chemical properties of the metal might influence your choice.

10 Identify why the alkali metals (group 1) are the most 'metallic' of all metals, but are not useful as structural metals.

11 Titanium is an important transition metal used in high-tech industries. Analyse the physical and chemical properties of titanium that make this metal particularly useful for the aerospace industry.

6.2 Metallic bonding

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recognise that the properties of metallic compounds can be explained by using the metallic bonding model to determine a metallic lattice structure
- understand the type of bonding within metallic substances
- analyse and interpret given data to evaluate the properties, structure and bonding of metallic compounds.

In this module, you will learn how the properties of metals can be explained in terms of the structure of the particles in metals. You will also learn about the bonding model that chemists have developed to explain these properties. The metallic bonding model has helped chemists and materials engineers understand why metals behave as they do and how metals can be modified to create useful new materials.

RELATING STRUCTURE AND PROPERTIES

The properties of metals are listed in Table 6.2.1. Each of these properties gives some information about the structure and bonding of particles in metals.

TABLE 6.2.1 Physical properties of metals and resulting conclusions about metal structure and bonding

Property	What this tells us about structure
Metals are usually hard and tend to have high boiling points.	The forces between the particles must be strong.
Metals conduct electricity in the solid state and in the molten liquid state.	Metals have charged particles that are free to move.
Metals are malleable and ductile.	The attractive forces between the particles must be stronger than the repulsive forces between the particles when the layers of particles are moved.
Metals generally have high densities.	The particles are closely packed in a metal.
Metals are good conductors of heat.	There must be a way of quickly transferring energy throughout a metal object.
Metals are lustrous or reflective.	Free electrons are present, so metals can reflect light and appear shiny.
Metals tend to react by losing electrons.	Electrons must be relatively easily removed from metal atoms.

Chemists have developed a model for the structure of metals to explain all the properties that have been mentioned so far. You can deduce from the information in Table 6.2.1 that the **metallic bonding model** must include:

- charged particles that are free to move and conduct electricity
- strong forces of attraction between atoms throughout the metal structure
- some electrons that are relatively easily removed.

METALLIC BONDING MODEL

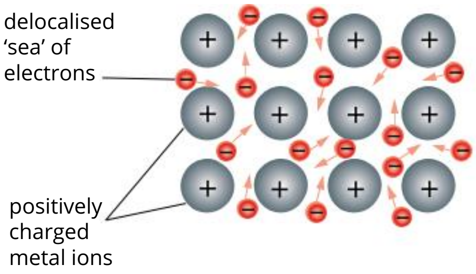


FIGURE 6.2.1 The metallic bonding model. Positive metal cations are surrounded by a mobile 'sea' of delocalised electrons. This diagram shows just one layer of metal ions.

Electrons are the particles that enable metals to conduct electricity. They are able to move between the lattice (tightly packed arrangement) of metal atoms. Negatively charged electrons can be lost from the outer shell of metal atoms, forming positive ions (cations). As shown in Figure 6.2.1, the freed electrons **delocalise** (spread through a large area) to form a 'sea' of electrons throughout the entire metal structure due to the strong attraction to the metal **cations**.

Metallic lattice

Chemists believe that a solid sample of a metal can be explained by the following ideas.

- Positive ions are arranged in a closely packed structure. This structure is described as a regular, three-dimensional network of positive ions. The cations occupy fixed positions in the lattice.
- Negatively charged electrons move freely throughout the lattice. These electrons are called **delocalised electrons** because they belong to the lattice as a whole, rather than staying in the shell of a particular atom.
- The delocalised electrons come from the outer shells of the atoms. Inner-shell electrons are not free to move throughout the lattice and remain firmly bonded to individual cations.
- The positive cations are held in the lattice by the electrostatic force of attraction between these cations and the delocalised electrons. This attraction extends throughout the lattice and is called **metallic bonding**.

Together, these ideas make up the metallic bonding model. An example of how a metal, such as sodium, could be represented using this model is shown in Figure 6.2.2.

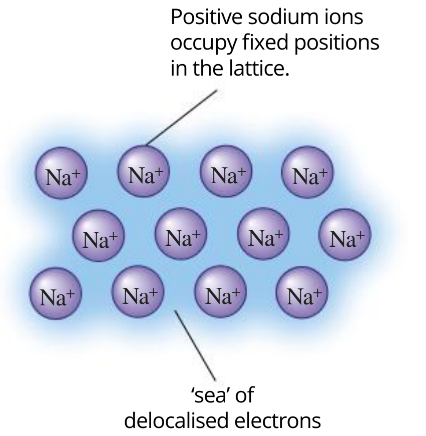
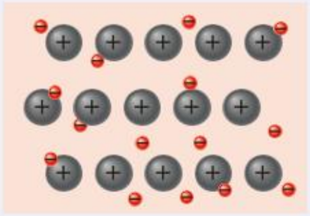
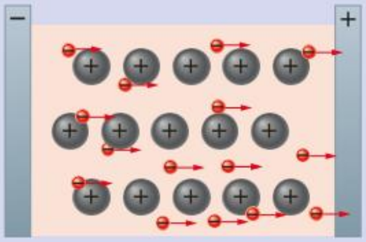
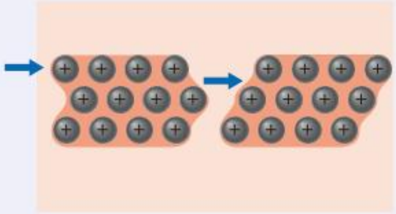


FIGURE 6.2.2 A representation of a sodium metal lattice. Each sodium atom loses its one valence electron. This electron is shared with all atoms in the lattice to form a sea of delocalised electrons.

In the metallic bonding model, positive metal cations are surrounded by a 'sea' of delocalised electrons.

Table 6.2.2 shows how the metallic bonding model is consistent with the relatively high boiling point, electrical conductivity, malleability and ductility of metals.

TABLE 6.2.2 Physical properties of metals and explanations from metallic bonding model

Property	Explanation	
Metals are hard and have relatively high boiling points.	Strong electrostatic forces of attraction between positive metal ions and the sea of delocalised electrons hold the metallic lattice together.	
Metals are good conductors of electricity.	Free-moving delocalised electrons will move towards a positive electrode and away from a negative electrode in an electric circuit.	
Metals are malleable and ductile.	When a force causes metal ions to move past each other, layers of ions are still held together by the delocalised electrons between them.	

Other properties of metals

Metals generally have a high density. The cations in a metal lattice are closely packed. The density of a metal depends on the mass of the metal ions, their radius and the way in which they are packed in the lattice.

Metals are good conductors of heat. When the delocalised electrons bump into each other and into the metal ions, they transfer energy to their neighbour. Heating a metal gives the ions and electrons more energy and they vibrate more rapidly. The electrons, being free to move, transmit this energy rapidly throughout the lattice.

Metals are lustrous. Because of the presence of free electrons in the lattice, metals reflect light of all wavelengths and appear shiny.

Metals tend to react by losing electrons. The delocalised electrons in metals may participate in reactions anywhere on the metal's surface. The **reactivity** of a metal depends on how easily electrons can be removed from its atoms. This was covered in more detail in Chapter 4.

Limitations of model

Although this model of metallic bonding explains many properties of metals, some cannot be explained so simply. These include the:

- range of melting points, hardness and densities of different metals
- differences in electrical conductivities of metals
- magnetic nature of metals such as cobalt, iron and nickel.

In order to deal with such questions, you would need a more complex model of metallic bonding, which is beyond the scope of this book.

Worked example 6.2.1

ELECTRON CONFIGURATION OF ALUMINIUM

With reference to the electron configuration of aluminium, explain why solid aluminium can conduct electricity.	
Thinking	**Working**
Using the atomic number of the element, determine the electron configuration of its atoms. (You may need to refer to a periodic table.)	Al has an atomic number of 13 and an electron configuration of $1s^22s^22p^63s^23p^1$.
From the electron configuration, find how many outer-shell electrons are lost to form cations that have a stable noble gas electron configuration. These electrons become delocalised.	Al has three electrons in its outer shell (the $3s^23p^1$ electrons). Al atoms will tend to lose these three valence electrons to form a cation with a charge of +3. The outer-shell electrons become delocalised and form the 'sea' of delocalised electrons within the metal lattice.
An electric current occurs when there are free-moving charged particles.	If the Al is part of an electric circuit, the delocalised electrons are able to move through the lattice towards a positively charged electrode.

WS 1.2.3

➤ Try yourself 6.2.1

ELECTRON CONFIGURATION OF MAGNESIUM

With reference to the electron configuration of magnesium, explain why solid magnesium can conduct electricity.

6.2 Review

SUMMARY

- Metallic bonding is the electrostatic force of attraction between a lattice of positive ions and delocalised valence electrons. The lattice of cations is surrounded by a 'sea' of delocalised electrons.
- The metallic bonding model can be used to explain the properties of metals, including their malleability, thermal conductivity, generally high melting point and electrical conductivity.

KEY QUESTIONS

Retrieval

1 List three properties common to most metals.

2 Recall, using the metallic bonding model, how metals can be lustrous.

Comprehension

3 The properties of calcium mean that it is classed as a metal.
  - **a** Draw a diagram to represent a calcium metal lattice.
  - **b** Describe the forces that hold this lattice together.

4 Barium is an element in group 2 of the periodic table. It has a melting point of 850°C and conducts electricity in the solid state. Describe how the properties of barium can be explained in terms of its bonding and structure.

5 Graphite is a non-metallic substance that can be lustrous and conducts electricity and heat. It is not malleable, but breaks if a force is applied.
  - **a** List the properties that graphite shares with metals.
  - **b** Describe the structure of graphite given it shares these properties with metals.

Analysis

6 A very useful property of metals is that they are malleable.
  - **a** Distinguish what specific part of the metallic bonding model enables metals to have this property.
  - **b** Analyse, with the aid of a diagram, what happens to a solid metal when a force is applied.

7 Determine why copper can conduct electricity, with reference to its electron configuration.

8 Aluminium is a group 13 metal. With reference to the electronic structure of aluminium:
  - **a** identify how many electrons it will donate to its 'sea' of delocalised electrons.
  - **b** compare its ability to withstand a force being applied to how sodium would withstand the same force.

6.3 Modifying metals

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- ➤ recognise that metals can be modified so that their physical and chemical properties can be altered
- ➤ compare the types and amounts of the substances that make up the modified metal to see what desirable properties can be obtained.

The way in which a metal can be used is determined by the metal's physical and chemical properties. Although some metals are valuable in their pure state, most metals need to be modified to make them more useful.

Modified metals have a wide range of applications. In Figure 6.3.1, you can see a number of uses of modified metals. In each case, a metal with desirable properties has been chosen and improved by the processes covered in this module.

A metal can be modified in one of three main ways:

- through alloy production
- by heat treatment
- with a coating.

FIGURE 6.3.1 (a) Modified metals are used in aeroplanes, which need to be made from materials that are strong, durable and light. (b) The Oscar statuettes are made of britannium (a mixture of tin, copper and antimony) and then coated successively in layers of copper, nickel, silver and gold. (c) Blacksmiths use heat and hammering tools to modify the properties of the metal used for horseshoes.

MAKING ALLOYS

Often, metals are mixed with small amounts of another substance, usually a metal or carbon. The substances are melted together, mixed and then allowed to cool. The resultant solid is an alloy.

By varying the composition of alloys, you can obtain materials with specific properties. Generally, an alloy is harder and melts at a lower temperature than the pure metal. This is because atoms of different sizes are now included in the metal lattice. As these atoms do not pack in the same way as the main metal, they will not allow the lattice to shift and bend in the same way. This disruption of the regular metallic lattice also accounts for the lowered melting point.

When other metals or carbon are added to melted metals, the substance that is produced is called an alloy. Alloys tend to have improved properties compared to the original metal, making them more useful in society.

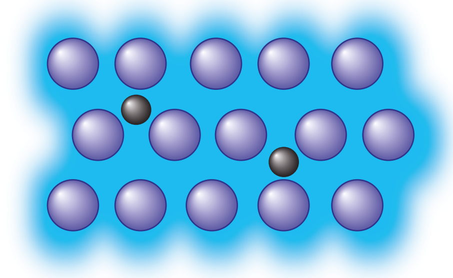


FIGURE 6.3.2 Steel is an interstitial alloy of iron and carbon. Note the relative sizes and position of the atoms in an interstitial alloy.

Alloys of iron: steel

Currently, almost all of the iron mined around the world is used to make the alloy steel. The simplest steel is made by adding a small amount of carbon to iron to make carbon steel.

Carbon steel is a type of **interstitial alloy**. In interstitial alloys, a small proportion of an element with significantly smaller atoms is added to a metal. The added atoms sit in interstices (very small spaces) between metal cations in the metallic lattice, as shown in Figure 6.3.2.

Carbon steel is generally harder and less malleable than pure iron. Varying the amount of carbon in the mixture produces steels with different properties, as shown in Table 6.3.1. This allows the steel with the best properties to be used in specific applications.

TABLE 6.3.1 A summary of the properties and uses of some different carbon steels

Type of carbon steel	Percentage of carbon	Properties	Typical use
low-carbon steel	less than 0.3%	strong, easily shaped	bridges, buildings, ships and vehicles
medium-carbon steel	0.3–0.45%	increased hardness and tensile strength, decreased ductility	large machinery parts
high-carbon steel	0.45–0.75%	very strong, more brittle	springs and high-strength wires
very high-carbon steel	up to 2.5%	hard, more brittle	cutting tools

Other elements can also be added in addition to carbon to make steels with improvements to different properties. Examples of some of these are listed in Table 6.3.2.

TABLE 6.3.2 Some elements that are alloyed with iron to make steels

Alloying element	General effects on properties	Example
manganese	• increases strength and toughness	bicycle frames
chromium	• increases hardness and tensile strength • resists corrosion	stainless steel for cutlery, kitchen sinks
nickel	• increases toughness, tensile strength and hardness • resists corrosion	stainless steel
cobalt	• improves magnetic properties • resists high temperatures	alnico magnets, jet propulsion engines

The steels in Table 6.3.2 are not considered to be interstitial alloys, as the added atoms are too big to fit into the spaces in the metallic lattice. The atoms of the elements added replace some of the iron cations and the mixture is called a **substitutional alloy**. In general, the metallic elements added to make substitutional alloys have fairly similar chemical properties and form cations of a similar size to the main metal.

Stainless steel is a substitutional alloy of nickel and chromium. The nickel and chromium atoms take the place of some of the iron atoms in the lattice (Figure 6.3.3). All the metal cations are attracted to the sea of electrons, so the lattice is still strongly bonded. However, because the different kinds of atoms are slightly different in size, the layers within the lattice cannot move as easily past each other. This makes the alloy harder and less malleable than pure iron.

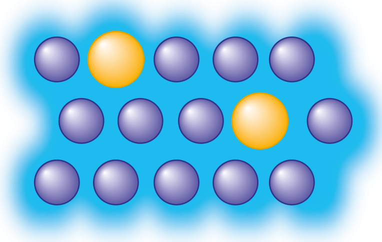


FIGURE 6.3.3 The iron, nickel and chromium atoms in stainless steel are relatively similar in size. The alloy is harder than any of these metals alone.

Common alloys

Many common materials are made of alloys in order to make the metal more suited to a particular purpose. Metals are blended together to combine properties, such as strength, colour, reflectivity and chemical stability. The resulting alloy will have different properties from the individual elements. A number of examples of substitutional alloys are shown in Table 6.3.3.

TABLE 6.3.3 Substitutional alloys found in everyday objects

Alloy	Metals used	
Australian 20-cent coin	copper and nickel	
Australian $2 coin	copper, aluminium and nickel	
dental mercury amalgam	mercury and zinc	
brass	copper and zinc	
bronze	copper and tin	

WORK HARDENING AND HEAT TREATMENT

The way a metal object is prepared also affects how it behaves. Many metals are prepared in the liquid state and then cooled. The rate at which a metal is cooled affects the properties of the solid.

The model that you have been developing for the structure of metals describes the arrangement of particles within a single metal **crystal**. A crystal is a region in a solid where the particles are arranged in a regular way. A sample of solid metal consists of many small crystals. Each crystal is a continuous regular arrangement of cations surrounded by a sea of delocalised electrons, but the arrangement of individual crystals with respect to one another is random like those shown in Figure 6.3.4. As you can see, the regular lattice is disrupted at the point where one crystal meets another.

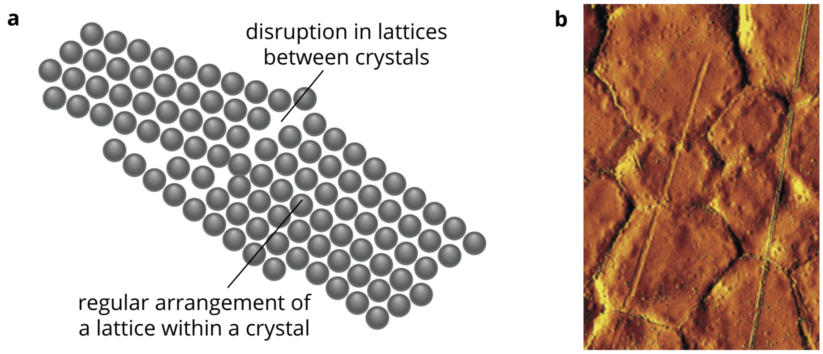


FIGURE 6.3.4 (a) Metals are made of a number of crystals, which have boundaries between them. (b) An electron micrograph shows the individual crystals that formed as the molten metal mixture cooled down and solidified.

The way a metal behaves—its malleability and brittleness—will depend on the size and arrangement of the crystals. Generally, smaller crystals result in harder metals as there is less free movement of layers of cations over each other. Smaller crystals also have more areas of disruption between them, and this usually means that these metals will be more brittle.

The crystal structure of metals can be altered in a number of ways. Two of these ways are **work hardening** and **heat treatment**.

FIGURE 6.3.5 The crystal structure of metals like this gold ring can be altered by work hardening. Care is needed as the gold becomes more brittle through this process.

Work hardening

Hammering or working cold metals causes the crystals to rearrange as they are pushed and deformed. This can result in the hardening of the metal as the crystals are flattened out and pushed closer together. Figure 6.3.5 shows a gold ring being hammered on a round tool, called a triblet, to harden and strengthen it.

Paperclips are also manufactured by work hardening processes. If you bend it once, it remains fairly pliable, but if you bend it backwards and forwards several times, it snaps. Bending rearranges the crystal grains, making the metal harder but more brittle.

Heat treatment

The physical properties of a metal can be altered by controlled heating and cooling. The three main methods of heat treatment (**annealing**, **quenching** and **tempering**) of metals and the effects on the metal's properties and structure are summarised in Table 6.3.4.

TABLE 6.3.4 Summary of the heat treatment methods and the effect on metal properties

Treatment	Process	Effect on metal structure	Effect on metal properties
annealing	A metal is heated to a moderate temperature and allowed to cool slowly.	Larger metal crystals form.	The metal is softer with improved ductility.
quenching	A metal is heated to a moderate temperature and cooled quickly (sometimes by plunging into water).	Tiny metal crystals form.	The metal is harder and brittle.
tempering	A quenched metal is heated (to a lower temperature than is used for quenching) and allowed to cool.	Crystals of intermediate size form.	The metal is hard but less brittle.

When metals are heated above a critical temperature, the individual crystals merge. When the metal is allowed to cool, the crystals re-form. The rate of cooling determines how large the new crystals will be. Faster cooling leads to smaller crystals; slower cooling allows more time for crystals to grow larger.

Steels respond well to heat treatment. Annealing steels reduces their strength or hardness, increases uniformity of crystals and so reduces stresses and restores ductility. Hardening of steels by quenching and tempering increases their strength and wear properties (Figure 6.3.6).

FIGURE 6.3.6 Japanese swordsmiths had to understand the properties of alloys and the effects of heat-treating and work hardening as they were working with low-quality steel.

COATINGS

Metals can be given a coating to make them even more suitable for their intended application. The coating can be decorative, functional or both. For example, stainless steel frying pans (Figure 6.3.7) are strong, hard and durable, have a high melting point and conduct heat very well. Coating the frying pan with the polymer Teflon makes the frying pan non-stick and more resistant to rust.

Iron and the other elements that make up steel alloys react with air and other substances, which can damage the steel object. Surface coatings, such as tin (Figure 6.3.8), paint and powder coating, protect the steel by forming a physical barrier on the surface to prevent exposure to damaging chemicals.

FIGURE 6.3.7 A stainless steel frying pan with a non-stick Teflon coating

FIGURE 6.3.8 'Tin' cans constructed from steel with a thin coating of tin make the can resistant to rusting.

Powder coating is a method of applying a decorative and protective finish. The powder is a mixture of finely ground particles of pigment and resin, which is sprayed onto a surface electrostatically (Figure 6.3.9). The electrically charged powder particles adhere to the surface. When heated, the particles fuse to form a smooth surface that is resistant to scratching and marking.

FIGURE 6.3.9 Powder coating uses fine particles of pigment and resin to protect the metal underneath.

Another approach involves coating steel with metals that chemically protect the surface. A common example is galvanising (Figure 6.3.10). Galvanised steel has a protective coating of zinc, which prevents rusting. The zinc coating reacts with oxygen and carbon dioxide in the air to slow corrosion.

The surface coatings applied to a car body consists of many layers that offer both physical and chemical barriers to protect the steel from rusting (Figure 6.3.11).

FIGURE 6.3.10 These galvanised nails have been coated with a protective layer of zinc.

FIGURE 6.3.11 The special paint used on cars is a decorative and functional coating.

6.3 Review

SUMMARY

- Most metals can be modified to alter their properties and make them more useful.
- An alloy is a mixture of a metallic element with other elements.
- Alloys are often harder and melt at a lower temperature than pure metals.
- Heat treatment changes the size and arrangement of crystals in metals.
- Different coatings can be applied to metals to provide additional decorative and/or functional properties.

KEY QUESTIONS

Retrieval

1 State the difference between a substitutional alloy and an interstitial alloy.

2 List three reasons why metals might need to be coated.

3 Recall the alloys that make up a 20-cent coin and high-carbon steel.
  a List the elements present.
  b Draw a diagram to show how the atoms of these elements are arranged in the alloy.

4 Describe the process of powder coating a metal.

5 Alloys are modified metals. In each of the following examples, identify the metal used and explain a property that makes the alloy different from the original metals.
  a an Australian 20-cent coin
  b solder
  c 9-carat gold
  d stainless steel
  e dental mercury amalgam

Comprehension

6 Horseshoes are often made from steel that is worked into shape by a process of heating and hammering the metal. Explain how this process results in a better horseshoe than one simply made of iron.

Analysis

7 Dentists fit partial dentures by means of small metal hooks. The hooks attach the denture to the remaining teeth. The hooks are easily bent at first to fit snugly in individual mouths. However, if the hooks are bent backwards and forwards too often, then they become brittle and snap. Deduce why this happens.

8 Before shaping aluminium into objects by hammering, it is usually annealed to make it more malleable. The final product of this shaping process is then usually quenched to increase its strength. Use the metallic bonding model to determine why:
  a annealing makes aluminium more malleable
  b quenching can increase the strength of the final aluminium object.

9 Compare the benefits in industry for a metal to be coated rather than heat-treated for durability.

10 A bicycle developer decides that manganese is too expensive for the frames he wishes to make. He looks at a range of metals as other options for the steel frame. For each metal listed below, determine the effects the addition of the metal to iron might have and whether these properties would be useful for the bike frame.
  a chromium
  b nickel
  c cobalt
  d lithium
  e aluminium

11 For each of the following alloys, determine the alloy type (interstitial or substitutional).
  a copper/beryllium
  b iron/chromium
  c aluminium/magnesium

12 Determine which of the properties listed below match the alloys provided.
    i high strength, lightweight
    ii high strength, excellent electrical conductivity
    iii increased strength, increased resistance to corrosion
  a copper/beryllium
  b iron/chromium
  c aluminium/magnesium

6.4 Extraction of iron from its ore

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- remember the chemical formula of Australia's most common iron ore, haematite, Fe_2O_3
- describe the smelting process for the production of iron
- understand the process for extracting iron from its ore
- recognise that the mining and extraction processes to obtain iron raise environmental, economic and social issues.

Australia is the world's largest exporter of iron **ore** (a natural compound containing a metal). Australia exported a record 757 million tonnes of iron ore in 2015–16, which was worth around A$47.8 billion.

Modern society is very dependent on iron. About 98% of world iron production is used to make steel. The steel in turn is used in bridges, buildings and all forms of transport. It also has many other uses as shown in Figure 6.4.1.

FIGURE 6.4.1 Steel is used in (a) the Sydney Harbour Bridge, (b) building frames in construction, (c) train tracks and (d) surgical instruments.

Deposits of iron ore occur in all Australian states and territories. Most of the identified deposits—almost 93% (totalling 64 billion tonnes)—are found in Western Australia. Massive deposits of iron ore in the Pilbara region of Western Australia are mined by open-cut methods.

Iron ore is composed mainly of **iron oxide** combined with rocky material. The iron must be extracted from the ore before it can be used to make steel.

There are various ways of extracting a metal from its ore. This module will examine, as a case study, the main method used to extract iron from its ore. It will also examine the environmental, economic and social issues associated with this method of extraction.

EXTRACTING IRON FROM ITS ORE

The raw materials used in the extraction of iron are:

- iron ore
- coke
- limestone
- air.

Iron ore

Iron ore is a **mineral** that occurs in the Earth's crust. Minerals are naturally occurring solid substances with a definite chemical composition, structure and properties.

Most deposits of iron ore occur in sedimentary rocks, which are more than 600 million years old. They formed from chemical reactions that combined iron and oxygen in marine and fresh waters.

In Australia, the iron oxides in iron ore are usually in the form of **haematite** (Fe_2O_3) (Figure 6.4.2). This mineral is dark and reflective when polished, but a red layer forms on its surface after exposure to oxygen. Many of Australia's deserts have red sands due to the presence of this red oxide.

Other common iron minerals are magnetite (Fe_3O_4) and pyrite (FeS_2). Magnetite is a black mineral that has been known since ancient times as lodestone. It is magnetic and was used to make early compasses. Pyrite is a shiny, yellow-coloured mineral that is often referred to as 'fool's gold' due to its striking resemblance to gold (Figure 6.4.3). Pyrite is used extensively as a source of sulfur by the chemical industry.

Iron ore is mined from the ground and taken in trucks for processing. In Australia, this is done by open-cut mining methods. A haul truck (large dump truck) capable of carrying up to 400 tonnes is shown moving iron ore within a mine in Figure 6.4.4.

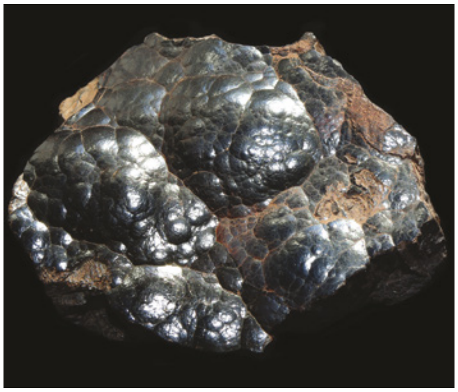


FIGURE 6.4.2 The mineral haematite (Fe_2O_3) is mined as the main source of iron. This sample has been partially polished, removing most of the red oxide coating that forms when the iron is exposed to air.

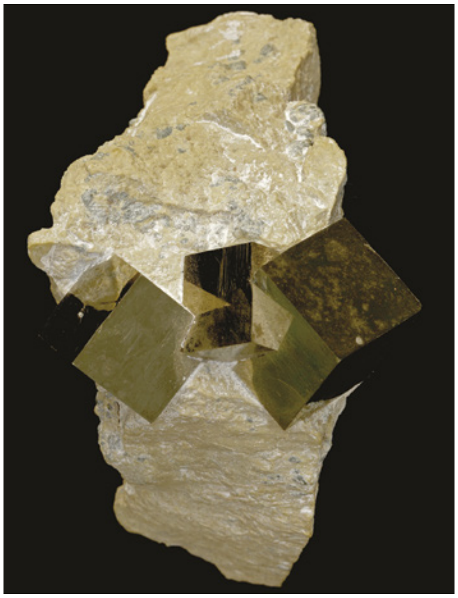


FIGURE 6.4.3 Cubic crystals of pyrite

FIGURE 6.4.4 Iron ore is sourced from open-cut mines in Australia.

Coke

Coke is a solid that contains 80–90% carbon. The coke used for iron production is made by strongly heating coal in air-tight ovens for about 15 hours. This is performed on the same site where iron is extracted from the ore. A number of useful by-products are obtained from this process, including:

- fuel gas, which is burnt to provide heat in various furnaces on the site
- tar, benzene, toluene and ammonium sulfate, which are used as the raw materials for other chemical industries.

Limestone

Limestone is a sedimentary rock that is mainly composed of calcium carbonate. Limestone is obtained by mining or quarrying and is crushed before use.

THE BLAST FURNACE

FIGURE 6.4.5 A blast furnace used to extract iron from iron ore. Iron ore is added continuously to the top of the furnace by the conveyor belts on the left-hand side.

Extraction of iron from ore and the removal of unwanted materials is carried out in a tall, bottle-shaped tower called a **blast furnace**. A building containing a modern blast furnace and external conveyer belts is shown in Figure 6.4.5. These furnaces are heated to very high temperatures and are operated continuously for many years.

Figure 6.4.6 shows a diagram of the inside of a blast furnace. Pre-heated air is blasted into the bottom part of the furnace while solid 'charges' (scoops) of iron ore, coke and limestone are continuously added to the top. The construction of the furnace causes different temperature zones where different reactions can take place.

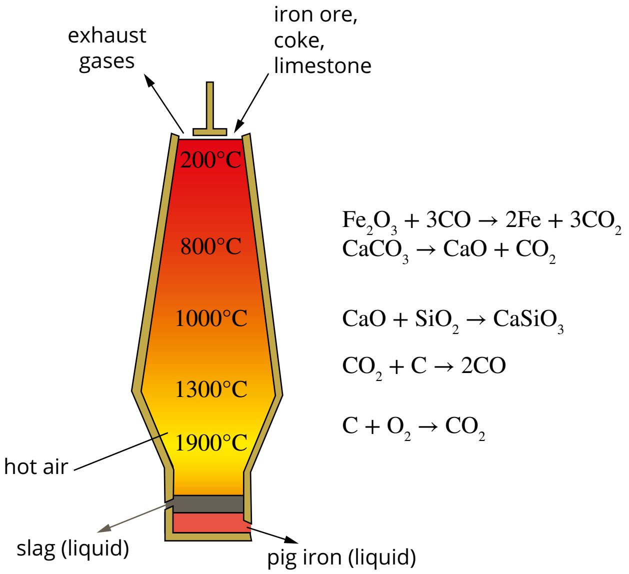


FIGURE 6.4.6 A diagram of a blast furnace. Hot air jets blast air in at the bottom. Iron ore, coke and limestone are added at the top.

Coke reacts with oxygen

As the air rises through the furnace and meets the descending charge, oxygen reacts with coke to produce carbon dioxide, which then reacts further with the coke to produce carbon monoxide. These reactions are represented by the following equations:

carbon + oxygen → carbon dioxide

$$C + O_2 \rightarrow CO_2$$

carbon dioxide + carbon → carbon monoxide

$$CO_2 + C \rightarrow 2CO$$

The first reaction releases considerable heat energy and helps to maintain the high temperature that provides the fast rate of reaction required for metal production in the furnace. No external source of heat is needed to achieve temperatures up to 1900°C within the furnace.

Carbon monoxide reacts with iron ore

Extraction of iron from iron oxide occurs in a series of steps in which carbon monoxide is the main reactant. The steps occur in different temperature zones within the blast furnace. Iron ore (Fe_2O_3) is first converted to another iron oxide (Fe_3O_4), then another (FeO) and finally to the metal (Fe). The process is summarised by the single chemical equation:

$$\text{carbon monoxide} + \text{iron oxide} \rightarrow \text{iron} + \text{carbon dioxide}$$

$$3CO + Fe_2O_3 \rightarrow 2Fe + 3CO_2$$

Function of limestone in the blast furnace

Iron ore is not pure iron oxide. It also contains rocky material such as silica (silicon dioxide), alumina and manganese oxides. These would not melt in the heat of the furnace and would eventually clog it up.

The limestone ($CaCO_3$) added to the top of the furnace breaks down in the furnace to form calcium oxide (CaO) and carbon dioxide. Calcium oxide can react with the unwanted materials to form a new compound called **slag**. Most of the slag is calcium silicate ($CaSiO_3$) due to the high amount of silica in iron ore.

Both the iron and slag are **molten** (melted) and sink to the bottom of the furnace. The slag is less dense and floats on top of the molten iron. This provides a cover that prevents the iron from reacting with the incoming air and re-forming iron oxide.

Collection of the iron

Holes at the base of the furnace are opened and the molten iron and slag are drained out and separated. In a steel works, the iron is usually transferred, while still molten, directly to a steel-making furnace. The slag may be used as a road-surfacing material or to manufacture cement.

ENVIRONMENTAL, ECONOMIC AND SOCIAL ISSUES

Modern society is literally built on and driven by iron and steel. Since the early 1960s, iron ore mining and export has shaped many aspects of life in Australia.

Mining and metal extraction processes raise environmental, economic and social issues. Some issues involved with the mining of iron ore and the production of iron are listed in Table 6.4.1.

TABLE 6.4.1 Some of the environmental, economic and social issues associated with the extraction of iron from iron ore

Environmental issues	Economic issues	Social issues
• loss of landscape due to mining, processing and transporting iron ore, coke and limestone • air pollution from extraction process • disposal of slag • noise pollution	• significant financial benefit for Australia • significant gaps can occur in wages paid to mining staff and those who work locally in service industries	• negotiations with Indigenous groups to ensure fair access to iron-rich land • land use conflicts between mining, agriculture and tourism • iron ore mines are located in remote locations. Workers usually fly to mines to work for a number of days before flying home, which has an impact on their family life

6.4 Review

SUMMARY

- The raw materials used in the extraction of iron are iron ore, coke, limestone and air.
- Iron ore is a mixture of iron oxides and other rocky material.
- Extraction of iron from iron ore occurs in a blast furnace. A series of reactions occur involving the iron ore, air and coke:

  carbon + oxygen → carbon dioxide

  $$C + O_2 \rightarrow CO_2$$

  carbon dioxide + carbon → carbon monoxide

  $$CO_2 + C \rightarrow 2CO$$

  carbon monoxide + iron oxide → iron + carbon dioxide

  $$3CO + Fe_2O_3 \rightarrow 2Fe + 3CO_2$$

- Limestone ($CaCO_3$) reacts with unwanted substances in the iron ore to produce slag, which is removed.
- Molten iron is collected from the bottom of the blast furnace.
- There are many environmental, economic and social issues associated with the mining and extraction of iron from ore.

KEY QUESTIONS

Retrieval

1 State the word equation and chemical equation for the production of iron from the reaction of carbon monoxide and the ore magnetite.

2 Identify the four raw materials used in the extraction of iron from its ore and explain the role of each.

Comprehension

3 One of the very important issues with the production of iron in a blast furnace is environmental pollution. Identify one source of gas pollution and represent this with a word equation in which it is produced.

4 Explain why iron is not found in its pure form in nature.

Analysis

5 In Australia, three iron oxides are commonly found: haematite, magnetite and pyrite.

  a Identify the formula and provide a brief description of each of these iron oxides.

  b The mineral magnetite has a higher iron content than the mineral haematite. However, while hematite ore generally contains large concentrations of hematite, magnetite ore generally holds low concentrations of magnetite. As a result, magnetite ore must be concentrated before it can be used to produce steel. Identify a property of magnetite that will readily enable the concentration and removal of magnetite from the mined ore.

6 When choosing a suitable location for the production of iron, miners consider other factors as well as the availability of iron ore. Explore what other factors might affect the selection of a suitable site and explain why they are important.

7 Classify each of the following statements about iron extraction in a blast furnace as true or false. Rewrite any false statements so that they become true.

  a Slag is denser than iron.

  b Iron ore is a rock that contains purely iron oxide.

  c Limestone is a source of calcium carbonate.

  d Air is pre-heated before being added through the bottom of the furnace.

  e The furnace must be continually heated by external sources to achieve the temperatures required to extract iron.

8 Identify useful by-products of the iron extraction process from iron ore and describe how they are useful.

6.5 Metallic nanomaterials

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- identify that metallic nanoparticles have different properties to large-scale metals
- describe the uses of metallic nanoparticles in health care, transport, agriculture and many other areas.

Metallic **nanomaterials** consist of metal atoms arranged to make nano-sized structures. Metallic nanomaterials have very different properties when compared to a bulk sample of the same type of metal. Scientists have long been fascinated by the effects of nanomaterials. For example, gold–silver alloyed **nanoparticles** change the colour of glass (Figure 6.5.1).

Emerging technologies have allowed scientists to manipulate metals at the nanoscale. This module will introduce you to some of the different forms of metallic nanoparticles, as well as some examples of their current applications.

FIGURE 6.5.1 As light passes through this stained glass window containing gold–silver alloyed nanoparticles, it causes the glass to change colour.

FORMS OF NANOMATERIALS

Nanomaterials exist in many different forms at the nanoscale level, just as they do on the macroscale level. These forms include:

- particles
- rods
- wires
- tubes.

Metallic nanoparticles

Nanoparticles range from about 1 to 100 nm in size (where 1 nm = 10^{-9} m). They have a very high surface area to volume ratio compared to the bulk material of similar volume.

The gold nugget shown in Figure 6.5.2 consists of many billions of cations and delocalised electrons. The properties of gold and other metals, such as electrical and thermal conductivity and metallic lustre, are explained in terms of the movement of delocalised electrons.

FIGURE 6.5.2 Bulk gold has a characteristic colour and typical metallic properties.

FIGURE 6.5.3 The red colour in this stained glass window is caused by gold nanoparticles trapped in the glass. The deep yellow colour is caused by silver nanoparticles.

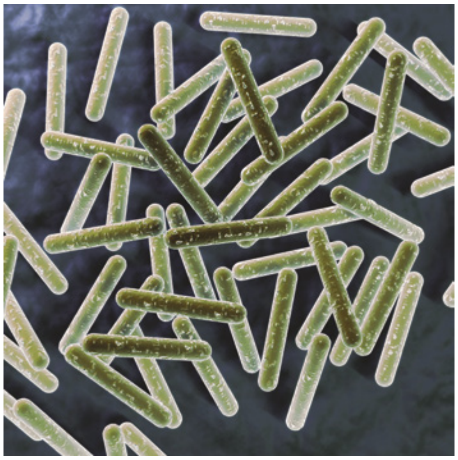


FIGURE 6.5.4 A group of nanorods

However, scientists have found that in a metal crystal with less than a certain number of atoms, the number of valence electrons released is too small to behave as a 'sea' of delocalised electrons. Metal nanoparticles that contain only 100 or so atoms have properties in-between those of metals and non-metals. In general, the nanoparticles have different optical properties and are more sensitive to heat.

For example, gold takes on a ruby colour when the gold particles are reduced to nano-size. You can see this in the stained glass window in Figure 6.5.3. The deep yellow colour is caused by silver nanoparticles.

Metallic nanorods

Nanorods are nanoscale rods in which each dimension ranges from 1 to 100 nm. They have a length to width ratio of 3:1–5:1 (Figure 6.5.4). Nanorods of metals such as gold and silver have been synthesised in the laboratory. Their applications are diverse, including in display technologies and microelectronics, powering everything from solar cells to mobile phones.

Metallic nanowires

A **nanowire** is a nano-sized wire. The diameter of a nanowire is measured on the nanoscale, but its length is unrestricted. Nanowires differ from nanorods because they are much longer, as shown in Figure 6.5.5. Platinum nanowires have promising applications as catalysts and in electronics.

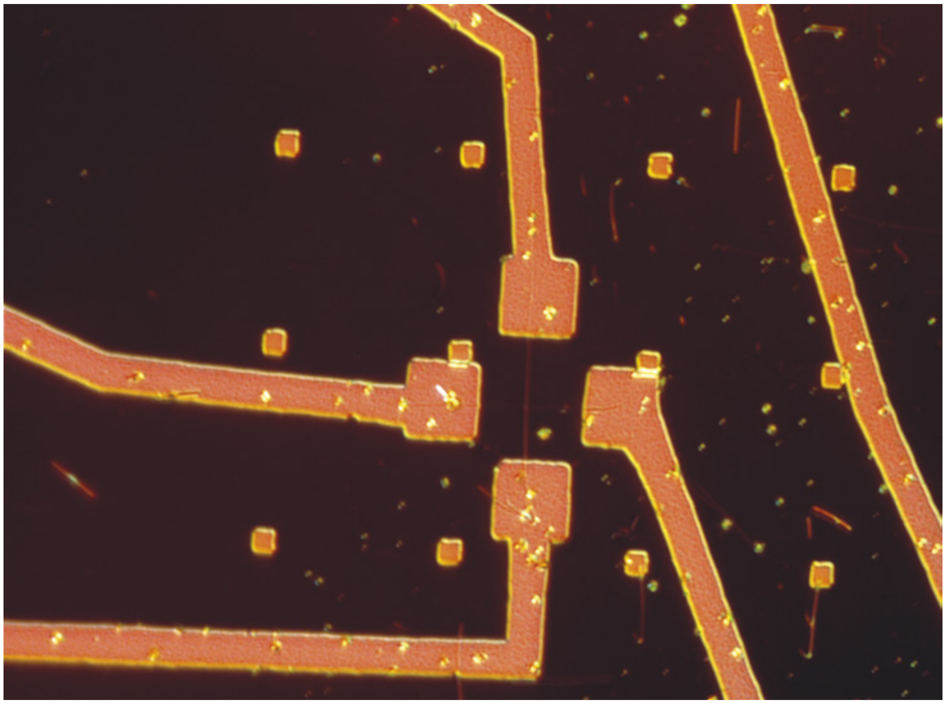


FIGURE 6.5.5 Zinc oxide nanowires

Nanomaterials

Nanomaterials have unique electrical, catalytic, magnetic, mechanical, thermal and imaging characteristics. This makes them attractive for use in medical, pharmaceutical, electronic and engineering sectors.

Gold nanoparticles in cancer treatment

Gold nanoparticles are the subject of substantial research with a wide range of applications (Figure 6.5.6). One area of development is in using gold nanoparticles as a targeted chemotherapy treatment method.

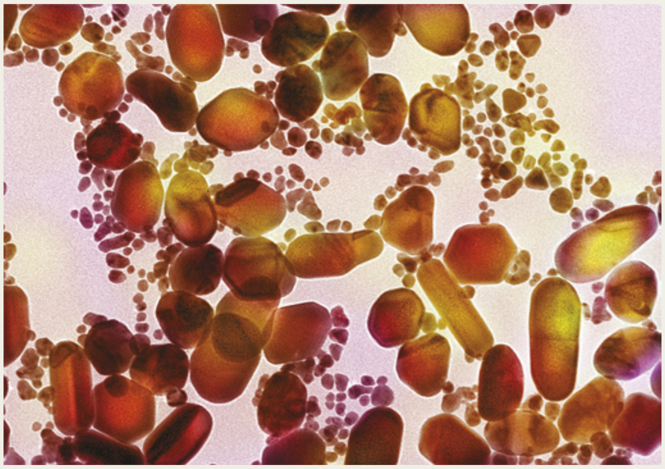


FIGURE 6.5.6 Applications of gold nanoparticles include cancer treatment research and detecting biological toxins.

Gold nanoparticles can be attached to molecules of a tumour-killing agent known as tumour necrosis factor (TNF). The nanoparticles hide the molecule from the body's immune system.

The nanoparticles carrying TNF tend to accumulate in cancer tumours, allowing TNF to destroy tumours. The nanoparticles do not appear to accumulate in other regions of the body, which means healthy cells are not affected.

Silver nanoparticles kill bacteria

Silver ions have long been known to kill bacteria. The ions can rapidly penetrate bacterial membranes and interact with proteins in the bacteria, destroying the cell structure of the bacteria and preventing them from reproducing.

Technology has enabled silver nanoparticles to be included in many different types of wound dressings. When the dressing (Figure 6.5.7) comes into contact with moisture from the wound, silver nanoparticles are slowly but continuously released from the wound pad. They then enter the wound and kill bacteria.

FIGURE 6.5.7 A wound dressing with silver nanoparticles to kill bacteria

In similar antibacterial applications, Samsung has created and marketed a material called Silver Nano, which adds silver nanoparticles to the surfaces of household appliances. Silver nanoparticles have been embedded in the surfaces of plastic storage bins, as well as in fabrics used by astronauts, babies and outdoor enthusiasts.

Copper nanoparticles go into space

Solder is a filler metal used to join two or more metals. Solders are essential to plumbing and metal constructions, including in satellites and spacecraft. For most of history, solders have contained a high amount of lead. Concerns about the toxicity of lead have driven the development of lead-free solder.

The complex electronics in satellites, such as the solar-powered satellite in Figure 6.5.8 on page 162, must be reliable and efficient over a very long time. Space scientists have developed a nanotechnology copper-based solder that offers far superior performance over the materials currently in use. It is expected that the new solder material will produce up to 10 times the electrical and thermal conductivity of current solders, with a wide range of space and defence applications.

FIGURE 6.5.8 A nanotechnology copper-based solder has advantages in situations where long life and reliability are critical.

Iron nanoparticles remove pollution in groundwater

Iron nanoparticles are being used to clean up the pollutant carbon tetrachloride (CCl_4) from groundwater (Figure 6.5.9). Carbon tetrachloride is a manufactured toxic chemical that has been shown to cause cancer in animals. Spills of carbon tetrachloride can spread through soil and create large areas of contamination.

Iron nanoparticles can quickly and effectively break down carbon tetrachloride to a mixture of relatively harmless products. It may be possible to inject the nano-sized iron deep into the ground where it can treat contaminated groundwater.

FIGURE 6.5.9 Nanoparticles of iron are being investigated as a way of eliminating a range of environmental pollutants.

Review

1. Scientists at Swinburne University of Technology in Victoria have shown that gold nanorods can inhibit cancer cell growth in cervical cancer. Explain how the gold nanorods inhibit the cancer cell growth. Determine if this has been shown to affect all body cells.
2. Carbon tetrachloride (CCl_4) is a toxic chemical that has been used in fire extinguishers, as a coolant in refrigeration and as a solvent in dry cleaning. Spills of carbon tetrachloride can spread through soil and create large areas of contamination. Explain how iron nanoparticles can be used to remove and break down this toxic CCl_4.

6.5 Review

SUMMARY

- Nanomaterials range in size between 1 and 100 nm, where 1 nm = 10^{-9} m.
- Metallic nanomaterials are nano-sized particles of metals.
- Different forms of metallic nanomaterials include nanoparticles, nanorods and nanowires.
- Metallic nanomaterials have applications in medical, pharmaceutical, electronic, environmental and engineering sectors.

KEY QUESTIONS

Retrieval

1 Iron nanoparticles must be stored away from oxygen as they will react quickly to produce iron oxide. State a word equation for the reaction of iron nanoparticles with oxygen.

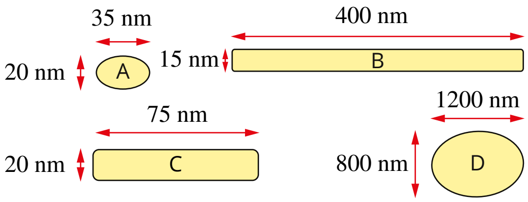
Comprehension

2 A scientist synthesises a new material in the laboratory. It is a solid substance with a diameter of 15 nm and a length of 2300 nm. Determine which form of nanomaterial this substance is.

3 A scientist synthesised a tiny particle for use in solar cells, with a diameter of 8.34×10^{-7} m. Explain if this particle could be classified as a nanomaterial.

Analysis

4 Classify the metallic nanoparticles shown in the illustration by type and size.



5 Determine why you would not expect a sample of pure gold nanoparticles to reflect light.

6 Determine why a nanowire can be longer than 3000 nm but still be considered a metallic nanomaterial.

Chapter review

06

KEY TERMS

alkali metal
alloy
annealing
blast furnace
brittle
cation
coke
conductor
crystal
delocalise
delocalised electron
density
ductile
haematite
heat treatment
interstitial alloy
ionisation energy
iron oxide
lattice
limestone
lustrous
malleable
metallic bonding
metallic bonding model
mineral
molten
nanomaterial
nanoparticle
nanorod
nanowire
ore
quenching
reactivity
slag
steel
substitutional alloy
tempering
tensile strength
transition metal
work hardening

KEY QUESTIONS

Retrieval

1 Select which of the following metals would have similar properties to beryllium.
Ca, Cs, Cu, Pb, Mg, Zn, Sr, K

2 Use the data in Table 6.1.2 on page 139 to answer the following questions.
  a State which metal is the best conductor of heat.
  b Explain why this metal is not used in saucepans.
  c Describe what metals are used to make saucepans.

3 State the meaning of the following terms.
  a iron ore
  b mineral
  c haematite
  d iron oxide

4 The boiling points of three metals—sodium, potassium and calcium—are given in the table below.

Metal	Boiling point (°C)
Na	892
K	760
Ca	1490

  a Recall in which group and period of the periodic table these metals are found.
  b Show the electron configuration for each of the three metals.
  c Use the metallic bonding model to suggest why:
    i sodium has a higher boiling point than potassium
    ii calcium has a much higher boiling point than potassium.

Comprehension

5 Use the metallic bonding model to explain each of the following observations.
  a Copper wire conducts electricity.
  b A metal spoon used to stir a boiling mixture becomes too hot to hold.
  c Iron has a high melting point, 1540°C.
  d Lead has a density of $11.4\,g\,mL^{-1}$, which is much higher than for a non-metal such as sulfur.
  e Copper can be drawn out to form a wire.

6 **a** Describe what is meant by the following terms based on your understanding of the metallic bonding model that is used to explain the structure and bonding of metals.
    i delocalised electrons
    ii a lattice of cations
    iii metallic bonding
  b Explain which electrons are delocalised in a metal.

7 Describe the arrangement of particles in a metal wire and explain how they allow the wire to conduct electricity.

8 Use a diagram to describe what is meant by the term ‘metallic lattice’.

9 Explain why iron is not found in its pure form in nature.

10 Represent the formation of slag (calcium silicate) from silica in a blast furnace using a word equation.

11 Some metals are found as elements in nature; others are found as compounds, combined with other elements such as oxygen and sulfur in ores. Australia has natural reserves of many metals and ores, with mining producing large quantities of metals, including aluminium, copper, gold, iron and silver.

- **a** Show the chemical symbol for each of these metals.
- **b** Determine in which group, period and block of the periodic table each of these metals is found.
- **c** Determine which of these metals are found in nature as elements rather than compounds.
- **d** Identify which of these metals are transition elements.
- **e** Identify which of these elements is the rarest.

Analysis

12 The atomic number of magnesium is 12, of aluminium is 13 and of sodium is 11.

- **a** For each metal, determine the number of valence electrons the atom has, the number of electrons its cation has, and the number of electrons each atom will donate to its sea of delocalised electrons.
- **b** Sort these metals from highest melting point to lowest melting point and explain why you have chosen this order.

13 Identify the four raw materials used in the extraction of iron from its ore and explain the role of each.

14 A student wishes to make an iron needle more malleable. She heats three needles strongly in a Bunsen burner flame. She then treats each needle differently.

- Needle 1 is allowed to cool slowly on the bench.
- Needle 2 is cooled quickly by dropping it into a beaker of cold water.
- Needle 3 is cooled quickly in the beaker of water and then reheated in the flame before being allowed to cool more slowly on the bench.

Sort the resulting needles in order of least malleable to most malleable.

15 Metals have many uses in modern society.

- **a** **i** Identify one metal that is used in large quantities in the building industry.
  - **ii** List the properties that make this metal suitable for this use.
  - **iii** Determine the properties of this metal that might limit its use in buildings.
- **b** **i** Identify one metal that is used in large quantities for electrical wiring.
  - **ii** List the properties that make this metal suitable for use in wires.
- **c** **i** Identify two metals that are used in large quantities in the jewellery trade.
  - **ii** List the properties that make them suitable for this use.

16 **a** Determine the electron configuration of an aluminium atom and the configuration of its most stable cation.

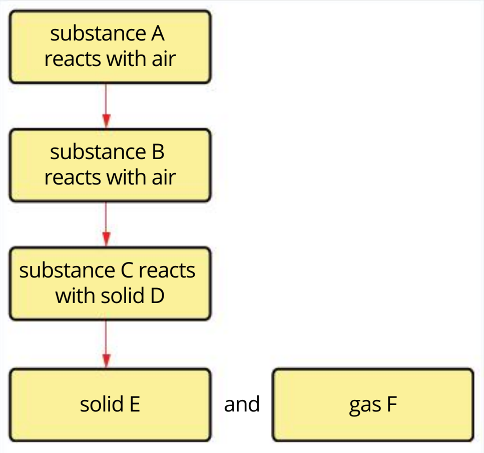
- **b** Identify a metal atom that would make a substitutional alloy with aluminium and explain your choice.

17 Gold nanoparticles have different properties from gold particles that are reduced to nano-size. Compare and contrast these differences and predict why they occur.

18 Apply your knowledge and understanding of the physical and chemical properties of metals to determine which is the most useful property of each of the following metals.

- **a** aluminium
- **b** copper
- **c** iron

19 Consider the following flow chart describing the extraction of iron from iron ore in a blast furnace.



- **a** Determine which two letters would represent carbon dioxide.
- **b** Determine what substance A is likely to be.
- **c** Predict if solid D is likely to be a mixture or an element.
- **d** Describe what other substance is also added to the blast furnace for proper operation.

20 Alloys are modified metals. In each example, identify the metals used and one property that is different from the original metals.

- **a** an Australian \$1 coin
- **b** solder
- **c** 9-carat gold
- **d** stainless steel
- **e** dental mercury amalgam

21 It has been shown using X-rays that metals have a crystalline structure in the solid state.

- **a** Deduce, using the metallic bonding model, why metals form crystals.
- **b** Identify the problems that can be caused by this crystalline structure of metals.

Knowledge utilisation

22 During the manufacture of a metal object, many different modifications may be needed. Devise a series of modifications to produce a steel chisel that has a flexible and strong shaft with a hard blade.

23 Aluminium is the most abundant metal in Earth's crust. Investigate and explain why aluminium was not available before 1886.

24 Metals can be galvanised for protection. This is sometimes referred to as 'sacrificial coating'. Research what galvanised steel is and determine why this process is called 'sacrificial coating'.

CHAPTER

07

Ionic compounds

Rocks, plates, molten lava and electrical insulators belong to a group of substances called ionic compounds. They form the majority of the Earth's crust and, when dissolved, are key components of biological systems. At the end of this chapter, you will be able to explain the structure and properties of these compounds.

Ionic compounds are made by the chemical combination of metallic and non-metallic elements. You will see that their properties are a direct result of the bonding between particles in the compound. The writing of chemical formulas and the naming of ionic compounds are other skills that you will learn in this chapter.

Syllabus subject matter

Topic 1 • Properties and structure of atoms

INTRODUCTION TO BONDING

- recognise that the properties of atoms, including their ability to form chemical bonds, are explained by the arrangement of electrons in the atom and by the stability of the valence electron shell
- understand that the number of electrons lost, gained or shared is determined by the electron configuration of the atom and recall that transitional elements can form more than one ion
- recognise that ions are atoms or groups of atoms that are electrically charged due to an imbalance in the number of electrons and protons and recognise that ions are represented by formulas which include the number of constituent atoms and the charge of the ion
- understand that chemical bonds are caused by electrostatic attractions that arise because of the sharing or* transfer of electrons between participating atoms and the valency is a measure of the number of bonds that an atom can form
- determine the formula of an ionic compound from the charges on the relative ions and name the compound

Topic 2 • Properties and structure of materials

BONDING AND PROPERTIES

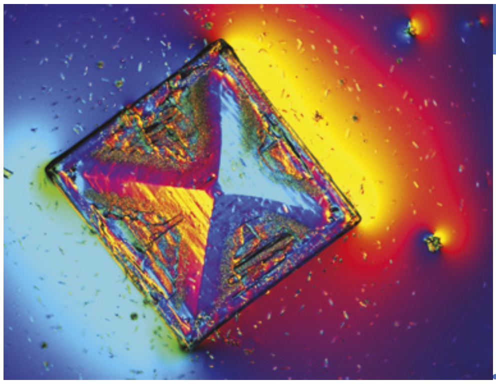
- recognise that the properties of ionic compounds, including high melting point, brittleness, and ability to conduct electricity when liquid or an aqueous solution, can be explained by modelling ionic bonding as ions arranged in a crystalline lattice structure with strong electrostatic forces of attraction between oppositely charged ions (metallic lattice, giant covalent networks, allotropes — carbon)*
- understand that the type of bonding within ionic, metallic and covalent* substances explains their physical properties, including melting and boiling point, thermal and electrical conductivity, strength and hardness
- analyse and interpret given data to evaluate the properties, structure and bonding of ionic, covalent and metallic compounds

SCIENCE AS A HUMAN ENDEAVOUR

- **Nanomaterials:** Development of organic and inorganic nanomaterials is important to meet a range of contemporary needs, including consumer products, health care, transportation, energy and agriculture.

*The greyed-out section of this dot point is addressed explicitly in another chapter.

7.1 Properties and structures of ionic compounds



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recognise that ionic substances have similar physical properties including high melting points, hardness, brittleness, solubility and an ability to conduct electricity when liquid or in aqueous solution but not as a solid
- understand that ionic bonding begins with the transfer of electron(s) between a metal and a non-metal atom
- understand that ionic compounds have a neutral charge and are composed of a crystal lattice made of metal cations and non-metal anions held together by strong electrostatic forces of attraction.

PROPERTIES OF IONIC COMPOUNDS

If you think about the characteristics of rocks (Figure 7.1.1), kitchen crockery and table salt, you will recognise that these materials, and therefore ionic compounds, have some properties in common. Table 7.1.1 lists some properties of a typical ionic **compound** that can be found in materials you might encounter in everyday life. Note that the compounds listed are simple ionic compounds, whereas rocks, ceramics and bricks contain more complex ionic compounds.

Aqueous means that the substance is dissolved in water, and is denoted by the bracketed state symbol (aq) that appears next to the aqueous species in a chemical equation.

Gypsum (calcium sulfate) is a pure ionic compound that is used to manage the chemistry of garden and farm soils. Table salt (sodium chloride) is also a pure ionic compound—**aqueous** sodium chloride is a small but key component of blood plasma, and is found in sport drinks designed to replenish body fluids. In the human body, sodium ions are essential for the conduction of nerve impulses, contraction and relaxation of muscles and maintenance of the body's overall fluid balance.

Data tables describing ionic compounds, such as Table 7.1.1, have allowed chemists to summarise their properties. Generally, ionic compounds have these properties.

- They have a high **melting point** and **boiling point**. They are all solids at room temperature.
- They are hard but **brittle**, unlike metals. Therefore, they are neither malleable nor ductile.
- They do not conduct electricity in the solid state.
- They are good conductors of electricity in the liquid or aqueous state.
- They vary from very soluble to insoluble in water. They are not soluble in **non-polar solvents** such as oil.

The properties of ionic compounds will be further explained in Module 7.5.

FIGURE 7.1.1 Glacial sand deposit containing different types of quartz

TABLE 7.1.1 Properties of typical ionic compounds

Ionic compound	Melting point (°C)	Conductive as solid	Conductive as liquid	Conductive in aqueous solution (0.1 mol L^{-1})	Solubility in water at 25°C (g/100 g water)	Example of commercially available product containing the compound
copper(II) sulfate	decomposes 110	no	yes	yes	22	bluestone spray (used to kill pathogens on fruit)
sodium chloride	801	no	yes	yes	36	food salt
calcium carbonate	1339*	no	yes	–	0.0013	main component in marble
zinc oxide	1975	no	yes	–	insoluble	zinc sunscreen
sodium hydroxide	318	no	yes	yes	114	oven cleaner

*Melting point determined under pressure to prevent decomposition of compound

THE IONIC BONDING MODEL

The properties described above are a result of the arrangement and interaction between particles in the compound. When metallic and non-metallic atoms react to form ionic compounds, the following two steps occur.

1 Metal atoms lose electrons to non-metallic atoms and so become positively charged **metal ions** (called **cations**). There are more protons than electrons and so a **charge imbalance** exists.

2 Non-metal atoms gain electrons from metal atoms and so become negatively charged **non-metal ions** (called **anions**). There are more electrons than protons and so a charge imbalance exists.

$X \rightarrow X^+ + e^-$

where X is any metal element.

$Y + e^- \rightarrow Y^-$

where Y is any non-metal element.

You will remember from Module 3.4 how electrons arrange themselves into shells around the nucleus. Atoms are at their most stable when their valance shell is full. This means there are eight electrons in the valence shell of atoms from period 2 onwards (Chapter 4). An atom's ability to form a cation or anion depends on how many electrons it needs to gain or lose to achieve this stable arrangement.

Newly formed cations and anions then arrange themselves in the following way.

- Large numbers of cations and anions combine to form a **three-dimensional lattice**.
- The three-dimensional lattice is held together strongly by electrostatic forces of attraction between the oppositely charged ions. The forces of **electrostatic attraction** holding the ions together is called **ionic bonding**.
- The attractive force between ions of opposite charge outweighs the repulsive force between ions of the same charge so the lattice is held together quite strongly.

$X^+ + Y^- \rightarrow XY$

Electrostatic forces of attraction exist between positive and negative charges.

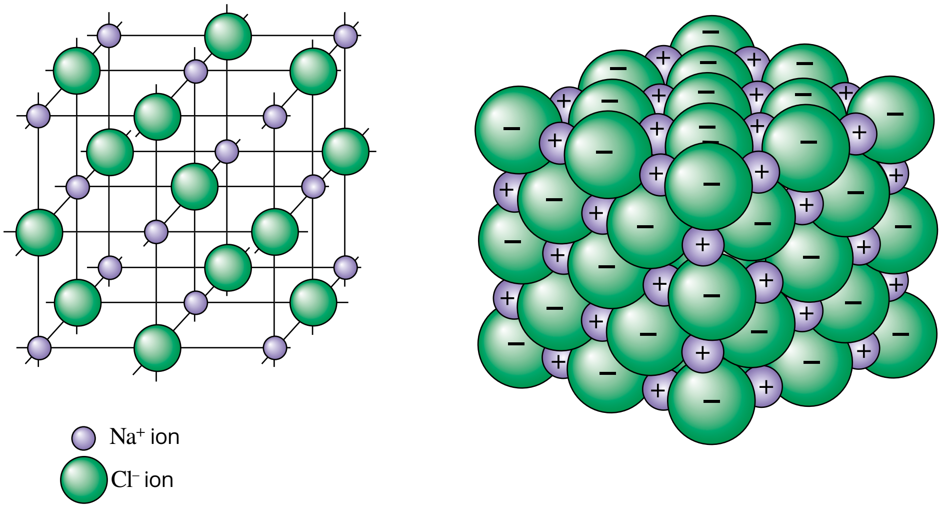
A neutral ionic compound is formed when these charged ions bond.

The formation of ionic compounds is described in more detail in Module 7.2.

The formula of sodium chloride

The **chemical formula** of sodium chloride is written as NaCl (Figure 7.1.2). However, it is important to note that in a solid sample of an ionic compound, such as sodium chloride, individual pairs of sodium and chloride ions do not exist. The solid is not built up of discrete NaCl molecules.

Instead, the solid is made up of a continuous lattice of alternating Na^+ and Cl^- ions. All sodium ions are an equal distance from six chloride ions and all chloride ions are an equal distance from six sodium ions. The chemical formula NaCl is therefore known as the **formula unit**, which shows that the overall ratio of sodium ions to chloride ions in the lattice is 1:1. Writing simple and more complex chemical formulas is described in Modules 7.3 and 7.4 respectively.



WS 1.2.4

FIGURE 7.1.2 Two representations of part of the crystal lattice of the ionic compound sodium chloride (NaCl). Forces of attraction (shown as lines in the first diagram) between oppositely charged ions result in strong bonding. Each positive sodium ion is surrounded by six negative chloride ions and each chloride ion is surrounded by six sodium ions.

7.1 Review

SUMMARY

- Ionic compounds form a crystal lattice made up of metals and non-metals.
- The particles that make up the crystal lattice are cations and anions.
- Each ion has a charge imbalance resulting from loss or gain of electrons
- The three-dimensional lattice is held together strongly by electrostatic forces of attraction between the cations and anions. The electrostatic forces of attraction are called ionic bonding.
- In the case of sodium chloride, in order to maximise the forces of attraction, each sodium ion (Na^+) is surrounded by six chloride ions (Cl^-) and each chloride ion is surrounded by six sodium ions.
- Ionic compounds have properties such as hardness and high melting points.
- Ionic compounds do not conduct electricity in the solid state but are good conductors in the liquid state or when dissolved in water. (Properties are described more fully in Module 7.5.)

KEY QUESTIONS

Retrieval

1 Some properties of four different substances are described below. Identify the substance that is most likely to be an ionic compound.
  - **A** Substance A has a melting point of 842°C and conducts electricity at 700°C.
  - **B** Substance B has a melting point of 308°C. It does not conduct electricity at 250°C but will conduct electricity at 350°C.
  - **C** Substance C has a melting point of 180°C and can be drawn into a wire.
  - **D** Substance D is a white solid that melts at 660°C and will not conduct electricity at 700°C.

2 Identify the two words within the word ‘electrostatic’ and state how this term correctly describes characteristics of anions and cations.

Comprehension

3 Explain why aluminium would be more likely to form a cation than an anion.

4 Using the periodic table on page 10, determine whether each of the following is likely to be an ionic substance.
  - **a** magnesium bromide
  - **b** carbon dioxide
  - **c** sulfur trifluoride
  - **d** lithium oxide

5 Explain why the crystal lattice is held together so strongly in sodium chloride.

6 Explain why ionic compounds are neither malleable nor ductile.

7 A crystal was found to contain 2.15×10^{10} magnesium atoms and 4.3×10^{10} fluorine atoms. Use this information to write the formula unit for this compound.

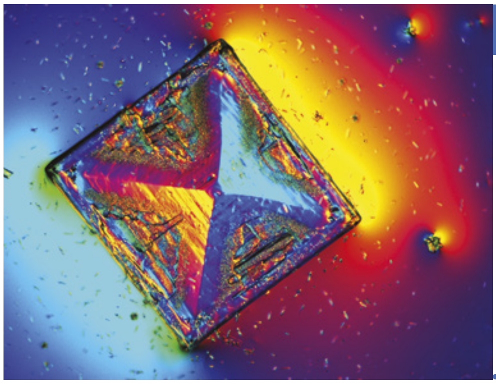
Analysis

8 Tap water is safe for drinking and has dissolved minerals (most of which are dissolved ionic substances) used by the body. Deduce why these substances also make tap water an electrocution hazard.

9 Analyse and explain the trend in melting points of the period 3 chlorides as shown in the table below.

Group	**1**	**2**	**3**	**4**	**5**	**6**	**7**
compound	$NaCl$	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_5	SCl_2	Cl_2
melting point (°C)	801	714	190	–70	–162	–121	–101

7.2 Formation of ionic compounds



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that metal atoms lose electrons to form positively charged ions (cations) and non-metal atoms gain electrons to form negatively charged ions (anions)
- determine the charge on an ion using its electron configuration and the octet rule
- understand that when an ionic compound is formed from cations and anions, the ions combine in proportions that produce a compound with an overall zero charge.

Some of the reactions that occur between metals and non-metals to form ionic compounds are very vigorous. The reaction between sodium and chlorine to form sodium chloride produces a lot of heat. You will remember from Chapter 4 that sodium is very reactive. At high temperatures, the production of sodium chloride from sodium metal and chlorine gas is very explosive, producing a flame and large amounts of energy.

FORMING IONS

When metal atoms react with non-metal atoms to form an ionic compound the following occur.

- Metal atoms lose electrons to form positively charged **ions** (cations).
- Non-metal atoms gain electrons to form negatively charged ions (anions).

From Chapter 4, you will remember that most metals have low ionisation energies and low **electronegativities**. This means that non-metal atoms are usually more electronegative than metal atoms. In other words, non-metals have a stronger attraction for electrons than metals. In reactions that form ionic compounds, metal atoms transfer one or more valence electrons to non-metal atoms.

Most importantly, the ions that are formed usually have eight electrons in their outer shell—a stable electron configuration. The tendency for **elements** to react in such a way that their atoms have eight electrons in their outer shell (**valence shell**) is known as the **octet rule**. Noble gases (group 18) are elements that already have the most stable valence shell configuration. Therefore, another way of thinking about the octet rule is that atoms tend to gain or lose electrons to gain a stable electron configuration identical to that of the noble gas nearest to them on the periodic table. The formation of stable ions is a powerful driving force in reactions between metals and non-metals when they transfer electrons and produce ionic compounds.

When an atom loses electrons, it becomes a cation (positively charged) as the number of protons no longer equals the number of electrons—there is a charge imbalance. The ion formed is written with a superscript + sign indicating the charge. When there are two electrons lost, its charge is written as 2+ in superscript.

$$X \rightarrow X^+ + e^-$$

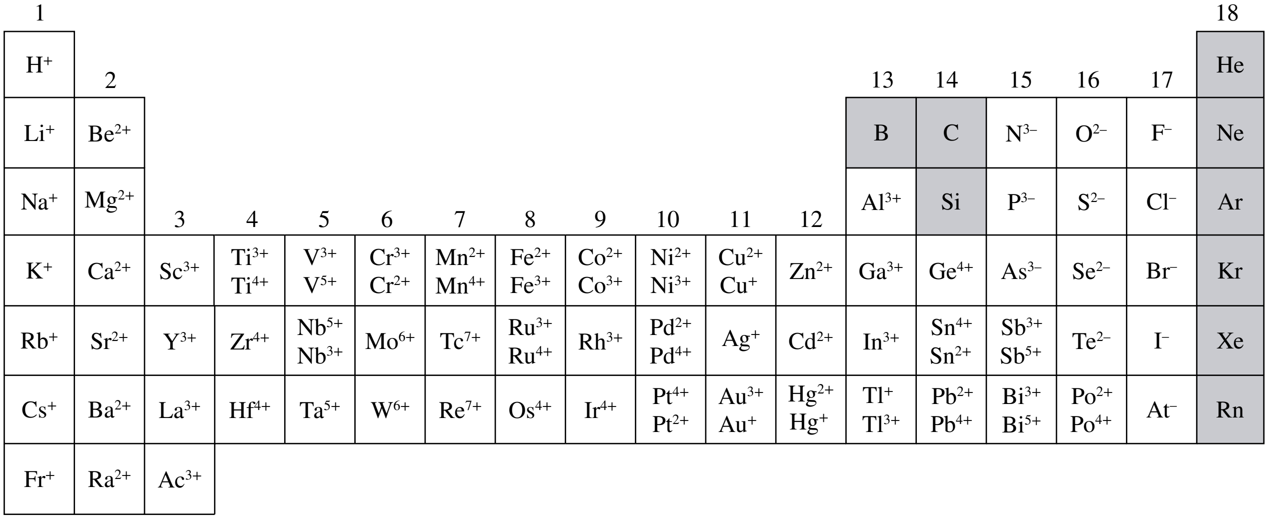
When an atom gains electrons, as in the case of a non-metal, it becomes an anion (negatively charged). This is written with a superscript –. If three electrons are gained, the charge is written as 3– in superscript.

$$Y + e^- \rightarrow Y^-$$

In Chapter 4, you learnt that atoms of elements in group 1 of the periodic table have electron configurations with one electron in the valence shell. Therefore, elements in this group often form cations with a charge of +1, as they readily lose this one outershell electron. For example, potassium has an electron configuration of 2,8,8,1. It will readily lose the one electron in the valence shell to have an identical electron configuration to argon (the nearest noble gas). The potassium ion now has an electron configuration of 2,8,8 and a charge of +1. Note that, for simplicity, you will look at the arrangement of electrons just in shells of atoms (rather than in subshells).

Metals in group 2 of the periodic table have electron configurations with two electrons in their valence shells. Therefore, they readily form ions with a charge of +2 as they lose these two valence electrons.

Non-metals in group 16 have six electrons in their valence shells. Therefore, they gain two electrons to form anions with a charge of −2. Group 17 non-metals have seven electrons in their valence shell. They readily gain one electron to fill the valence shell and form anions with a charge of −1. All ions have a full valence shell according to the octet rule. This means they form anions with a charge of −1. Therefore, the periodic table can be used as a quick reference point to begin forming ionic compounds and writing their formulas. Figure 7.2.1 shows the charge on the most common ions formed relative to the group of the element.



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H^+																	He
Li^+	Be^{2+}											B	C	N^{3-}	O^{2-}	F^-	Ne
Na^+	Mg^{2+}											Al^{3+}	Si	P^{3-}	S^{2-}	Cl^-	Ar
K^+	Ca^{2+}	Sc^{3+}	Ti^{3+} Ti^{4+}	V^{3+} V^{5+}	Cr^{3+} Cr^{2+}	Mn^{2+} Mn^{4+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^{2+} Cu^+	Zn^{2+}	Ga^{3+}	Ge^{4+}	As^{3-}	Se^{2-}	Br^-	Kr
Rb^+	Sr^{2+}	Y^{3+}	Zr^{4+}	Nb^{5+} Nb^{3+}	Mo^{6+}	Tc^{7+}	Ru^{3+} Ru^{4+}	Rh^{3+}	Pd^{2+} Pd^{4+}	Ag^+	Cd^{2+}	In^{3+}	Sn^{4+} Sn^{2+}	Sb^{3+} Sb^{5+}	Te^{2-}	I^-	Xe
Cs^+	Ba^{2+}	La^{3+}	Hf^{4+}	Ta^{5+}	W^{6+}	Re^{7+}	Os^{4+}	Ir^{4+}	Pt^{4+} Pt^{2+}	Au^{3+} Au^+	Hg^{2+} Hg^+	Tl^+ Tl^{3+}	Pb^{2+} Pb^{4+}	Bi^{3+} Bi^{5+}	Po^{2+} Po^{4+}	At^-	Rn
Fr^+	Ra^{2+}	Ac^{3+}															

FIGURE 7.2.1 Periodic table showing the common ions formed by each element. The dominant form is shown on the top.

Transition metal ions

A **transition metal** forms one or more stable ions which have incompletely filled *d* orbitals. For example, copper can form Cu^+ ions with a charge of +1, and Cu^{2+} ions with a charge of +2. This is because the *s*-, *p*- and *d*-subshells of these transition metals are of similar energy levels. Two or more electron configurations of the metal ion have similar stability, although one is usually more common than the other.

Some other metals with ions of variable charges are shown in Figure 7.2.1 and below:

- lead (Pb^{2+} and Pb^{4+})
- iron (Fe^{2+} and Fe^{3+})
- tin (Sn^{2+} and Sn^{4+}).

For compounds of transition metal ions, you need to specify the charge on the metal ion by writing a Roman numeral in brackets immediately after the metal in the name of the compound. For example:

- iron(II) chloride contains the Fe^{2+} ion
- iron(III) chloride contains the Fe^{3+} ion
- copper(I) sulfide contains the Cu^+ ion.

FORMING IONIC COMPOUNDS

When sodium (metal atom) reacts with chlorine (non-metal atom) an **ionic compound** is formed in two stages. Firstly, as seen in the previous module, each sodium atom loses one electron and each chlorine atom gains one electron.

After the electron is transferred:

- the sodium ion (Na^+) has the stable electron shell configuration of 2,8 (the same as a neon atom)
- the chloride ion (Cl^-) has the stable electron shell configuration of 2,8,8 (the same as an argon atom).

Figure 7.2.2 illustrates how, when sodium reacts with chlorine, an electron is lost by a sodium atom and gained by a chlorine atom. A diagram of this type is called an **electron transfer diagram**.

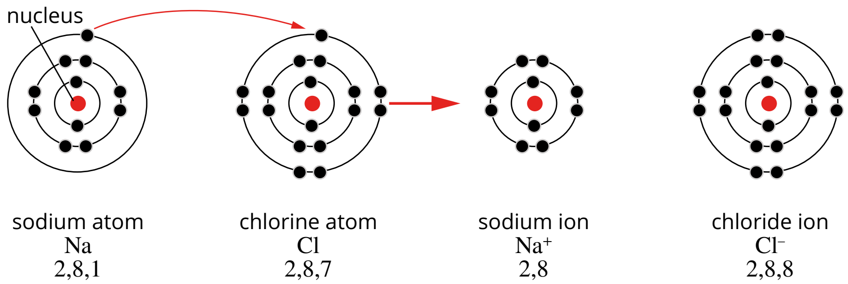


FIGURE 7.2.2 Electron transfer diagram showing the formation of sodium and chloride ions

In the second stage the strong electrostatic attraction between Na^+ and Cl^- draws the two ions together and forms an ionic bond as shown in the equation below. A neutral ionic compound containing charged particles is formed.

$$Na^+ + Cl^- \rightarrow NaCl$$

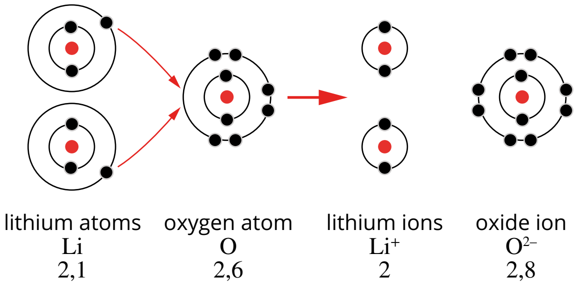
Ionic compound formation

Stage one $X + Y \rightarrow X^+ + Y^-$

Stage two $X^+ + Y^- \rightarrow XY$

Let us look at the formation of two other ionic compounds, lithium oxide and magnesium chloride. The reaction between lithium and oxygen atoms is illustrated in Figure 7.2.3. In this reaction, an oxygen atom needs to gain two electrons to have eight electrons in its outer shell and form a stable ion. Lithium has a single electron in its valence shell and loses one electron. To allow this to happen, one oxygen atom will react with two lithium atoms, taking one electron from each atom.

After the reaction, there are just two electrons in what is now the outer shell of the Li^+ ion, which is the same as the electron configuration of a helium atom (2). The oxygen ion has the same electron configuration as neon (2,8).



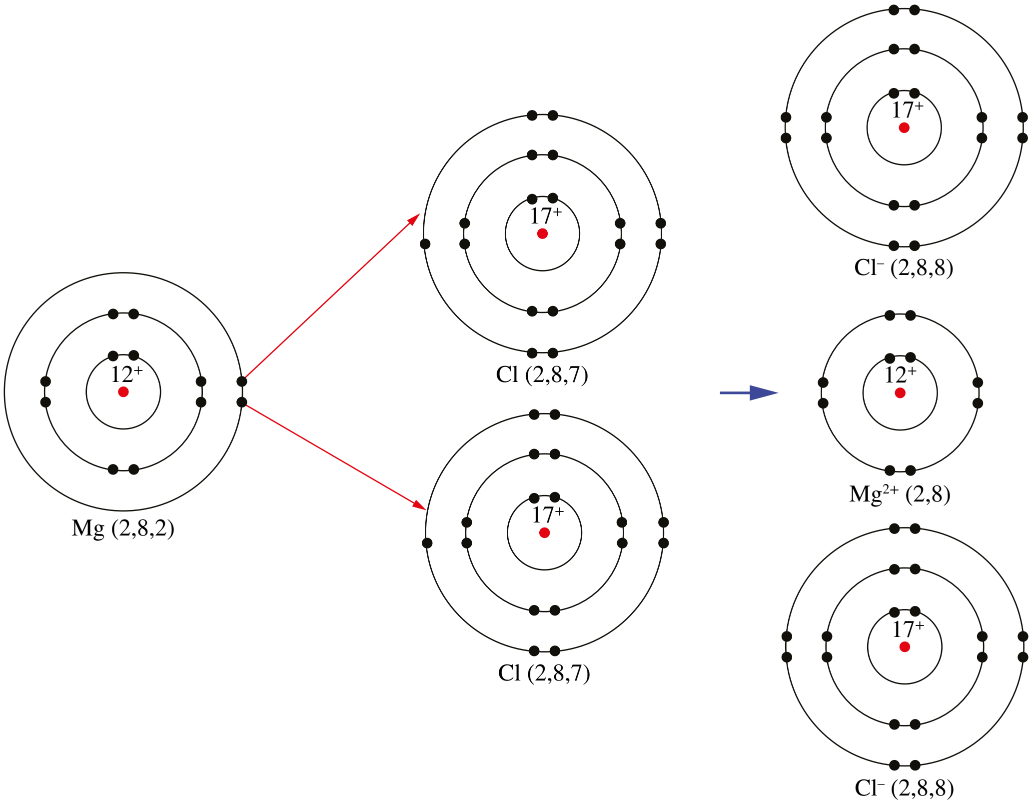
$$2Li^+ + O^{2-} \rightarrow Li_2O$$

FIGURE 7.2.3 An electron transfer diagram showing the formation of lithium and oxide ions

The strong electrostatic attraction between the two Li^+ ions and one O^{2-} ion forms the ionic bond as shown in Figure 7.2.3. The lithium symbol has a 2 as a subscript to show that two lithium ions are present in the compound. (This is explained further in Module 7.3.)

Figure 7.2.4 shows the electron transfer diagram and chemical equation for the reaction of magnesium with chlorine. A magnesium atom has an electron configuration of 2,8,2 so it will readily lose its two valence electrons. A magnesium ion (Mg^{2+}) is formed so that it now has the electron configuration of the nearest noble gas, neon (2,8).

Each chlorine atom has an electron configuration of 2,8,7, so each will gain one outer-shell electron. A chloride ion (Cl^-) has the electron configuration of the noble gas argon (2,8,8). Because magnesium readily donates two electrons, it is found with *two* chlorine atoms, each chlorine atom taking *one* electron each from magnesium. The 2 subscript next to chlorine in the chemical formula shows the number of chloride ions present (explained further in Module 7.3).



$$Mg + 2Cl \rightarrow Mg^{2+} + 2Cl^-$$
$$Mg^{2+} + 2Cl^- \rightarrow MgCl_2$$

FIGURE 7.2.4 Electron transfer diagram and chemical equations showing the formation of magnesium chloride from one magnesium atom and two Cl atoms

You will learn to write simple and more complex chemical formulas for ionic compounds in Modules 7.3 and 7.4.

Worked example 7.2.1

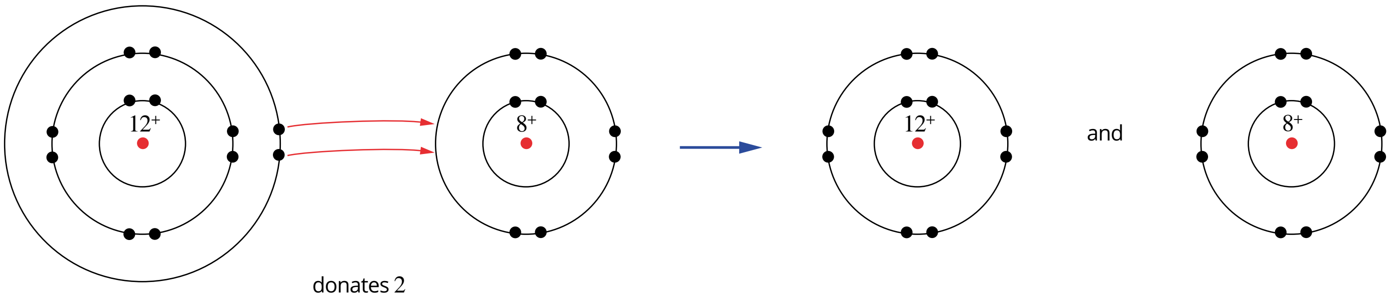
WRITING EQUATIONS FOR FORMATION OF AN IONIC COMPOUND BY REACTION BETWEEN A METAL AND NON-METAL ATOM

Develop an equation for the reaction between magnesium and oxygen atoms to form magnesium oxide. Show the electron configurations for each element before and after the reaction and the chemical equation for the combination of the ions. Include an electron transfer diagram for the reaction.

Thinking	**Working**
Write the symbol and electron configuration of the two atoms forming the ionic compound.	Mg (2,8,2) and O (2,6)
Write the symbol, electron configuration and charge of the two ions forming the ionic compound.	Mg^{2+} (2,8) and O^{2-} (2,8)
Calculate the lowest common multiple of the two numbers in the charges of the ions.	Number of charges is the same for both ions.

Calculate how many positive ions are needed to equal the negative charges.	one Mg^{2+} ion
Calculate how many negative ions are needed to equal the lowest common multiple.	one O^{2-} ion
Use the answers from the previous two steps to write the formula for the ionic compound using subscripts to show number of ions. Write the symbol of the positive ion first. (Note that 1 is not written as a subscript.)	MgO
Write the chemical equation for the combination of magnesium and oxygen.	Mg (2,8,2) + O (2,6) $\rightarrow$ Mg^{2+} (2,8) + O^{2-} (2,8) $Mg^{2+} + O^{2-} \rightarrow MgO$

Draw the electron transfer diagram using arrows to show transfer of electrons and electron configuration to show change to a full valence shell for both ions.



Mg (2,8,2)	donates 2 electrons to	O (2,6)	to form	Mg^{2+} (2,8)	and	O^{2-} (2,8)
Mg (2,8,2)	+	O (2,6)	→	Mg^{2+} (2,8)	+	O^{2-} (2,8)
Mg^{2+}	+	O^{2-}	→	MgO		

➤ Try yourself 7.2.1

WRITING EQUATIONS FOR FORMATION OF AN IONIC COMPOUND BY REACTION BETWEEN A METAL AND NON-METAL ATOM

Develop an equation for the reaction between sodium and phosphorus atoms to form sodium phosphide. Show the electron configurations for each element before and after the reaction and the chemical equation for the combination of the ions. Include an electron transfer diagram for the reaction. (Hint: these atoms do not combine in a 1:1 ratio. Refer to the charge on each ion.)

7.2 Review

SUMMARY

- During the formation of ionic compounds an imbalance in the numbers of electrons and protons gives atoms a negative or positive charge.
- Metal atoms lose electrons to form positively charged ions (cations).
- Non-metal atoms gain electrons to form negatively charged ions (anions).
- The ions present in an ionic compound have a stable electron configuration identical to that of the noble gas nearest to them on the periodic table (the octet rule).
- For transition metals (and several other metals) that form ions with different charges, the charge on the ion is shown by placing a Roman numeral after the name of the metal.
- When an ionic compound is formed from cations and anions, the ions combine in proportions that produce a compound with an overall zero charge (refer to Module 7.3).

KEY QUESTIONS

Retrieval

1 Identify the number of electrons in each of the following ions.

- **a** K^+
- **b** Cl^-
- **c** Mg^{2+}
- **d** H^+
- **e** Al^{3+}
- **f** O^{2-}

Comprehension

2 Explain whether the following atoms will form cations or anions and identify the ion formed.

- **a** calcium
- **b** fluorine
- **c** aluminium
- **d** phosphorus

3 Use electron transfer diagrams to show the formation of ions when electrons are transferred between magnesium and fluorine.

4 Use the octet rule and refer to noble gas configuration to explain why group 2 metals of the periodic table are likely to form cations with a charge of +2.

5 Determine the number of electrons lost or gained by copper when forming cations of +1 and +2 charges. Explain why copper forms two cations.

6 Determine the anions and cations present in the following.

- **a** AgBr solution
- **b** molten FeI_2
- **c** Na_3N solution

7 Represent the reaction between the following metal and non-metal atoms with an equation. Show the final electron configurations for each ion after the reaction.

- **a** sodium and chlorine atoms
- **b** magnesium and oxygen atoms
- **c** aluminium and sulfur atoms

Analysis

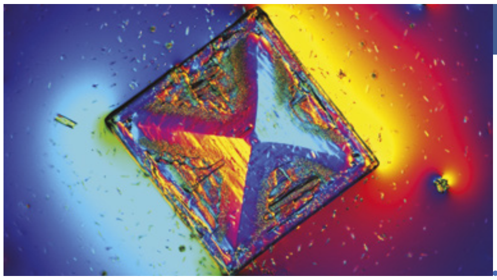
8 Show, using chemical equations:

- **i** the formation of ions
- **ii** ionic compounds

for the reaction for the atom pairs below.

- **a** calcium and fluorine
- **b** sodium and oxygen
- **c** aluminium and phosphorus

7.3 Chemical formulas of simple ionic compounds



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- write formulas for ionic substances by determining the number of each ion in the formula unit based on the charge on the ion and the overall neutrality of the compound
- use rules to write formulas using symbols and subscripts
- use the chemical formula to determine the name of the compound.

You have seen that ionic compounds contain oppositely charged ions that are arranged in three-dimensional lattices. The charges on these ions arise from the transfer of electrons between atoms. An ionic crystal may contain billions of these ions in an immense three-dimensional lattice, so a formula is used to describe the simplest ratio of ions in a neutral unit of the ionic compound.

WRITING THE FORMULA OF AN IONIC COMPOUND

Ionic compounds are electrically neutral because they are composed of ions formed by the initial transfer of electrons followed by the combining of the ions into a compound. This means that the total number of positive charges on the metal ions must equal the total number of negative charges on the non-metal ions. This principle will guide you when trying to work out the formula of an ionic compound.

An example of this is the formula of the ionic compound magnesium chloride.

- A magnesium ion (Mg^{2+}) has a +2 charge.
- A chloride ion (Cl^{-}) has a −1 charge.

Therefore, in a crystal of magnesium chloride, two chloride ions are needed to provide two negative charges so that they balance the +2 charge on every magnesium ion. Therefore, the ratio of magnesium ions to chloride ions in the crystal is 1:2 and the formula of magnesium chloride is $MgCl_2$. The electron transfer diagram of this reaction is shown in Figure 7.2.4 on page 175.

Figure 7.3.1 illustrates how formulas for some other ionic compounds can be determined. Tables 7.3.1 and 7.3.2 list some of the more common positively and negatively charged ions. You may use these when you are writing formulas for ionic compounds. Remember that an ionic substance contains millions of ions, depending on the size of the crystal. Therefore, the ionic formula is known as a formula unit indicating the simplest whole number ratio of ions making up the lattice.

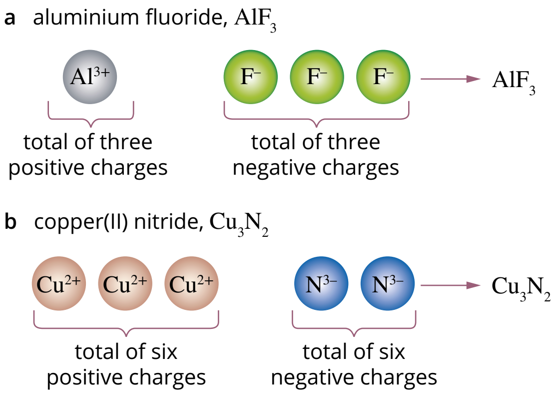


FIGURE 7.3.1 How to deduce chemical formulas from the charges on ions

In Module 7.2, you learnt that transition metals are able to form more than one ion. As seen in Table 7.3.1, the charge of the ion of a transition metal is taken from the Roman numeral that is in brackets after the name of the ion.

TABLE 7.3.1 Names and formulas of some common cations

Charge			
+1	**+2**	**+3**	**+4**
caesium, Cs^+	barium, Ba^{2+}	aluminium, Al^{3+}	lead(IV), Pb^{4+}
copper(I), Cu^+	cadmium(II), Cd^{2+}	chromium(III), Cr^{3+}	tin(IV), Sn^{4+}
gold(I), Au^+	calcium, Ca^{2+}	gold(III), Au^{3+}	
lithium, Li^+	cobalt(II), Co^{2+}	iron(III), Fe^{3+}	
potassium, K^+	copper(II), Cu^{2+}		
rubidium, Rb^+	iron(II), Fe^{2+}		
silver, Ag^+	lead(II), Pb^{2+}		
sodium, Na^+	magnesium, Mg^{2+}		
	manganese(II), Mn^{2+}		
	mercury(II), Hg^{2+}		
	nickel, Ni^{2+}		
	strontium, Sr^{2+}		
	tin(II), Sn^{2+}		
	zinc, Zn^{2+}		

TABLE 7.3.2 Names and formulas of some common monatomic anions

Charge		
–1	**–2**	**–3**
bromide, Br^-	oxide, O^{2-}	nitride, N^{3-}
chloride, Cl^-	sulfide, S^{2-}	
fluoride, F^-		
iodide, I^-		

Rules for writing chemical formulas

Here are some simple rules to follow when you are writing chemical formulas.

- Write the symbol for the positively charged ion first.
- Use subscripts to indicate the number of each ion in the formula. Remember to add all negative charges and positive charges to check the compound is neutral overall. Use the lowest common multiple of the two numbers in the charges of the ions. Write the subscripts after the ion they refer to.
- If there is just one ion present in the formula, omit the subscript '1'.
- Do not include the charges on the ions in the balanced formula.

These rules are illustrated in Figure 7.3.2.

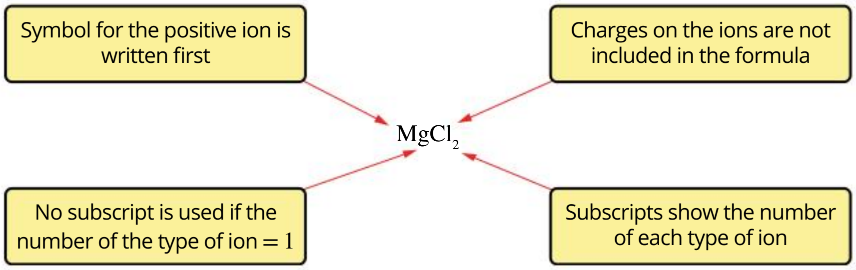


FIGURE 7.3.2 Summary of the information provided by a chemical formula

Worked example 7.3.1

STEPS IN WRITING A CHEMICAL FORMULA

Determine the chemical formula of the ionic compound formed between zinc and nitride ions. You may need to refer to Tables 7.3.1 and 7.3.2.

Thinking	**Working**
Write the symbol and charge of the two ions forming the ionic compound.	Zn^{2+} and N^{3-}
Calculate the lowest common multiple of the two numbers in the charges of the ions.	$2 \times 3 = 6$
Calculate how many positive ions are needed to equal the lowest common multiple.	three Zn^{2+} ions
Calculate how many negative ions are needed to equal the lowest common multiple.	two N^{3-} ions
Use the answers from the previous two steps to write the formula for the ionic compound using subscripts to show number of ions. Write the symbol of the positive ion first. (Note that 1 is not written as a subscript.)	Zn_3N_2

➤ Try yourself 7.3.1

STEPS IN WRITING A CHEMICAL FORMULA

Determine the chemical formula of the ionic compound formed between barium and fluoride ions. You may need to refer to Tables 7.3.1 and 7.3.2.

NAMING SIMPLE IONIC COMPOUNDS

There are some basic conventions that are followed when naming ionic compounds.

- The first part of the name is the name of a positively charged metal ion (cation). This is the same as the name of the metal. For example, the cation of a sodium atom is called a sodium ion; the cation of an aluminium atom is an aluminium ion.
- For simple monatomic anions, the second part of the name is the non-metal ion (anion). The name of this ion is similar to that of the atom, but ends in '-ide'. For example, the anion of the chlorine atom is chloride; the anion of the oxygen atom is oxide. (Some names for more complex ions are derived differently and are described in Module 7.4.)

Worked example 7.3.2

STEPS IN WRITING THE NAME OF AN IONIC COMPOUND

Determine the name of Zn_3N_2 (without using Tables 7.3.1 and 7.3.2).

Thinking	**Working**
Write the name of the first element (metal) in the formula.	zinc
Write the name of the second element (non-metal) in the formula.	nitrogen
Change the ending of the non-metal element to 'ide'.	nitride
Write the two names together (with a space between each word).	zinc(II) nitride

➤ Try yourself 7.3.2

STEPS IN WRITING THE NAME OF AN IONIC COMPOUND

WS 1.1.4

Determine the name of $CaBr_2$ (without using Tables 7.3.1 and 7.3.2).

7.3 Review

SUMMARY

- When determining the formula of an ionic compound, the total number of positive charges on the metal ion(s) must equal the total number of negative charges on the non-metal ion(s).
- When writing formulas of ionic compounds:
  - the symbol for the positively charged ion is written first
  - subscripts are used to indicate the number of each ion in the formula
  - the charges on the ions are not included in the balanced formula.
- When naming a simple ionic compound, the name of the metal element is written first followed by the name of the non-metal element with the ending changed to 'ide'.

KEY QUESTIONS

Retrieval

1 Use the information in Tables 7.3.1 and 7.3.2 on page 179 to name the ionic compounds with the following formulas.

a LiCl
b CaO
c MgS
d K_2O
e NaF

Comprehension

2 Listed below are pairs of metal and non-metal ions. Determine the ratio of the metal ion to the non-metal ion when they combine to form ionic compounds.

a K^+ and S^{2-}
b Al^{3+} and F^-
c Ca^{2+} and N^{3-}
d Al^{3+} and P^{3-}
e Mg^{2+} and Cl^-

3 Use the information in Tables 7.3.1 and 7.3.2 on page 179 to write formulas for the following ionic compounds.

a sodium chloride
b potassium bromide
c zinc chloride
d potassium oxide
e barium bromide
f aluminium iodide
g silver bromide
h zinc oxide
i barium oxide
j aluminium sulfide

4 Use the rules on page 179 and without referring to Tables 7.3.1 and 7.3.2 on page 179, write the names of the following ionic compounds.

a KCl
b CaO
c FeO
d AgI
e Al_2S_3
f $HgBr_2$
g Ca_3N_2
h CuF
i $PbCl_4$

Analysis

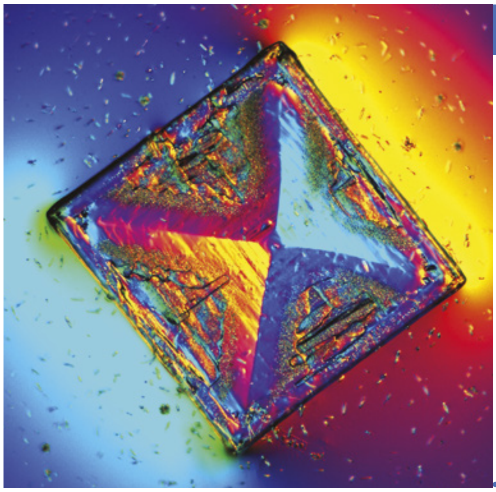
5 Using the periodic table on page 173 and without referring to Tables 7.3.1 and 7.3.2 on page 179, determine the formulas for the ionic compounds made from the following metals and non-metals.

a potassium and chlorine
b calcium and iodine
c bromine and aluminium
d oxygen and magnesium
e rubidium and phosphorus
f selenium and beryllium
g nitrogen and lithium

6 A_2B is an ionic compound. Both of the ions in A_2B have the same electron configuration as an argon atom. Identify A_2B from the options below. Explain your answer.

A potassium sulfide
B calcium chloride
C calcium sulfide
D sodium oxide

7.4 Writing formulas of more complex ionic compounds



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that ions can be composed of more than one atom. These are called polyatomic ions
- understand that a polyatomic ion has an overall charge like other anions and cations and so (as before) will be combined with an ion of the opposing charge in a fixed ratio to form a neutral compound
- use brackets to surround the polyatomic ion in a chemical formula when more than one polyatomic ion is required to make a neutral ionic compound. A subscript after the bracket indicates the number of polyatomic ions required in the formula
- use a naming convention for the polyatomic ion that is similar to that used for simple covalent substances.

The chemical formulas you have written for ionic compounds so far contain simple ions—ions that contain only one atom of an element. In this module, you will look at the formulas of some compounds containing more complex polyatomic ions. You will also learn rules to help you name the ionic compounds of these polyatomic ions.

FORMULAS CONTAINING POLYATOMIC IONS

Simple monatomic ions contain only one atom, as you have seen in Module 7.3. However, other ions contain atoms of two or more elements. These ions are called **polyatomic ions**.

In polyatomic ions:

- different atoms are present combined in a fixed ratio
- the group of atoms behaves as a single unit with a specific charge
- subscripts are used to indicate the number of each kind of atom in the ion.

For example, a carbonate ion (CO_3^{2-}) contains one carbon atom and three oxygen atoms covalently bonded to form an ion. The carbonate ion has a charge of –2. Other common polyatomic ions are nitrate (NO_3^-), hydroxide (OH^-) and phosphate (PO_4^{3-}). The formulas of common polyatomic ions can be seen in Table 7.4.1.

TABLE 7.4.1 Common polyatomic cations and anions

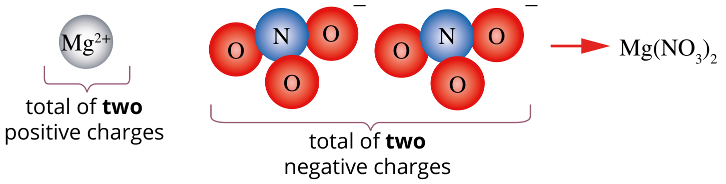
Charge			
+1	–1	–2	–3
ammonium, NH_4^+	• cyanide, CN^- • dihydrogen phosphate, $H_2PO_4^-$ • ethanoate, CH_3COO^- • hydrogen carbonate, HCO_3^- • hydrogen sulfide, HS^- • hydrogen sulfite, HSO_3^- • hydrogen sulfate, HSO_4^- • hydroxide, OH^- • nitrite, NO_2^- • nitrate, NO_3^- • permanganate, MnO_4^-	• carbonate, CO_3^{2-} • chromate, CrO_4^{2-} • dichromate, $Cr_2O_7^{2-}$ • hydrogen phosphate, HPO_4^{2-} • oxalate, $C_2O_4^{2-}$ • sulfite, SO_3^{2-} • sulfate, SO_4^{2-}	• phosphate, PO_4^{3-}

If more than one polyatomic ion is required in a formula to balance the charge, then it is placed in brackets with the required number written as a subscript after the brackets. Some examples are:

- magnesium nitrate, $Mg(NO_3)_2$. The brackets and subscript show that for every Mg^{2+} ion there are two nitrate (NO_3^-) ions present in the lattice.
- iron(III) sulfate, $Fe_2(SO_4)_3$. The (III) in the formula means the cation has a +3 charge. The brackets and subscript show that for every three sulfate ions (SO_4^{2-}) there are two iron(III) ions present in the lattice.

The formation of these two compounds containing polyatomic ions is diagrammatically shown in Figure 7.4.1.

a magnesium nitrate, $Mg(NO_3)_2$



b iron(III) sulfate, $Fe_2(SO_4)_3$

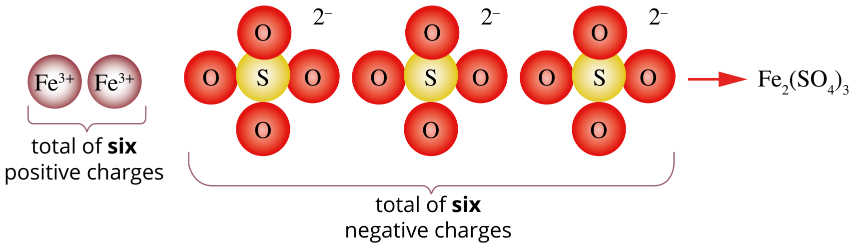


FIGURE 7.4.1 The chemical formulas of ionic compounds containing polyatomic ions

Polyatomic ions are made up of two or more different atoms that have an overall charge. They need to be written within brackets if there is more than one present in an ionic compound. Subscripts are used to indicate the ratio of ions in the crystal lattice.

Note that brackets are not required for the formula of sodium nitrate, $NaNO_3$, where there is only one NO_3^- ion present for each sodium ion.

Worked example 7.4.1

STEPS IN WRITING A CHEMICAL FORMULA INCLUDING POLYATOMIC IONS

<table>
<tr><td colspan="2">Determine the chemical formula of the ionic compound formed between potassium and chromate ions. You may need to refer to Tables 7.3.1, 7.3.2 and 7.4.1.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Write the symbol and charge of the two ions forming the ionic compound.</td><td>K^+ and CrO_4^{2-}</td></tr>
<tr><td>Calculate the lowest common multiple of the two numbers in the charges of the ions.</td><td>$1 \times 2 = 2$</td></tr>
<tr><td>Calculate how many positive ions are needed to equal the lowest common multiple.</td><td>two K^+ ions</td></tr>
<tr><td>Calculate how many negative ions are needed to equal the lowest common multiple.</td><td>one CrO_4^{2-} ion</td></tr>
<tr><td>Use the answers from the previous two steps to write the formula for the ionic compound using subscripts to show number of ions. Write the symbol of the positive ion first.
(Note that 1 is not written as a subscript.)</td><td>K_2CrO_4</td></tr>
</table>

➤ Try yourself 7.4.1

STEPS IN WRITING A CHEMICAL FORMULA INCLUDING POLYATOMIC IONS

Determine the chemical formula of the ionic compound formed between barium and hydroxide ions. You may need to refer to Tables 7.3.1, 7.3.2 and 7.4.1.

Naming ionic compounds containing polyatomic ions

In Module 7.3, you learnt that when naming simple ionic compounds the name of the metal element is the first part of the name. This rule also holds for ionic compounds containing polyatomic ions. However, the name of the polyatomic anion usually does not end in 'ide'. Polyatomic anions containing oxygen will end in '-ite' or '-ate'. The name of the ion with the smaller number of oxygen atoms usually ends in '-ite' and the one with the larger number of oxygen atoms ends in '-ate'. For example, the NO_2^- ion is called a nitrite ion and the NO_3^- ion is called a nitrate ion.

Some ions that are widely used have a common name. The common name for the ethanoate ion is acetate, and the common name for the hydrogen carbonate ion is bicarbonate.

Worked example 7.4.2

WRITING THE NAME OF AN IONIC COMPOUND CONTAINING A POLYATOMIC ION

Determine the name of Na_3PO_4. You may need to refer to Table 7.4.1.

Thinking	**Working**
Write the name of the first ion in the formula.	sodium
Write the name of the second ion in the formula.	phosphate
Write the two names together (with a space between).	sodium phosphate

➤ Try yourself 7.4.2

WRITING THE NAME OF AN IONIC COMPOUND CONTAINING A POLYATOMIC ION

State the name of $CaCO_3$. You may need to refer to Table 7.4.1.

7.4 Review

SUMMARY

- Ions that contain two or more atoms of different elements are called polyatomic ions.
- When writing formulas for polyatomic ionic compounds, these rules apply.
  - The symbol for the positively charged ion is written first.
  - Subscripts are used to indicate the number of each ion in the formula.
  - If a chemical formula contains more than one polyatomic ion, the formula of the ion is placed in brackets with the number of ions written as a subscript after the brackets.
  - Charges on ions are not shown in the chemical formula.
- When naming polyatomic ionic compounds, the following rules apply.
  - The name of the metal ion is the same as the name of the metal.
  - Polyatomic anions containing oxygen usually end in '-ite' or '-ate'.

KEY QUESTIONS

Retrieval

1 The section ‘Naming ionic compounds containing polyatomic ions’ describes two examples of ions with common names. Use Table 7.4.1 on page 182 to identify two further examples.

2 Identify a link between the ‘left and right’ of the periodic table and the position of an ion in the chemical formula.

3 Determine the formulas of the following polyatomic ions. Use the information in Table 7.4.1 on page 182.

- **a** hydrogen sulfite
- **b** phosphate
- **c** carbonate
- **d** cyanide

Comprehension

4 Use the information in Table 7.3.1 on page 179 and Table 7.4.1 on page 182 to name the following ionic compounds.

- **a** $Mg(OH)_2$
- **b** Na_2CO_3
- **c** $FePO_4$
- **d** $CuSO_4$
- **e** $Ba(NO_3)_2$
- **f** Cu_2SO_4
- **g** $Fe(CH_3COO)_3$
- **h** NH_4NO_3
- **i** Na_2HPO_4

5 Use the information in Table 7.3.1 on page 179 and Table 7.4.1 on page 182 to determine the chemical formulas for the following ionic compounds.

- **a** sodium carbonate
- **b** barium nitrate
- **c** aluminium nitrate
- **d** calcium hydroxide
- **e** zinc(II) sulfate
- **f** potassium hydroxide
- **g** potassium nitrate
- **h** zinc(II) carbonate
- **i** potassium sulfate
- **j** barium hydroxide

6 Use the information in Tables 7.3.1 and 7.3.2 on page 179 and Table 7.4.1 on page 182 to determine the chemical formulas for the following ionic compounds. Identify compounds where there are the same number of cations as anions in the formula unit of the compound.

- **a** copper(I) chloride
- **b** iron(III) oxide
- **c** magnesium dichromate
- **d** chromium(III) sulfate
- **e** iron(II) sulfite
- **f** lead(II) nitrate
- **g** lead(IV) oxide
- **h** tin(II) hydroxide

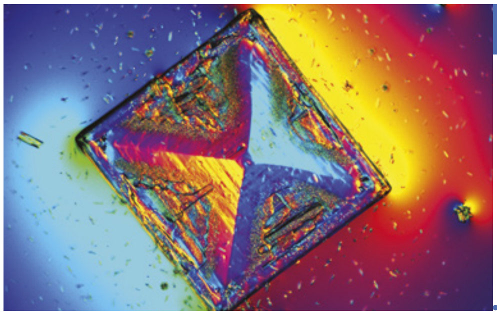
Analysis

7 Consider the formulas for ammonium hydrogen phosphate and aluminium carbonate. Summarise in the table below, for each compound, the number of atoms of each element found in the formula unit of each of the compounds. The first one has been started for you.

Compound: ammonium hydrogen phosphate	
Formula:	
Element	**Number of atoms in formula unit**
N	2
H	
P	1
O	

Compound: aluminium carbonate	
Formula:	
Element	**Number of atoms in formula unit**

7.5 Properties of ionic substances



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand key properties of ionic substances—high melting points, hardness and brittleness, electrical conductivity and solubility
- relate the physical properties of ionic substances to the structure of the ionic lattice
- understand ways that the properties of ionic substances can be utilised for our health and wellbeing.

You have seen that ionic compounds are formed from positive and negative ions combined in various ratios. A wide array of compounds is possible by the combination of cations and anions into a strong lattice structure. The abundance of ionic compounds on Earth and their properties such as hardness make them essential materials for everyday use. Most rocks, minerals and gemstones (Figure 7.5.1) are ionic compounds. Weathered rocks form clays, which can be made into ceramics, kitchen crockery and bricks.

When comparing physical properties of ionic substances to other bonding types, you can see that ionic compounds show high melting points common with metallic substances. Also, ionic substances have low electrical **conductivity** in common with covalent substances. These and other properties can be related to the structure of the ionic compound.

FIGURE 7.5.1 Many gemstones are made from ionic compounds.

HIGH MELTING POINTS

To melt an ionic solid such as sodium chloride, you must provide energy to allow the ions to break free and move. Sodium chloride has a high melting point (801°C). This indicates that a large amount of energy is needed to overcome the electrostatic attraction between oppositely charged ions and allow them to move freely. Therefore, the bonds in the lattice must be strong.

In Modules 7.3 and 7.4, you learnt that ions can have various charges and can be made from a single element or more than one element (polyatomic ion). These differences will affect the alignment of atoms (shape of the crystals) and the strength of the electrostatic attractions in the crystal structure. A lower melting point indicates weaker ionic bonding.

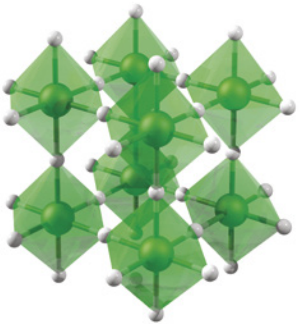


FIGURE 7.5.2 Rhenium oxide (ReO_3) has a lower melting point than most ionic compounds (400°C). It has a more complex lattice structure than NaCl.

Ions that are formed from the transfer of more than one electron (for example Fe^{2+}) show strong electrostatic attraction and generally high melting points. Also, a large difference in the size of the anion and cation results in incomplete shielding of particles with the same charge. This allows them to repel each other and weaken the lattice. Therefore, ionic compounds composed of polyatomic ions or very large ions combined with a small ion often have lower melting points than compounds where the anion and cation are single atoms. An example of this is the compound (ReO_3) shown in Figure 7.5.2.

The high melting point of ionic compounds is put to use in the bricks that line furnaces and kilns (Figure 7.5.3) and in the **ceramic** materials used to make brake discs for high-performance cars (Figure 7.5.4).

FIGURE 7.5.3 Bricks made from the ionic compound magnesium oxide are used to line furnaces and kilns.

FIGURE 7.5.4 Ceramic brake discs work more effectively than steel ones at high temperatures. A ceramic brake disc contains ionic compounds that have very high melting temperatures and withstand the heat produced by braking better than metals.

HARDNESS AND BRITTLENESS

There are strong electrostatic forces of attraction between ions in an ionic compound, so a strong force is needed to disrupt the **crystal lattice**. Therefore, one of the properties of ionic compounds is that they are hard. This means that a sodium chloride crystal cannot be scratched easily.

The strength of house bricks, concrete bridges and cobbled streets can be attributed to the ionic bonding within their structures.

Some of the earliest tools used by humans were axes, spearheads and coarse needles used for weaving. Each application required a material that was hard and could be shaped. Certain types of rocks that are composed of ionic compounds served this purpose well. Figure 7.5.5 shows a primitive axehead used during the Stone Age. The Stone Age ended at different times in different parts of the world, as humans learnt to smelt (fuse or melt) metals such as copper from their ores to create more refined and lighter tools. Their metallic nature meant that they were not as hard as the original tools made of rock composed of ionic compounds.

FIGURE 7.5.5 The hardness of ionic compounds is the reason why axes were once made from rocks.

The characteristic hardness of ionic compounds allows them to have other key roles.

- Calcium phosphate is a constituent of bone tissue that gives it strength.
- Calcium sulfate in the form of gypsum is used to make plasterboard for lining the walls and ceilings of houses.
- Granite, limestone and sandstone are used as building stone.

Although a salt crystal is hard, a strong force such as a hammer blow will shatter the crystal. Therefore, it is said to be brittle. This is because the layers of ions will move relative to each other due to the force of the blow. During this movement, ions of like charge are shifted so they are next to each other, as seen in Figure 7.5.6. The resulting repulsion between the similarly charged ions causes the crystal to shatter.

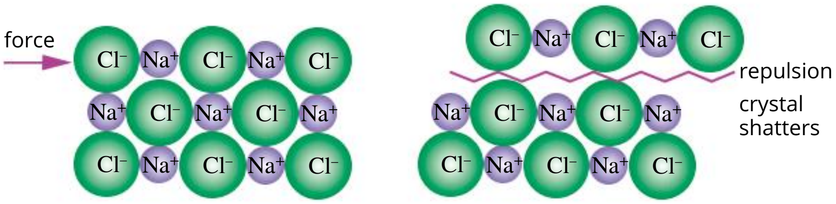


FIGURE 7.5.6 A lattice of an ionic compound shattering. Note that just before shattering, the Cl^- ions are adjacent to other Cl^- ions and the Na^+ ions are also next to each other.

FIGURE 7.5.7 A coffee cup being smashed

Materials made from clay, such as kitchen crockery (Figure 7.5.7), ceramic tiles and bricks are hard, but they are also brittle. Porcelain is a type of chinaware made from a clay called kaolin. Clay is weathered rock and consists of a mixture of complex ionic compounds. The chemical formula for kaolin is $Al_2Si_2O_5(OH)_4$. The ions in kaolin are aluminium, (Al^{3+}), silicate, ($(Si_2O_5)^{2-}$), and hydroxide, (OH^-).

ELECTRICAL CONDUCTIVITY

In the solid form, ions in sodium chloride are held in the crystal lattice and are not free to move, so solid sodium chloride does not conduct electricity. Remember that for a substance to conduct electricity, it must contain charged particles that are free to move. Figure 7.5.8 shows how the particles are arranged in an ionic compound in solid form.

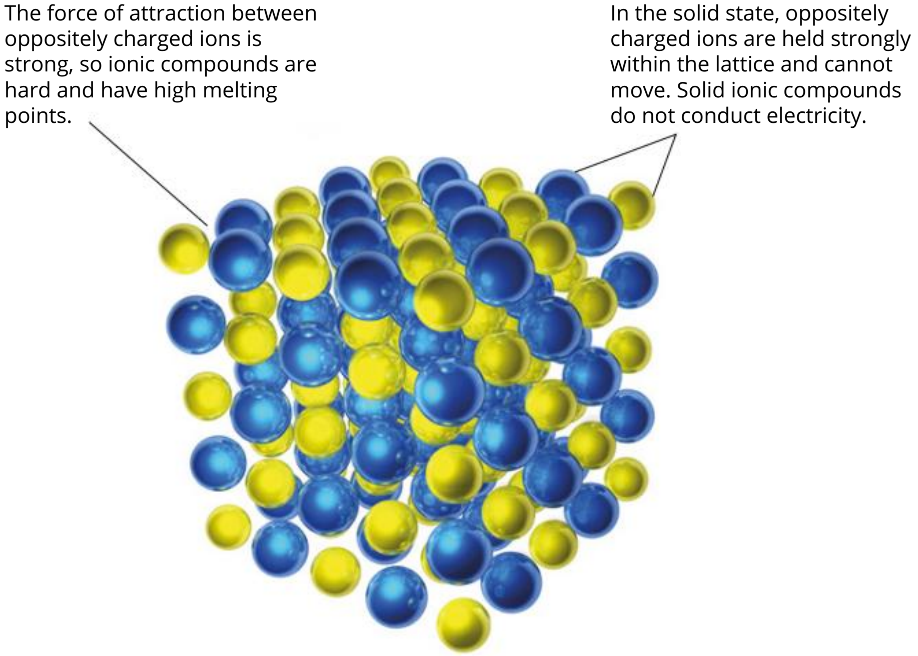


FIGURE 7.5.8 The arrangement of ions within a solid ionic compound. The strong crystal lattice structure explains its physical properties.

FIGURE 7.5.9 A ceramic insulator on the post of an electric fence

The non-conducting property of ionic compounds is used in **ceramic insulators**, which are used to keep high-voltage powerlines insulated from electricity poles and electric fence wires (Figure 7.5.9).

When solid ionic compounds **melt**, the ions become free to move, enabling the cations and anions in the molten compound to conduct electricity.

Similarly, when ionic compounds **dissolve** in water, ionic bonds in the lattice are broken and the ions are separated and move freely in solution.

When an electric **current** is applied to either a molten ionic compound or a solution of the compound in water, positive ions move towards the negatively charged **electrode** and negative ions move towards the positively charged electrode, resulting in an electric current as shown in Figure 7.5.10. A solution or molten substance that conducts electricity by means of the movement of ions in the liquid is called an **electrolyte**.

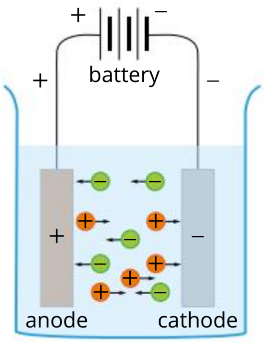


FIGURE 7.5.10 A molten ionic compound will conduct an electric current.

For an object to conduct electricity, it must contain charged particles that are free to move. Solid ionic compounds are made up of a crystal lattice, so the particles are not free to move and they cannot conduct electricity. When molten (melted) or dissolved in water, the charged particles are free to move and conduct.

Electrolysis

When an electric current passes through a molten ionic compound, or a solution of the compound, chemical reactions occur at the positive and negative electrodes. This process is known as **electrolysis** and is used to make a variety of chemicals. Chemicals formed by electrolysis are often difficult to obtain by other means.

For example, the final stage in the industrial production of aluminium involves passing an electric current through a molten liquid containing aluminium oxide (alumina). Aluminium metal is deposited at the negatively charged electrode.

Sodium metal is a starting material for the manufacture of a range of organic chemicals, including dyes. It is also produced on a large scale by electrolysis. In this case, electricity is passed through molten sodium chloride to make sodium. During this process, sodium ions move towards the negative electrode where they become sodium metal, and chloride ions move to the positive electrode and are converted into chlorine gas (Figure 7.5.11).

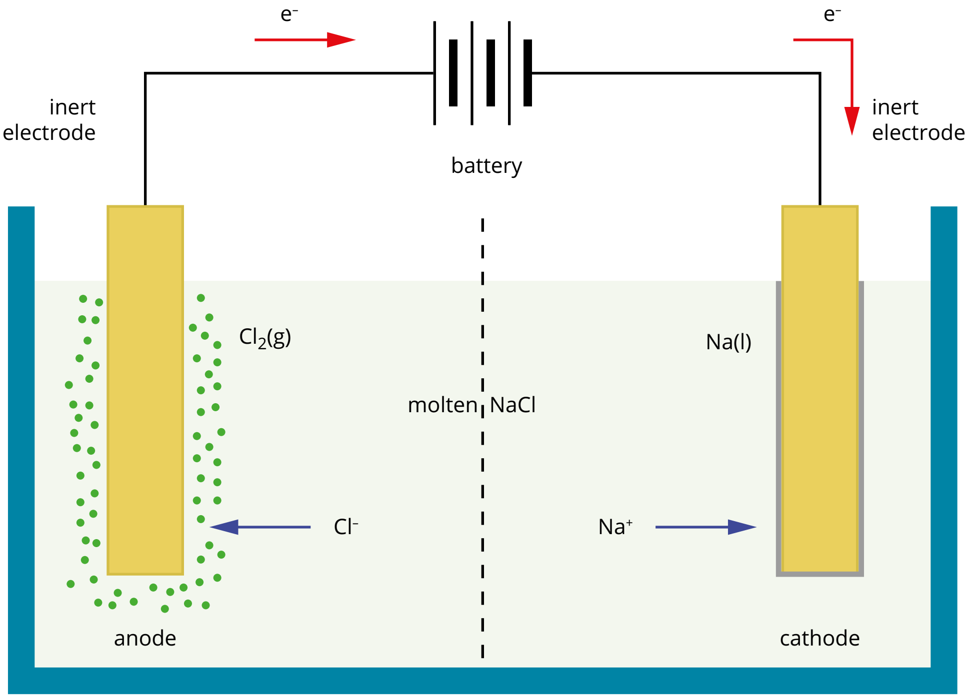


FIGURE 7.5.11 The production of sodium metal and chlorine gas by the electrolysis of molten sodium chloride

SOLUBILITY

Some ionic compounds are very **soluble** in water, whereas others are very **insoluble**. When a soluble ionic compound is added to water, the ions are pulled away from the ionic lattice and mix with the water molecules. If an insoluble compound is added to water, the ions remain bonded together in the ionic lattice and do not form a solution.

Whether an ionic compound is soluble or insoluble depends on the relative strength of the forces of attraction between the:

- positive and negative ions in the lattice
- water molecules and the ions.

Sodium hypochlorite is an example of a highly soluble ionic compound. It is used as a bleach and in swimming pools to kill microorganisms. You will look at the solubility of ionic compounds in water in more detail in Chapter 15.

Nanocomputers

Professor Michelle Simmons was awarded Australian of the Year in 2018). Her research group at the Australian Research Council Centre of Excellence for Quantum Computation and Communication Technology (UNSW, Sydney) has made the world's smallest transistor by positioning a single phosphorus atom inside a silicon structure (Figure 7.5.12). Another atomic precision device the group has made is a conducting wire which is only four atoms wide but conducts with the same efficiency as copper wire. A PhD student, Bent Weber, published this finding in the journal *Science*. These and other projects will allow Professor Simmons' group to, in the future, construct a 10-qubit quantum computer which enables faster and more efficient processing of large amounts of information and databases. They have already made 2-, 3- and 4-qubit structures.

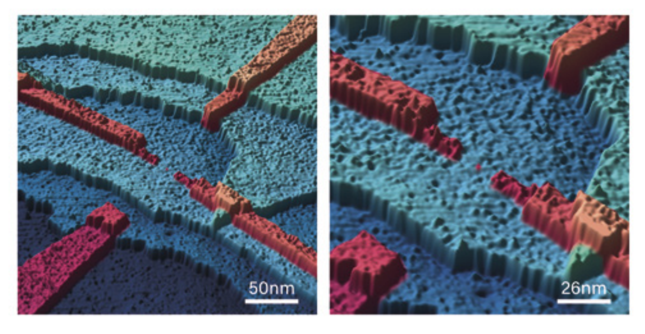


FIGURE 7.5.12 A tiny transistor with a central phosphorus atom sitting on silicon atoms. This image was produced by a one tonne scanning tunnelling microscope (STM).

Using the scanning tunnelling microscope (STM), individual atoms, as seen in Figure 7.5.13, are picked up and moved around with an iridium wire with copper or platinum atoms on the tip. When it is lowered near the base plate the non-metal (phosphorus) attempts to ionically bond with the metal atoms on the probe, and can be lifted and moved. Ionic crystals are used to move the tip of the probe nanometres at a time in various directions (back/forth, up/down). They can move the copper probe extremely accurately because the crystals are piezoelectric. This means that they do not conduct electricity, but when a voltage is applied the crystal changes shape very slightly. This movement is enough to move the scanning probe tip very small distances in a controlled way.

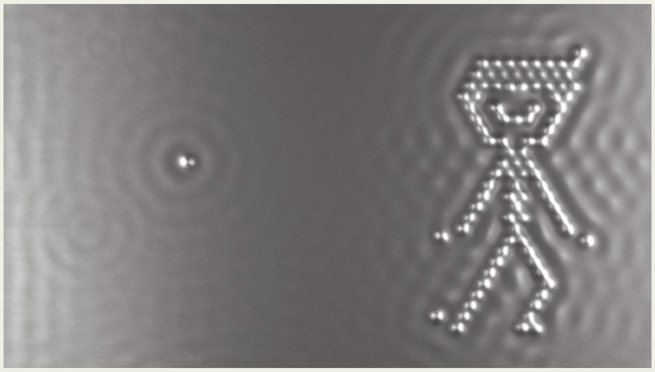


FIGURE 7.5.13 While practising moving atoms using an STM, scientists at the technology company IBM in the United States have made an entertaining stop-motion nano-movie called 'A boy and his atom'. The background is a copper sheet and the 'atoms' are actually molecules of carbon monoxide.

Review

1 Piezoelectricity has two opposite effects—it has the movement of electric charge through an ionic crystal under pressure, or the change to the shape of an ionic crystal when an electric current is applied. For both effects, a change in shape of an ionic crystal is linked to electric current through the crystal. Research an example of the use of ionic crystal(s) where pressure results in an electric current.

2 Review Figure 7.5.13 and identify evidence that the surface atoms are actually dimers of carbon monoxide, not single atoms.

3 Propose a reason why a scanning tunnelling microscope creates the clearest images at –269°C. (Hint: use kinetic energy ideas.)

7.5 Review

SUMMARY

- Ionic compounds are hard and have high melting and boiling points. This is because of the strong forces of attraction between the positively and negatively charged ions in the ionic lattice.
- When an ionic compound is hit, the ions move within the lattice so that like-charged ions line up opposite each other and then repel, causing the lattice to be disrupted. This makes ionic compounds brittle.
- Ionic compounds do not conduct electricity in the solid state. Although the solid ionic lattice contains charged particles, the particles are not free to move.
- When ionic compounds are added to water or are in molten form, the charged particles are free to move, which means they can conduct electricity.
- In water, ionic compounds vary from very soluble to insoluble. The solubility depends on whether the forces between the water molecules and the ions in the lattice are strong enough to pull the ions out of the lattice.
- Ionic compounds are useful because of their physical properties such as hardness and high melting points.

KEY QUESTIONS

Retrieval

1 Use the ionic bonding model to explain why ionic compounds that have been heated above their melting point conduct electricity but will not when in solid form.

2 State why ionic compounds have higher melting points.

Comprehension

3 **a** Describe what happens to ions in an ionic compound when the compound is hit with a heavy blow.

b Explain why this makes ionic compounds brittle.

Analysis

4 Australia is a major exporter of metal ores (metal oxides) and metals. Consider one of the ores below and identify the cation(s) and anion(s) present as well as the metal that is recovered.

- bauxite
- sphalerite
- haematite

5 Complete this summary table of properties of ionic compounds and the information this provides about its structure.

Property	**What this tells us about structure**
	Forces between the particles are strong.
hard, brittle crystals	
does not conduct electricity in the solid state	No free-moving charged particles are present in the solid.
conducts electricity in the molten state or in aqueous solution	
	The strength of the interaction of ions with water particles varies as does the strength of bonds within the lattice.

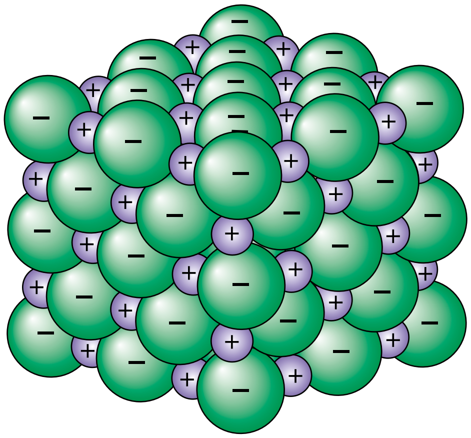
6 The melting points of ionic compounds containing sodium and a range of anions (simple and polyatomic) are shown in the table below.

a Based on the data in the table below, predict and explain whether the melting point of NaF will be higher or lower than NaCl.

b As melting point is related to the strength of the bonds in the ionic compound, explain why compounds containing polyatomic ions have lower melting points.

	NaF	**NaCl**	**NaI**	**$NaNO_2$**	**NaH_2PO_4**
Melting point (°C)	?	801	661	271	250
Type of anion	simple	simple	simple	polyatomic	polyatomic

You may use this diagram of the structure of NaCl below as a starting point for your discussion.



Chapter review

WS 1.1.8

WS 1.1.9

07

KEY TERMS

anion
aqueous
boiling point
brittle
cation
ceramic
ceramic insulator
charge imbalance
chemical formula
compound
conductivity
crystal lattice
current
dissolve
electrode
electrolysis
electrolyte
electron transfer diagram
electronegative
electrostatic attraction
element
formula unit
insoluble
ion
ionic bonding
ionic compound
melt
melting point
metal ion
non-metal ion
non-polar solvent
octet rule
polyatomic ion
soluble
three-dimensional lattice
transition metal
valence shell

KEY QUESTIONS

Retrieval

1 State the electron configurations of the following ions.
- **a** Na^+
- **b** O^{2-}
- **c** Mg^{2+}
- **d** N^{3-}

2 Describe what the subscripts in the formula of an ionic compound tell you about the arrangement of cations and anions.

3 Use Table 7.4.1 page 182 to identify:
- **a** the anion containing the most atoms
- **b** the anion with the greatest charge.

4 Use the ionic bonding model to explain the following properties of ionic compounds.
- **a** They generally have high melting points.
- **b** They are hard and brittle.
- **c** They do not conduct electricity in the solid state but will conduct when molten or dissolved in water.

5 Use the periodic table in Figure 7.2.1 on page 173 and your understanding of its arrangement of metal and non-metal elements to identify which of the following are likely to be ionic substances.
- **A** sodium bromide
- **B** magnesium chloride
- **C** silicon dioxide
- **D** phosphorus trichloride

Comprehension

6 Determine which of the following is *not* a property of ionic compounds.
- **A** poor conductor of electricity
- **B** high melting point
- **C** soluble
- **D** ductile

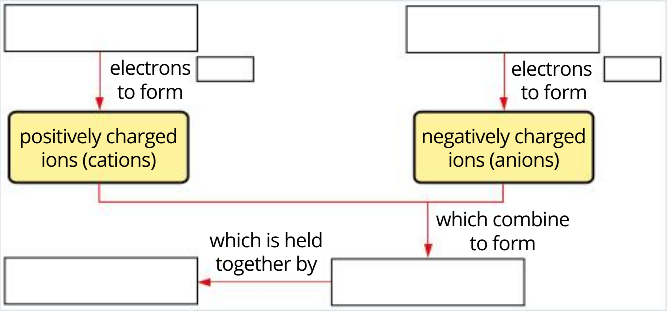
7 Select from the following options what the change to electrical conductivity is due to when ionic compounds are molten or dissolved.
- **A** the effect of heat in removing electrons from ions
- **B** water molecules carrying charge
- **C** ions becoming free to move (carrying charge)
- **D** electrons passing through the lattice structure

8 Select from the following options what the charge on X in XI_2 is.
- **A** +2
- **B** +1
- **C** –1
- **D** –2

9 Determine which of the following are ionic compound(s).
- **A** copper(II) nitrate
- **B** ammonia NH_3
- **C** potassium oxide
- **D** boron trichloride

10 Determine whether the following properties relate to metallic lattices only, ionic lattices only or both metallic and ionic lattices.
- **a** They contain both positively and negatively charged particles.
- **b** The lattice is held together by forces of attraction between positively and negatively charged particles.
- **c** They are hard.
- **d** They are brittle.
- **e** They conduct electricity in the molten state.

11 Show what happens when ionic compounds are formed from metallic and non-metallic elements by completing the following diagram.



12 Determine what the ionic compounds with the following chemical formulas are.

- **a** $(NH_4)_2CO_3$
- **b** $Cu(NO_3)_2$
- **c** $CrBr_3$

13 Describe what happens to the structure of solid sodium chloride when it melts.

14 From the following list of some of the uses of ionic compounds explain what property of ionic compounds enables them to be used in this way.

- **a** insulators on electrified fences
- **b** bricks for building the wall of a house
- **c** one of the chemicals in a battery

15 Use electron configuration and a balanced chemical equation to show the electron transfer that occurs when:

- **a** lithium reacts with chlorine
- **b** magnesium reacts with fluorine
- **c** potassium reacts with sulfur
- **d** magnesium reacts with nitrogen

16 Explain why elements in group 17 of the periodic table are likely to forms ions with a –1 charge.

17 Write the chemical formula for the ionic compound formed in each of the following reactions.

- **a** potassium and bromine
- **b** magnesium and iodine
- **c** calcium and oxygen
- **d** aluminium and fluorine
- **e** calcium and nitrogen

18 Show the chemical formula that represents each of the following ionic compounds.

- **a** copper(I) chloride
- **b** silver(I) oxide
- **c** lithium nitride
- **d** potassium iodide

19 Complete the table below by writing the chemical formulas for four ionic compounds formed using combinations of cations and anions as shown.

	K^+	Ca^{2+}
F^-		
N^{3-}		

20 Use the diagrammatic representations shown in Figure 7.4.1 on page 183 to show the formation of magnesium sulfate.

21 Draw a concept map to show the connection between the terms: metal, non-metal, atom, valence electron, anion, cation, electrostatic attraction, ionic bonding.

Analysis

22 Apply the concept of a formula unit to explain why the chemical formula of an ionic compound is not just the actual number of atoms in the compound.

23 The electron configurations of some metallic and non-metallic elements are given below. (The symbols shown for the elements are not their real ones.) Write formulas for the compounds they are most likely to form if they react together. The first example has been done for you.

a	A: 2,1	B: 2,6	A_2B
b	C: 2,8,3	D: 2,7	
c	E: 2,8,6	F: 2,8,8,2	
d	G: 2,5	H: 2,8,8,1	
e	K: 2,8,2	L: 2,6	

24 Refer to the periodic table on page 173 and, for each general formula given, identify two elements that will react to form an ionic compound with that formula. (Remember the metal ion, as represented by M, is written first in each formula.)

- **a** MY_2
- **b** MY
- **c** M_2Y
- **d** M_3Y
- **e** MY_3
- **f** M_3Y_2

25 Represent the transfer of electrons between the following metal and non-metal atoms using an equation and write a second equation showing formation of the ionic compound.

- **a** magnesium and chlorine atoms
- **b** aluminium and oxygen atoms

26 For the *same amount* of the following non-metals, identify which one of the following will react with potassium so that the maximum number of metal atoms is converted to a compound. Explain your answer.

- **A** sulfur
- **B** oxygen
- **C** chlorine
- **D** nitrogen

27 Use the information in Tables 7.3.1 and 7.3.2 on page 179 and Table 7.4.1 on page 182 to write formulas for the following ionic compounds.

- **a** iron(III) oxide
- **b** copper(II) oxide
- **c** chromium(III) sulfate
- **d** iron(II) oxide
- **e** lead(II) nitrate
- **f** lead(IV) sulfide
- **g** tin(II) nitrate

28 The elements M, Y and Z form ionic compounds when they react with other elements. The following compounds are formed: Ca_3M_2, Y_2CO_3 and Al_2Z_3.

- **a** Determine the charge on the ion formed by:
  - **i** element M
  - **ii** element Y
  - **iii** element Z
- **b** Use these charges on the ions to deduce correct chemical formulas for the:
  - **i** sulfate salt of Y
  - **ii** potassium salt of Z
  - **iii** ionic compound formed between M and Y
  - **iv** ionic compound formed between Y and Z

29 Identify the formula unit for the following ionic compounds.

- **a** copper(I) nitrate
- **b** chromium(II) fluoride
- **c** potassium carbonate
- **d** magnesium hydrogen carbonate
- **e** nickel(II) phosphate

30 The melting point of sodium chloride is 801°C, whereas that of magnesium oxide is 2800°C.

- **a** Identify the particles present in the two solids.
- **b** Conclude which solid has the stronger forces between its particles.
- **c** Suggest how the differences between the two sets of ions in each lattice contribute to the differences in the melting points.

Knowledge utilisation

31 The formula of sodium perchlorate is $NaClO_4$ and that of potassium ferrocyanide is $K_4Fe(CN)_6$. Propose the formula for the following, using this information:

- **a** calcium perchlorate
- **b** aluminium ferrocyanide
- **c** iron(III) perchlorate
- **d** ammonium ferrocyanide

32 Determine how the shape of copper sulfate crystals and sodium chloride crystals shown below reflect their ionic structure. Propose possible reasons for differences and explain the particles making up the lattice in each structure.

crystals of copper sulfate

crystals of sodium chloride

33 Propose an experiment you could carry out to demonstrate each of the following properties of the compounds given. In each case:

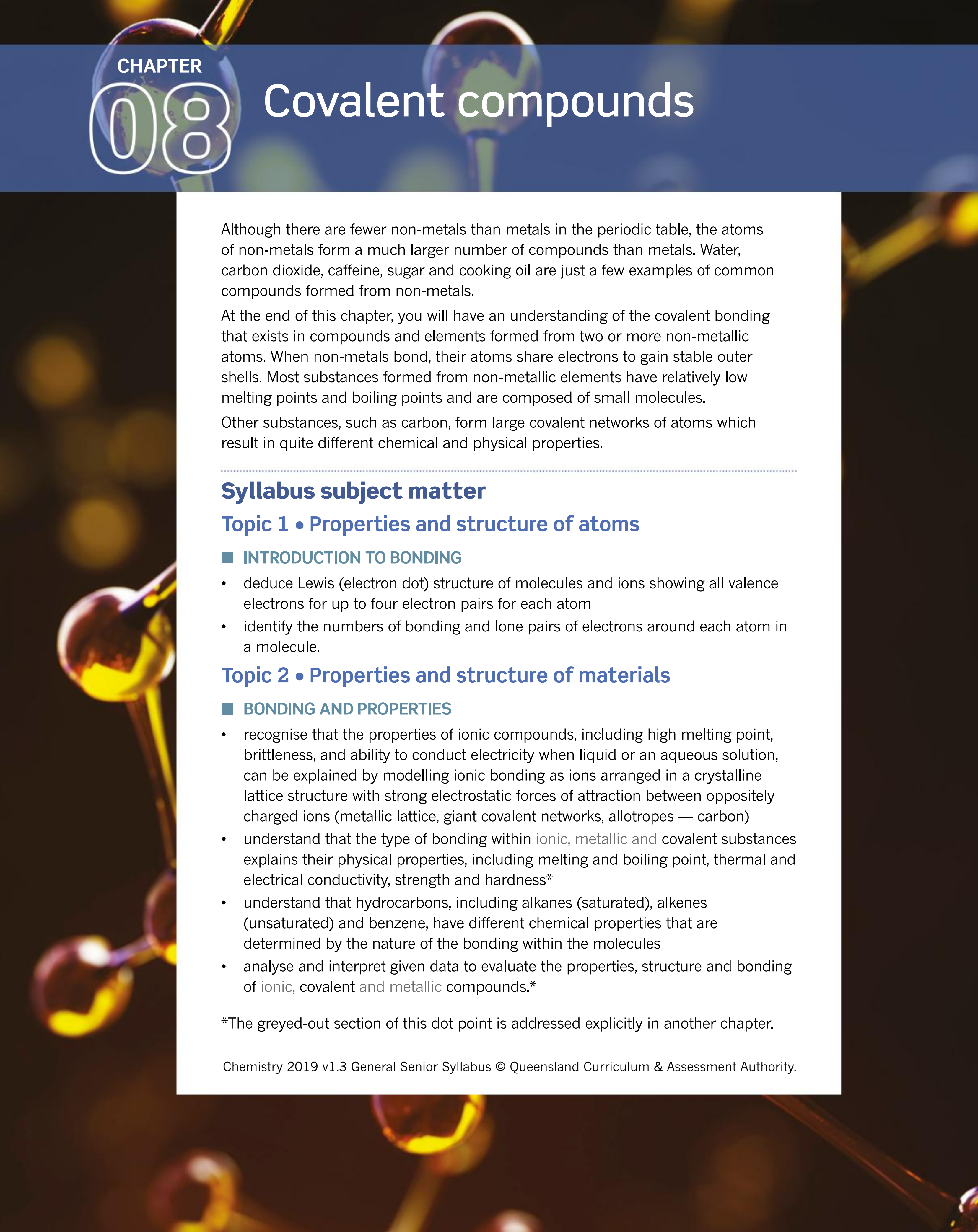
- **i** sketch the equipment you would use
- **ii** describe what you would expect to observe.

- **a** Solid magnesium chloride does not conduct electricity.
- **b** Dissolved sodium chloride does conduct electricity.
- **c** Solid sodium chloride is hard and brittle.

34 A student analyses the structure and bonding in a metal and an ionic compound. From this analysis, she makes the statements below.

- **a** Metals and ionic compounds both contain positive ions in a regular arrangement.
- **b** In metals and ionic compounds, there is attraction between one particle and all the neighbouring particles of opposite charge.
- **c** In metals and ionic compounds, there will be forces of repulsion between particles with like charges.
- **d** A metal will conduct electricity, whereas an ionic compound will not because electrons are much smaller than negative ions.

Judge whether you agree or disagree with each statement. Provide reasons for your decisions.



CHAPTER 08

Covalent compounds

Although there are fewer non-metals than metals in the periodic table, the atoms of non-metals form a much larger number of compounds than metals. Water, carbon dioxide, caffeine, sugar and cooking oil are just a few examples of common compounds formed from non-metals.

At the end of this chapter, you will have an understanding of the covalent bonding that exists in compounds and elements formed from two or more non-metallic atoms. When non-metals bond, their atoms share electrons to gain stable outer shells. Most substances formed from non-metallic elements have relatively low melting points and boiling points and are composed of small molecules.

Other substances, such as carbon, form large covalent networks of atoms which result in quite different chemical and physical properties.

Syllabus subject matter

Topic 1 • Properties and structure of atoms

INTRODUCTION TO BONDING

- deduce Lewis (electron dot) structure of molecules and ions showing all valence electrons for up to four electron pairs for each atom
- identify the numbers of bonding and lone pairs of electrons around each atom in a molecule.

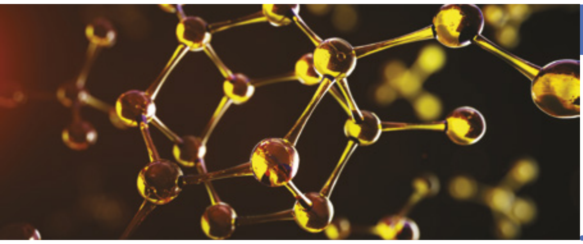
Topic 2 • Properties and structure of materials

BONDING AND PROPERTIES

- recognise that the properties of ionic compounds, including high melting point, brittleness, and ability to conduct electricity when liquid or an aqueous solution, can be explained by modelling ionic bonding as ions arranged in a crystalline lattice structure with strong electrostatic forces of attraction between oppositely charged ions (metallic lattice, giant covalent networks, allotropes — carbon)
- understand that the type of bonding within ionic, metallic and covalent substances explains their physical properties, including melting and boiling point, thermal and electrical conductivity, strength and hardness*
- understand that hydrocarbons, including alkanes (saturated), alkenes (unsaturated) and benzene, have different chemical properties that are determined by the nature of the bonding within the molecules
- analyse and interpret given data to evaluate the properties, structure and bonding of ionic, covalent and metallic compounds.*

*The greyed-out section of this dot point is addressed explicitly in another chapter.

8.1 Properties of non-metallic substances



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand some of the properties common to non-metallic substances
- describe a molecule
- recognise the differences between intermolecular bonds and intramolecular bonds.

In Chapter 6, you saw that the bonding between atoms in metallic elements is called metallic bonding. In Chapter 7, you saw that when metal atoms combine with atoms of non-metallic elements, the compounds contain another form of bonding, ionic bonding. In this chapter, you will look at the chemical bonding that occurs when atoms of non-metals combine with each other.

PROPERTIES OF NON-METALLIC ELEMENTS AND COMPOUNDS

The properties of compounds formed by non-metallic elements can tell you a lot about their chemical structures. Table 8.1.1 lists some of the properties of a range of common substances. The substances chosen include metallic elements, non-metallic elements, ionic compounds and molecular compounds. The information in Table 8.1.1 is useful because it allows you to:

- identify the characteristic properties of each category of substances. For example, metals conduct electricity as solids and liquids
- deduce information about the chemical structures of these substances. If a substance has a high melting point, the bonding between particles must be strong. If a substance conducts electricity, it must contain charged particles that are free to move.

TABLE 8.1.1 Properties of some common metals, ionic compounds and non-metallic substances

Substance	Formula	Melting point (°C)	Boiling point (°C)	Conducts electricity as a solid?	Conducts electricity as a liquid?
ammonia	NH_3	−77	−33	no	no
copper	Cu	1084	2562	yes	yes
chlorine	Cl_2	−102	−34	no	no
copper(II) chloride	$CuCl_2$	498	993	no	yes
nitrous oxide	N_2O	−90	−88	no	no
sodium chloride	NaCl	801	1413	no	yes
oxygen	O_2	−219	−183	no	no
water	H_2O	0	100	no	no

Table 8.1.1 shows some of the properties of common substances.

- Metals such as copper have high melting points and conduct electricity both as solids and as liquids. As you saw in Chapter 6, this indicates that strong bonds exist in metals and that they contain charged particles (delocalised electrons) that are free to move along the metal.
- The ionic compounds copper(II) chloride and sodium chloride have high melting points and conduct electricity as liquids but not as solids. In Chapter 7, you learnt that ionic solids have strong bonds and that they contain charged particles (ions) that are free to move in a liquid form but not in a solid form.
- The non-metallic elements chlorine and oxygen have low melting points and do not conduct electricity. The same is true of the molecular compounds ammonia, nitrous oxide and water. Non-metallic substances can also be soft in the solid state. Two conclusions about bonding in non-metallic substances that can be drawn from these properties are shown in Figure 8.1.1.

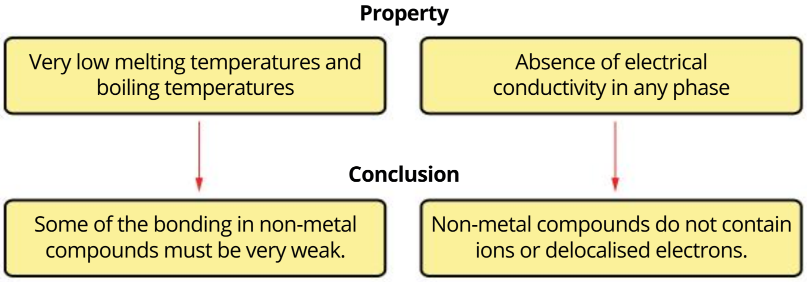
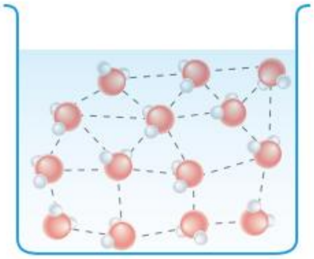


FIGURE 8.1.1 These conclusions can be drawn from the properties of non-metallic substances.

Molecules

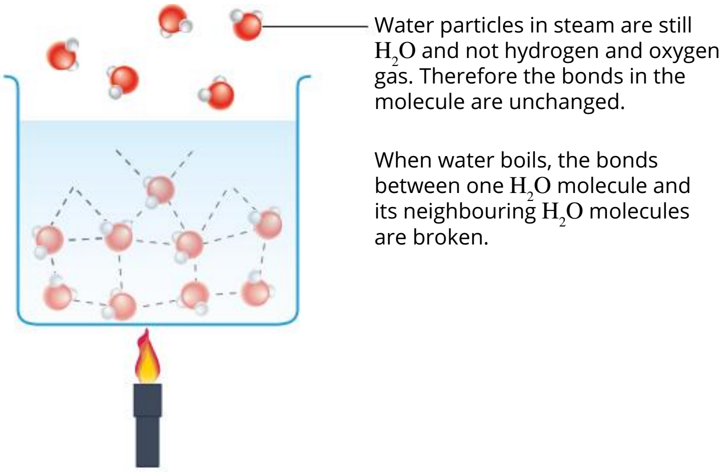
You gain further insight into the chemical structure of non-metallic substances by considering what happens when you boil water in a beaker. Figure 8.1.2 shows that each water **molecule** contains two hydrogen atoms bonded to one oxygen atom. In between the molecules are weak bonds holding the molecules to each other. This type of intermolecular bonding will be covered in more detail in Chapter 12.



Water in a beaker contains two types of bonds:
1. the bonds between hydrogen and oxygen atoms within each water molecule
2. the bonds holding one molecule of water to another.

FIGURE 8.1.2 This representation shows the bonds between water molecules in a beaker.

When water boils, the weak bonds between the molecules must break (Figure 8.1.3). However, the water molecules do not separate into hydrogen and oxygen. Rather, the water vapour that is formed still contains water molecules in which two hydrogen atoms are bonded to one oxygen atom. This indicates that liquid water must contain more than one form of bonding.


Water particles in steam are still H_2O and not hydrogen and oxygen gas. Therefore the bonds in the molecule are unchanged.

When water boils, the bonds between one H_2O molecule and its neighbouring H_2O molecules are broken.

FIGURE 8.1.3 Changes occur to the bonding in water when water starts to boil.

The particles of water shown in Figure 8.1.3 are examples of a molecule. A molecule is a definite and discrete group of atoms chemically bonded together. The atoms in molecules are non-metallic atoms bonded to other non-metallic atoms. Molecules can be found in some elements as well as in some compounds. The element oxygen, for example, forms O_2 molecules. Water can be referred to as a **molecular compound** with a **molecular formula** of H_2O. The bonds between the hydrogen and oxygen atoms within water molecules are referred to as **intramolecular bonds** and the bonds between water molecules are referred to as **intermolecular bonds**.

The intramolecular bonds between the hydrogen and oxygen atoms in water molecules are strong compared to the intermolecular bonds between the water molecules. It is the intermolecular bonds that are broken when molecular substances such as water boil or melt. This allows the molecules to separate from each other while the atoms within the molecules remain bound to one another.

8.1 Review

SUMMARY

- Many substances contain only non-metal atoms. Some of these substances are elements, while others are compounds.
- Non-metallic elements and compounds usually have low boiling points and do not conduct electricity.
- Many non-metallic elements and compounds are composed of molecules. A molecule is a definite and discrete group of atoms chemically bonded together. The atoms in molecules are non-metallic atoms bonded to other non-metallic atoms.
- There are two important types of bonds in molecular compounds—the bonds within the molecules (intramolecular bonds) and the bonds between molecules (intermolecular bonds).
- The melting and boiling points of non-metallic elements and compounds depend on the bonding between molecules. Generally, molecular substances have low melting and boiling points. This indicates they have weak bonds between molecules. The weak bonds mean that solid compounds of non-metallic elements can be soft.
- In general, non-metallic elements and compounds do not conduct electricity because they do not contain free-moving charged particles (neither delocalised electrons nor ions).

KEY QUESTIONS

Retrieval

1 Identify the non-metallic elements and compounds from this list.
  - **a** $CaSO_4$
  - **b** N_2O_5
  - **c** H_2
  - **d** Cu
  - **e** $FeCl_2$
  - **f** I_2

2 Define the term 'molecule'.

Comprehension

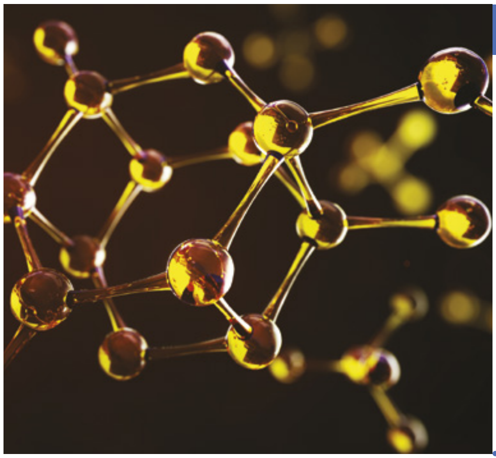
3 Explain the following general properties of non-metallic substances.
  - **a** Non-metallic substances tend not to conduct electricity.
  - **b** Non-metallic substances have low melting points and boiling points.

Analysis

4 Explore the difference between the boiling points of methane (CH_4, b.p. = –161°C) and water (H_2O, b.p. = 100°C).

5 When sugar (i.e. sucrose, $C_{12}H_{22}O_{12}$) is gently heated, it turns into a clear liquid. If the liquid is heated strongly, it turns black and a gas is produced. Determine what is happening to the bonds in sugar when it is heated. Use the terms 'intermolecular bonds' and 'intramolecular bonds' in your answer.

8.2 Covalent bonding



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that covalent bonds involve sharing of valence electrons between atoms
- recall that non-metallic atoms can bond with other non-metallic atoms via single, double or triple bonds
- identify trends in bond lengths and bond strengths depending on whether covalent bonds between atoms are single, double or triple
- represent the structure of polyatomic molecules and ions using Lewis structures
- identify the number of bonding pairs of electrons and non-bonding pairs of electrons in a molecule or ion.

This module examines a series of simple molecules to help you to understand the concept of a **covalent bond**, which is formed when non-metallic atoms share electrons. Using your knowledge of the valence shell electron arrangements of non-metallic atoms, you will be able to predict the molecules that different elements can form.

COVALENT BONDS

In Chapter 3, you saw that many atoms become more stable if they gain an outer shell of eight electrons by combining with other atoms (the **octet rule**).

Commonly, when atoms of non-metals combine, electrons are shared so that each atom has eight electrons in its outer shell. Molecules formed in this way are more stable than the separate atoms. The major exception to the octet rule is hydrogen which can only accommodate a maximum of two electrons in its outer shell.

Non-metallic atoms have a relatively high number of electrons in their outer shells and they tend to share rather than to transfer electrons. Covalent bonding occurs when electrons are shared between atoms.

Single covalent bonds

When atoms share two electrons, one from each atom, the covalent bond formed is called a **single covalent bond**. Two examples of substances that contain single bonds are hydrogen and chlorine.

Hydrogen

Hydrogen atoms have one electron. The valence shell for a hydrogen atom can hold a maximum of two electrons. A hydrogen atom can bond to another hydrogen atom to form a molecule of H_2, as shown in Figure 8.2.1.

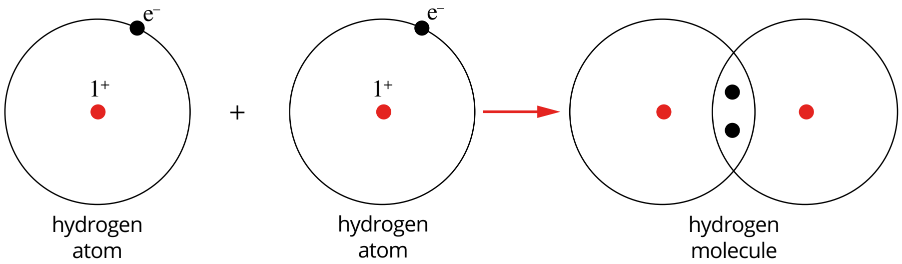


FIGURE 8.2.1 A covalent bond is formed when two hydrogen atoms share two electrons, one from each atom.

In the molecule that is formed:

- two hydrogen atoms share two electrons, one from each atom, to form a single covalent bond
- the atoms of hydrogen are now strongly bonded together by two electrons (an electron pair) in their outer shells.

The hydrogen molecule can be represented as H_2. Molecules that contain two atoms are called **diatomic molecules**.

Two alternative ways of representing a hydrogen molecule are shown in Figure 8.2.2.

In a hydrogen molecule, the two electrons are attracted to both protons of the H_2 structure. This means the two electrons will spend most of their time between the two nuclei instead of orbiting their original nucleus. Even though the protons in the two nuclei still repel each other (remember, like charges repel), the electrostatic attraction to the electrons (which are closer) keeps the molecule held together.

This model of a hydrogen molecule is consistent with the observed properties of hydrogen gas. The covalent bonds in the molecules are strong, but the intermolecular forces in hydrogen (the attractions between one molecule and the surrounding molecules) are weak. This explains the low melting point of −259°C for hydrogen. Hydrogen does not conduct electricity as it does not contain ions or delocalised electrons.

The hydrogen molecule (H_2) is an example of a covalent molecular substance.

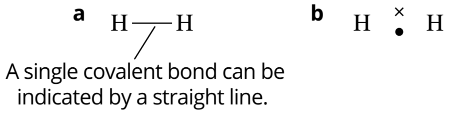


FIGURE 8.2.2 Two representations of a hydrogen molecule: (a) the single covalent bond is shown as a single straight line or (b) the electrons in the covalent bond are represented using dots or crosses. In this case, the dot and cross represent electrons that originated from two different hydrogen atoms.

Chlorine

A chlorine atom has an electron configuration of 2,8,7. It requires one more electron to achieve eight electrons in its outer shell. (Note: throughout this chapter, shell-model electron configurations and diagrams will be used to represent the electron arrangement in atoms as simply as possible.)

One chlorine atom can share an electron with another chlorine atom to form a molecule of chlorine with a single covalent bond. As a result, both atoms gain outer shells of eight electrons as shown in Figure 8.2.3. This is an example of the application of the octet rule.

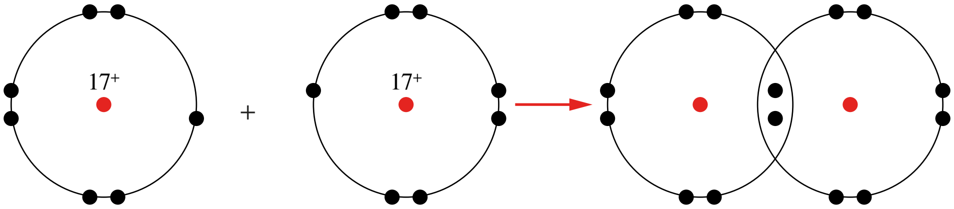


FIGURE 8.2.3 Two chlorine atoms share one electron each to form a chlorine molecule. Note that only outer-shell electrons are shown in these diagrams.

Electron dot diagrams and Lewis structures

Chemists often use **electron dot diagrams** (also known as electron dot formulas) and **Lewis structures** to simplify the drawing of molecules.

Electron dot diagrams show the valence shell electrons of an atom, as only electrons in the valence shells of atoms are involved in bonding. In a Lewis structure, lines are used to represent the two electrons in a covalent bond. One line represents one pair of bonding electrons.

The Lewis structure also allows you to distinguish between bonding electrons and non-bonding electrons. A chlorine molecule has one pair of bonding electrons represented by the straight line between the two Cl atoms. The outer-shell electrons that are not involved in bonding are called the **non-bonding electrons**. Each chlorine atom has six non-bonding electrons, present as three pairs of electrons. Pairs of non-bonding electrons are known as **lone pairs**.

FIGURE 8.2.4 Electron dot diagrams of two chlorine atoms (on the left) combining to form the Cl_2 molecule via the sharing of one electron from each atom. The structure on the right is the Lewis structure for the Cl_2 molecule.

Figure 8.2.4 shows the electron dot diagram for two separate chlorine atoms and the Lewis structure for a molecule of chlorine. It is easy to see that each chlorine atom in the molecule has a total of eight electrons in its outer shell—six non-bonding electrons (represented as three pairs of dots) and a pair of bonding electrons (represented as a straight line).

Double covalent bonds

In a **double covalent bond**, two pairs of electrons (four electrons in total) are shared between the atoms, rather than just one pair.

The oxygen molecule, O_2, contains a double covalent bond. The electron configuration of an oxygen atom is 2,6. Each oxygen atom requires two electrons to gain a stable outer shell containing eight electrons. Therefore, when one oxygen atom bonds to another one, each atom shares two electrons.

As you can see in Figure 8.2.5, each oxygen atom in the molecule now has eight outer-shell electrons, four of these are bonding electrons and four are non-bonding electrons.

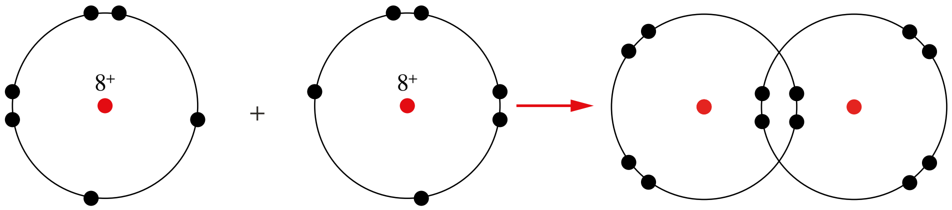


FIGURE 8.2.5 In oxygen molecules, each oxygen atom contributes two electrons to the bond between the atoms.

FIGURE 8.2.6 Electron dot diagrams of two oxygen atoms (on the left) combining to form the O_2 molecule via the sharing of two pairs of electrons. The structure on the right is the Lewis structure for the O_2 molecule.

Figure 8.2.6 shows the electron dot diagram for separate oxygen atoms (on the left) and the Lewis structure of the oxygen molecule, O_2, on the right. The two straight lines represent two pairs of bonding electrons between the oxygen atoms, while the four pairs of dots represent four pairs of non-bonding electrons.

Triple covalent bonds

A **triple covalent bond** occurs when three electron pairs are shared between two atoms. The nitrogen molecule, N_2, contains a triple bond. The electron configuration of nitrogen is 2,5. A nitrogen atom requires three electrons to achieve eight electrons in its outer shell. When it bonds to another nitrogen atom, each atom contributes three electrons to the bond that forms, as shown in Figure 8.2.7. Figure 8.2.8 shows the Lewis structure of the nitrogen molecule. The three pairs of bonding electrons are represented by the three straight lines and two pairs of non-bonding electrons, represented by the dots.

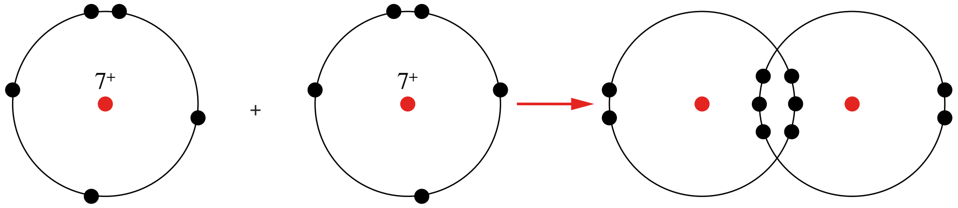


FIGURE 8.2.7 Nitrogen atoms contribute three electrons each to form a triple covalent bond in a molecule of N_2.

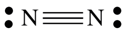


FIGURE 8.2.8 The triple covalent bond for the Lewis structure of the nitrogen molecule, N_2, is drawn as three parallel lines.

The triple covalent bond in nitrogen gas, N_2, is relatively strong and not easily broken. This means that nitrogen gas is relatively unreactive. Nitrogen is an essential element in living organisms because it is a major component of proteins and other biological molecules. Even though 78% of air is nitrogen gas, very few organisms can make use of the nitrogen because it is so unreactive. Only nitrogen-fixing microorganisms are able to convert nitrogen gas into soluble nitrogen-containing compounds. These compounds are then absorbed by plants, allowing nitrogen to then be passed up the food chain.

In the early twentieth century, German chemist Fritz Haber invented a process for converting nitrogen gas and hydrogen gas into ammonia, which is used to make synthetic fertilisers. This allowed more food to be grown to feed a growing world population.

Hydrocarbons containing single, double and triple bonds

Hydrocarbons are molecular compounds containing the elements hydrogen and carbon only. They have the general formula of C_xH_y where the subscripts x and y can take on a wide range of numerical values depending on the specific compound. Figure 8.2.9 shows the Lewis structure for three different hydrocarbons; namely, ethane, ethene and ethyne. Each contains two carbon atoms but each differs in the number of bonds joining them. Ethane contains a single covalent bond between the two carbon atoms, while ethene and ethyne contain a double and triple bond respectively.

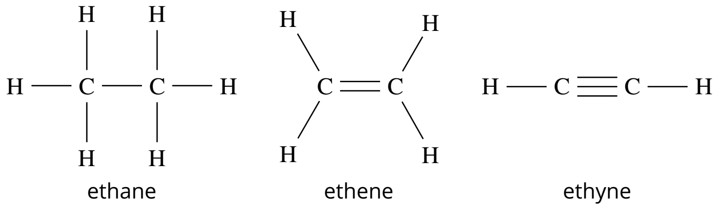


FIGURE 8.2.9 Lewis structures for three simple hydrocarbons showing single, double and triple bonds between carbon atoms

The presence of multiple bonds can greatly affect the chemical and physical properties of compounds. For example, Table 8.2.1 shows that the carbon–carbon bond lengths and bond strengths vary according to the number of bonds joining the two carbon atoms. The bond energy values in Table 8.2.1 are a measure of bond strength and you can see that the carbon–carbon triple bond is the strongest of the three; it is also the shortest. In general, triple bonds are stronger and shorter than double bonds; and double bonds are stronger and shorter than single bonds.

TABLE 8.2.1 Bond properties of hydrocarbons with single, double and triple bonds

Hydrocarbon	C–C bond type	C–C bond energy (kJ mol^{-1})	Bond length (pm)
ethane	single	346	154
ethane	double	602	134
ethyne	triple	835	120

Benzene (C_6H_6) is another important hydrocarbon containing multiple bonds. The Lewis structure for benzene is often represented as a cyclic hexagon structure with alternating single and double bonds (Figure 8.2.10a). This structure suggests that the carbon–carbon bond lengths in benzene are different with the single bonds being longer (i.e. ~154 pm) than the double bonds (~134 pm).

However, spectroscopic analysis confirms that all C–C bond lengths in benzene are equivalent and have a length of ~140 pm, i.e. between that of a double and single bond. A better representation for benzene is shown in Figure 8.2.10b. It shows that the six electrons associated with the three double bonds are actually spread (or delocalised) across the entire cyclic structure, making each C–C bond equal in both length and strength and a hybrid between a single and double bond.

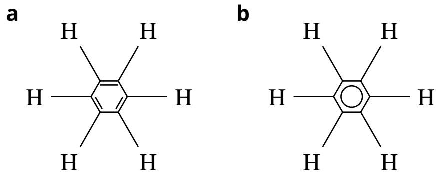


FIGURE 8.2.10 Two different representations of the benzene molecule: (a) a six carbon cyclic structure with alternating single and double bonds, and (b) a six carbon cyclic structure showing the pairs of bonding electrons spread (or delocalised) across the entire molecule

PA
1.2.3

DRAWING LEWIS STRUCTURES FOR MOLECULAR COMPOUNDS

A diatomic molecule contains two atoms. The diatomic molecules discussed so far have been elements, i.e. H_2, Cl_2, O_2 and N_2. Being elements these molecules contain only one type of atom.

Covalent bonds can also form between different types of atoms to form molecular compounds. Hydrogen chloride (HCl) is a simple example (Figure 8.2.11). A hydrogen atom requires one electron to gain a stable outer shell, as does a chlorine atom. They can share an electron each and form a single covalent bond.

FIGURE 8.2.11 Electron dot diagrams of a hydrogen and a chlorine atom (on the left) combining to form the HCl molecule via the sharing of one pair of electrons. The structure on the right is the Lewis structure for the HCl molecule.

Polyatomic molecules

If the atoms of one element have a different number of valence electrons from the atoms of the element it is bonding with, the molecule that is formed may not be a simple diatomic molecule. Molecules made up of more than two atoms are called **polyatomic molecules**. Two examples of polyatomic molecules are water and methane.

Water

When a compound forms between hydrogen and oxygen, an oxygen atom shares two electrons and a hydrogen atom shares one. To resolve this imbalance, two hydrogen atoms each share one electron with an oxygen atom.

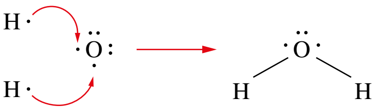


FIGURE 8.2.12 Electron dot diagrams of two hydrogen atoms and an oxygen atom (on the left) combining to form the H_2O molecule via the sharing of one electron from each H atom. The structure on the right is the Lewis structure for the H_2O molecule.

As you can see in Figure 8.2.12, a water molecule contains:

- two single covalent bonds, each containing a shared electron pair (i.e. a total of two bonding pairs of electrons)
- four non-bonding electrons on the oxygen atom (i.e. two non-bonding pairs).

Methane

When a compound forms between carbon and hydrogen, four hydrogen atoms are needed to provide the four electrons required in order to have eight electrons in the outer shell of a carbon atom (Figure 8.2.13). The molecule formed has a chemical formula of CH_4 and is called methane. A methane molecule has four bonding pairs of electrons and zero non-bonding pairs.

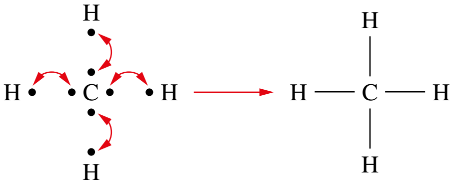


FIGURE 8.2.13 Electron dot diagrams of four hydrogen atoms and an carbon atom (on the left) combining to form the CH_4 molecule via the sharing of one electron from each H atom. The structure on the right is the Lewis structure for the CH_4 molecule.

Steps for drawing Lewis structures

The Lewis structures for HCl, H_2O and CH_4 (Figures 8.2.11, 8.2.12 and 8.2.13) are relatively straightforward examples. When drawing Lewis structures of more complex polyatomic molecules and ions the following steps provide a useful guide:

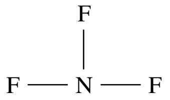
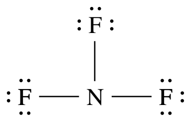
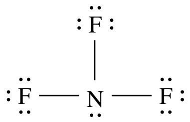
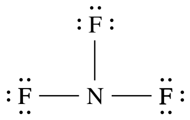
1 Sum the valence electrons from all atoms shown in the molecular formula.
  a For polyatomic anions, add the value of the negative charge to this total.
  b For polyatomic cations, subtract the value of the positive charge from this total.
2 Write atomic symbols for atoms involved to show which atoms are connected to which.
  a You must define a central atom. The central atom is typically the most electropositive (or least electronegative) atom. Electronegativity values are included in Appendix E on page 595.
  b Connect the atoms using a single line representing a single pair of bonding electrons (i.e. a single bond).
  c If the molecule contains double bonds you will address this in rule 5.
3 Complete octets for all peripheral atoms bonded to the central atom.
  a You do this by adding three pairs of electrons to the atomic symbol of the peripheral atom thus giving each atom a total of eight electrons (including the two represented by the single bond).
  b The exception here is hydrogen, which can only accommodate two electrons in its valence shell—for the H atom you must think in terms of a duet rather than an octet.

4 Place any leftover electrons on the central atom.
   a You do this even if it gives the central atom more than an octet. (Exceptions to the octet rule are discussed later.)
5 If there aren't enough electrons to give the central atom an octet, form multiple bonds.
   a You do this by using one or more pairs of non-bonding electrons from the peripheral atoms (choosing the most electropositive, or least electronegative, atom).

Drawing Lewis structures is best shown by example.

Worked example 8.2.1

DRAWING LEWIS STRUCTURES FOR POLYATOMIC MOLECULES

Draw the Lewis structure for nitrogen trifluoride (NF_3).	
Thinking	**Working**
Step 1: Sum the valence electrons from all atoms shown in the molecular formula.	N has the electron configuration of 2,5 and so has five valence electrons. F has the electron configuration of 2,7 and so has seven valence electrons. Sum of all valence electrons in $NF_3 = 5 + (3 \times 7) = 26$
Step 2: Write atomic symbols for all atoms to show which atoms are connected to which. The central atom will be the most electropositive.	In this case, N is the most electropositive and the central atom; the three F atoms will be peripheral atoms.  A total of six electrons have been used so far representing the three pairs of bonding electrons in the three single bonds.
Step 3: Complete octets for all peripheral atoms bonded to the central atom.	Do this by adding three pairs of electrons to the atomic symbol of the peripheral atoms.  A total of 24 electrons have been used in total so far; three bonding pairs on the central N atom and nine non-bonding pairs across the three F atoms.
Step 4: Place any leftover electrons on the central atom.	 So far, 24 of the available 26 valence electrons have been used. There are two leftover electrons, which need to be placed on the central N atom.
Step 5: If there are not enough electrons to give the central atom an octet, form multiple bonds. The Lewis structure is complete.	 This step is not required since the central atom has a total of eight electrons surrounding it; three bonding pairs and one non-bonding pair.

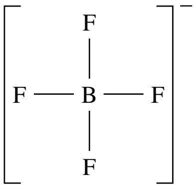
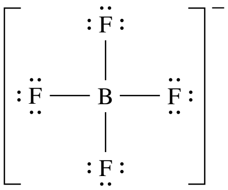
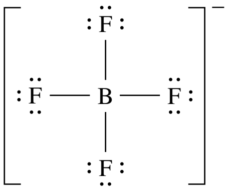
➤ Try yourself 8.2.1

DRAWING LEWIS STRUCTURE FOR POLYATOMIC MOLECULES

Draw the Lewis structure for ammonia (NH_3).

Worked example 8.2.2

DRAWING LEWIS STRUCTURES FOR POLYATOMIC IONS

Draw the Lewis structure for the boron tetrafluoride ion (BF_4^-).	
Thinking	**Working**
Step 1: Sum the valence electrons from all atoms shown in the formula.	B has the electron configuration of 2,3 and so has three valence electrons. F has the electron configuration of 2,7 and so has seven valence electrons. Note, in this case, there is a negatively charged ion with a charge of −1. You therefore need to add one valence electron to the total. Sum of all valence electrons in $BF_4^- = 3 + (4 \times 7) + 1 = 32$
Step 2: Write atomic symbols for all atoms to show which atoms are connected to which. The central atom will be the most electropositive.	In this case, B is the most electropositive and the central atom; the four F atoms will be peripheral atoms. For polyatomic ions, you also need to put square brackets around the structure and the charge of the ion at the top right of the brackets.  A total of eight electrons have been used so far representing the four pairs of bonding electrons in the four single bonds.
Step 3: Complete octets for all peripheral atoms bonded to the central atom.	Do this by adding three pairs of electrons to the atomic symbol of the peripheral atoms.  A total of 32 electrons have been used in total so far; four bonding pairs on the central B atom and 12 non-bonding pairs across the three F atoms.
Step 4: Place any leftover electrons on the central atom. The Lewis structure is complete.	This step is not required since all 32 valence electrons have been used. 
Step 5: If there are not enough electrons to give the central atom an octet, form multiple bonds.	This step is not required since the central atom has a total of eight electrons surrounding it.

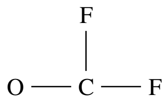
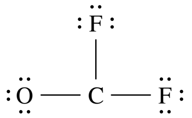
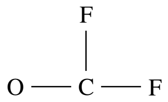
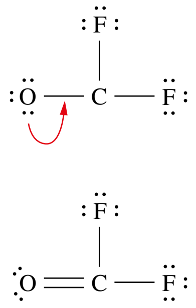
➤ Try yourself 8.2.2

DRAWING LEWIS STRUCTURES FOR POLYATOMIC IONS

Draw the Lewis structure for the ammonium ion (NH_4^+).

Worked example 8.2.3

DRAWING LEWIS STRUCTURES FOR POLYATOMIC MOLECULES WITH MULTIPLE BONDS

Draw the Lewis structure for carbonyl fluoride (COF_2).	
Thinking	**Working**
Step 1: Sum the valence electrons from all atoms shown in the molecular formula COF_2.	C has the electron configuration of 2,4 and so has four valence electrons. O has the electron configuration of 2,6 and so has six valence electrons. F has the electron configuration of 2,7 and so has seven valence electrons. Sum of all valence electrons in $COF_2 = 4 + 6 + (2 \times 7) = 24$
Step 2: Write atomic symbols for all atoms to show which atoms are connected to which. The central atom will be the most electropositive.	In this case, C is the most electropositive and the central atom; the O atom and the two F atoms will be peripheral atoms.  A total six electrons have been used so far representing the three pairs of bonding electrons in the three single bonds.
Step 3: Complete octets for all peripheral atoms bonded to the central atom.	Do this by adding three pairs of electrons to the atomic symbol of the peripheral atoms.  A total of 24 electrons have been used in total so far; three bonding pairs on the central C atom, three non-bonding pairs on the O atom, and six non-bonding pairs across the two F atoms.
Step 4: Place any leftover electrons on the central atom.	There are no left over electrons. All of the 24 available valence electrons have been used. 
Step 5: If there are not enough electrons to give the central atom an octet, form multiple bonds.	The central C atom has only six electrons around it. Therefore you need to form one multiple bond (i.e. a double bond) by using a bonding pair of electrons from one of the peripheral atoms. The atom chosen must be the least electronegative, which, in this case, is the O atom. A non-bonding pair of atoms from the O atom is used to create a double bond between the C and O atoms, giving all atoms in the structure an octet. The Lewis structure is now complete. 

➤ Try yourself 8.2.3

DRAWING LEWIS STRUCTURES FOR POLYATOMIC MOLECULES WITH MULTIPLE BONDS

Draw the Lewis structure for hydrogen cyanide (HCN). (Hint: the H atom can only accommodate two electrons in its outer shell and, therefore, cannot be a central atom.)

SHAPES OF MOLECULES

The shapes of diatomic molecules such as H_2 and HCl are simple. With two atoms only, a diatomic molecule is linear, as you can see in Figure 8.2.14.

FIGURE 8.2.14 H_2 and HCl are diatomic molecules containing only two atoms, so they are linear in shape.

Lewis structures can be used as a starting point in determining the shape of molecules. These diagrams show the outer-shell electron pairs in the molecule. Electron pairs have a negative charge and they repel each other. The electron pairs arrange themselves as far away from each other as possible.

In the case of methane, the Lewis structure (Figure 8.2.15) shows four outer-shell electron pairs around the central carbon atom (represented by the four straight lines).

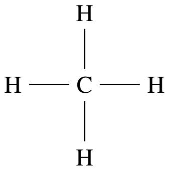


FIGURE 8.2.15 The carbon atom in methane has four pairs of electrons in its outer shell represented by the four straight lines in the Lewis structure. These electron pairs repel each other.

Representations of molecules

A Lewis structure does not show the shape of a molecule but there are several other accepted styles for representing molecules that do. A **ball-and-stick model**, as shown in Figure 8.2.16, displays the shape and the type of bonds (single, double or triple) in a molecule. The atoms are represented by spheres that are connected by rods, representing the bonds. The spheres represent the centre of the atom, not the whole space the atom occupies.

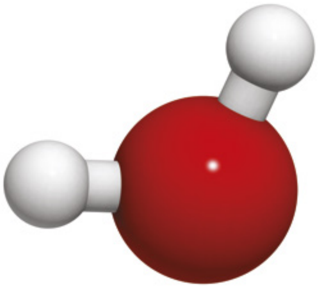


FIGURE 8.2.16 Ball-and-stick model of a water molecule

A **space-filling model** (Figure 8.2.17) also uses spheres but this time the spheres represent the whole atom, including its electron cloud. Different-coloured spheres are used in both types of models to identify the different elements.

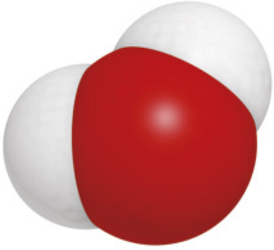


FIGURE 8.2.17 Space-filling model of a water molecule

Repulsion between the electron pairs means that the hydrogen atoms form a **tetrahedral shape** around the carbon atom (Figure 8.2.18). In this molecule, the bonds between the carbon and hydrogen atoms are at an angle of 109° to each other.

You will have a more detailed look at the shapes of molecules and intermolecular bonding in Chapter 12.

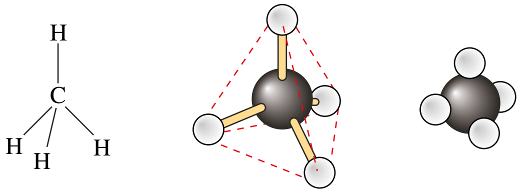
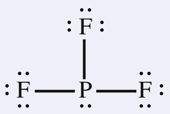
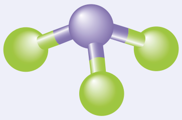
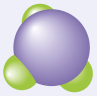


FIGURE 8.2.18 A Lewis structure and two models of a methane molecule show that hydrogen atoms form a tetrahedron around the central carbon.

Each different representation of a molecule has its advantages. A Lewis structure is often drawn before the shape of the molecule is determined and can be used to determine the bonds present, while the ball-and-stick and space filling models show the shape of the molecule and the position of atoms. Table 8.2.2 compares different representations of a PF_3 molecule.

WS 1.1.5

TABLE 8.2.2 Comparison of different representations of phosphorus fluoride, PF_3

	Lewis structure	Ball-and-stick	Space-filling
Diagram			
Model best used for	simplifying the electron dot diagram	displaying the molecule shape	showing the relative size and position of the atoms in the molecule
Limitation of model	does not show the relative size of atoms or shape of the molecule	shows the shape but not the relative sizes of the atoms	shows the relative size and position of the atoms but does not show bond angles or types of bonds

8.2 Review

SUMMARY

- The atoms in non-metallic molecules are held together by covalent bonds.
- A covalent bond involves the sharing of electrons.
- Covalent bonds form between non-metallic atoms, often enabling the atoms to gain outer shells containing eight electrons (except hydrogen which gains an outer shell containing two electrons).
- A single covalent bond forms when two atoms share a pair of electrons.
- A double covalent bond forms when two atoms share four electrons (two pairs).
- A triple covalent bond forms when two atoms share six electrons (three pairs).
- Outer-shell electrons that are not involved in bonding are called lone pairs.
- Lewis structures show the valence electron arrangements of atoms in a molecule and can be used to readily identify the number of bonding pairs and non-bonding pairs of electrons in a molecule.

KEY QUESTIONS

Retrieval

1 Recall the octet rule.

2 Describe a covalent bond.

3 State the difference between a single, double and triple covalent bond.

4 List ethane, ethene and ethyne in order of:
  a increasing carbon–carbon bond length
  b increasing carbon–carbon bond strength

Comprehension

5 Represent the covalent bonds formed between atoms in these diatomic molecules.
  a H_2
  b N_2
  c O_2
  d F_2

6 Determine the maximum number of covalent bonds an atom of each of the following elements can form.
  a F
  b O
  c N
  d C
  e H
  f Ne

7 Represent each of the following molecules using Lewis structures.
  a fluorine (F_2)
  b hydrogen fluoride (HF)
  c water (H_2O)
  d tetrachloromethane (CCl_4)
  e phosphine (PH_3)
  f carbon dioxide (CO_2)

Analysis

8 Identify the number of bonding pairs and non-bonding pairs of electrons surrounding the central atom in the following molecules or ions.
  a PCl_3
  b NH_4^+
  c PH_3
  d H_2S
  e CCl_4

9 Identify the type of representation of a molecule that best fits the following descriptions.
  a shows the shape of the molecule and type of atoms present
  b shows the order of attachment of atoms in a molecule and arrangement of bonding electrons between atoms as straight lines
  c shows the relative size and position of atoms and the shape of molecule

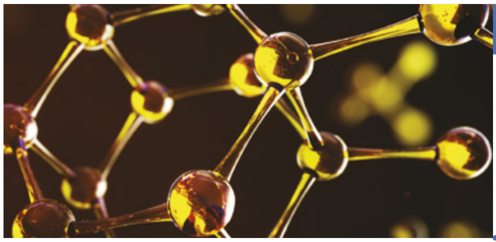
10 Predict the most likely formula of the compound formed between the following pairs of elements.
  a C, Cl
  b N, Br
  c Si, O
  d H, F
  e P, F

11 Atom X has an electron configuration of 2,6. Atom Y has a configuration of 2,7. Deduce the likely molecular formula of the compound that will form between X and Y.

12 Compare and contrast the bonding of the oxygen atom in water (H_2O) and in methanal (CH_2O).

13 Differentiate between the non-metallic elements argon (Ar) and chlorine (Cl).

8.3 Carbon lattices and carbon nanomaterials



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- identify the different allotropes of carbon
- understand that the different bonding structures in different allotropes of carbon result in different physical and chemical properties
- recognise the potential applications of different carbon nanomaterials.

You saw in Chapter 2 that covalent substances can exist as small discrete molecules or as large networks (or lattices) of non-metallic atoms bonded together. Diamonds are highly valued in our society. **Diamonds** are a form of pure carbon, but they are not the only form that carbon can take. Graphite, charcoal, graphene and **fullerenes** are also made entirely from carbon but their properties are very different from those of diamond. Scientists have found many exciting uses for these other forms of carbon. As the properties of these forms of carbon are quite different, it is not surprising that their chemical structures are also different.

In this module, you will look more closely at non-metallic networks and in particular at different forms of carbon lattices and carbon nanomaterials.

ABUNDANCE AND PROPERTIES OF CARBON

Carbon is a fascinating element for many reasons.

- It is a vital component of all living systems.
- It is the eleventh most abundant element in the universe.
- It has three isotopes: ^{12}C (98.9% abundant), ^{13}C (1.1% abundant) and ^{14}C (traces).
- It undergoes **sublimation** at temperatures above 3550°C.
- It is a non-metal, but a number of forms can conduct electricity.
- It can form single, double and triple covalent bonds with several other elements.
- It can form large molecules and lattice structures by bonding to itself.

ALLOTROPES

Some elements can exist with their atoms in several different structural arrangements, called **allotropes**, that give them different physical forms. In different allotropes, the atoms are bonded to each other in different, specific ways. This gives them significantly different properties from other allotropes of the same element. Oxygen and carbon are two of the most common elements that have different allotropic forms.

Oxygen has two allotropes. Oxygen gas consists of diatomic molecules with the formula O_2. Each oxygen atom in this arrangement is bound to one other oxygen atom. Ozone is another molecule containing only oxygen. Ozone molecules have the formula O_3 and consist of a central oxygen atom bound to two other oxygen atoms. Figure 8.3.1 shows the structure of these two molecules. As both contain only oxygen atoms, they are both allotropes of oxygen. The rest of this module will focus on the different allotropes of carbon.

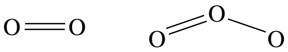


FIGURE 8.3.1 Oxygen and ozone are two molecules that contain only oxygen atoms.

Allotropes are different forms of the same element.

Diamonds (Figure 8.3.2) might be 'a girl's best friend' but it is unlikely that graphite (Figure 8.3.3) will ever be held in the same esteem. Both of these minerals are made of the same single element—carbon. Graphene and fullerenes are new materials that are also allotropes of carbon.

Table 8.3.1 summarises some information about the structure, properties and uses of the two most common allotropes of carbon: diamond and graphite.

TABLE 8.3.1 Comparison of properties of some of the allotropes of carbon

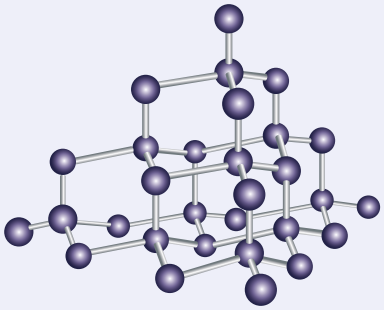
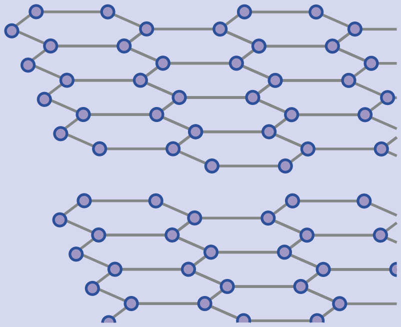
Allotrope	Structure	Properties	Uses
Diamond	 covalent network lattice, each carbon surrounded by four other carbon atoms in a tetrahedral arrangement	- very hard - sublimes - non-conductive - brittle	- jewellery - cutting tools - drills
Graphite	 covalent layer lattice, each carbon bonded to three other carbons, one delocalised electron per carbon atom	- conductive - slippery - soft - greasy material	- lubricant - pencils - electrodes - reinforcing fibres

FIGURE 8.3.2 Diamond is the hardest naturally occurring substance.

FIGURE 8.3.3 Natural graphite is soft and black.

Diamond is made up of carbon atoms that bond with four neighbouring carbon atoms forming a covalent network lattice. This structure makes diamond extremely hard.

Diamond

Diamond is the hardest naturally occurring substance known.

Diamond does not contain small, discrete (individual) molecules. Instead, the carbon atoms bond to each other to form a continuous three-dimensional structure called a **covalent network lattice**. There are no weak intermolecular forces present, only strong covalent bonds. This is what gives diamond its strength.

In general, substances that have a network lattice structure have very high melting points or decomposition temperatures. They are also very hard because the atoms are held firmly in fixed positions in the lattice.

As you will see in Chapter 12, when an atom has four electron pairs in its outer shell, the electron pairs position themselves as far away from each other as possible in a tetrahedral shape. In the covalent network lattice for diamond shown in Figure 8.3.4, you can also see that individual atoms within diamond form single covalent bonds to four other carbon atoms in a tetrahedral arrangement.

The structure of diamond is directly related to its properties.

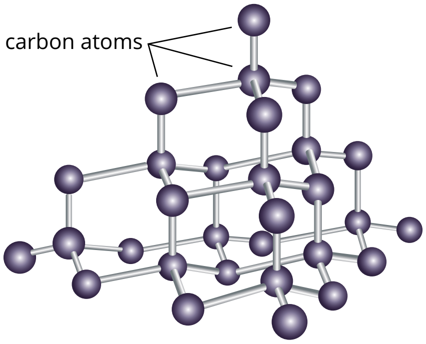


FIGURE 8.3.4 The structure of diamond showing each carbon atom with four single covalent bonds to neighbouring atoms

FIGURE 8.3.5 Diamond-tipped drills are used to drill through rock in the fracking industry.

- Single covalent bonds between carbon atoms are strong bonds. The entire structure of a diamond consists of a continuous network of these bonds, making diamond very hard and rigid.
- There are no small molecules in diamond, so there are no weak forces between the atoms. There are only strong covalent bonds between carbon atoms and this makes the sublimation point very high (about 3500°C).
- The rigidity means that diamonds are brittle and break rather than bend.
- Diamond does not conduct electricity because it does not contain any charged particles that are free to move.
- Because the atoms in diamonds are held together very strongly, the thermal conductivity is extremely high. It is five times greater than that of copper, leading to some specialty electronic uses where diamond is used to transfer heat away from some important electrical components.

The crystalline appearance of diamonds and their high refractive index make them sparkle and have made them extremely popular as jewellery, but the hardness of diamond also lends itself to industrial uses. Many industrial cutting and drilling tools for working with tough materials are diamond tipped. The drill tips in Figure 8.3.5 are used to drill through rock in the fracking industry. They contain small pieces of diamond that improve the hardness and durability of the tool.

Graphite

Graphite is a very different form of carbon. As you can see in Figure 8.3.6, the carbon atoms in graphite are in layers. There are strong covalent bonds between the carbon atoms in each layer. However, there are weak intermolecular forces between the layers. As a consequence, it is hard in one direction but quite slippery and soft in another direction. The structure of graphite is referred to as a **covalent layer lattice**.

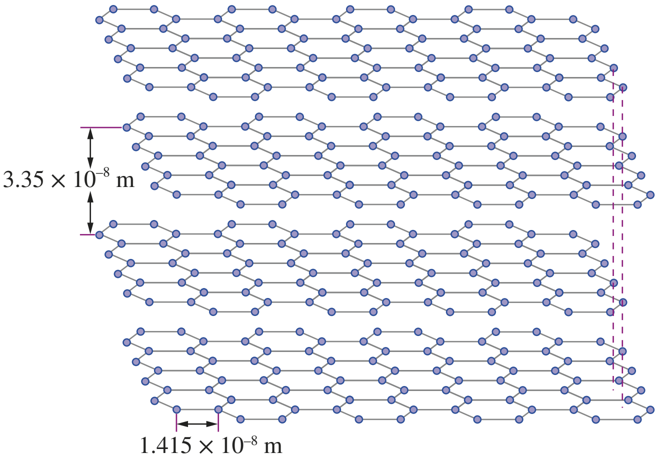


FIGURE 8.3.6 Graphite has a covalent layer lattice structure. The carbon atoms within each layer are covalently bonded to each other. Weak dispersion forces exist between the layers.

ℹ In graphite, each carbon atom is covalently bonded to three other carbon atoms. The layered network structure contains delocalised electrons. Bonds within the layers are strong but bonds between layers are weak dispersion forces.

The covalent layer lattice structure of graphite also explains some of its other properties.

- The strong covalent bonds between the atoms in each layer explain graphite's resistance to heat. Graphite sublimes at a temperature of about 3600°C.
- Each carbon atom is bonded to three other carbon atoms. The fourth valence electron from each atom is able to move within the layer. The electrical conductivity of graphite is due to these delocalised electrons.

The conductivity of graphite makes it suitable for applications such as battery electrodes where conductivity is required but a metal is not suitable.

Graphite can also be used as a lubricant. The weak intermolecular forces between layers allow these layers to slide over each other and to reduce the friction between moving parts, such as in locks or machinery.

Graphite is also used as an additive to improve the properties of rubber products and it can be woven into a fibre. This helps to reinforce plastics. Figure 8.3.7 shows spun graphite fibre, which can be used to make strong composite materials such as those used in tennis racquets, fishing rods and racing car shells.

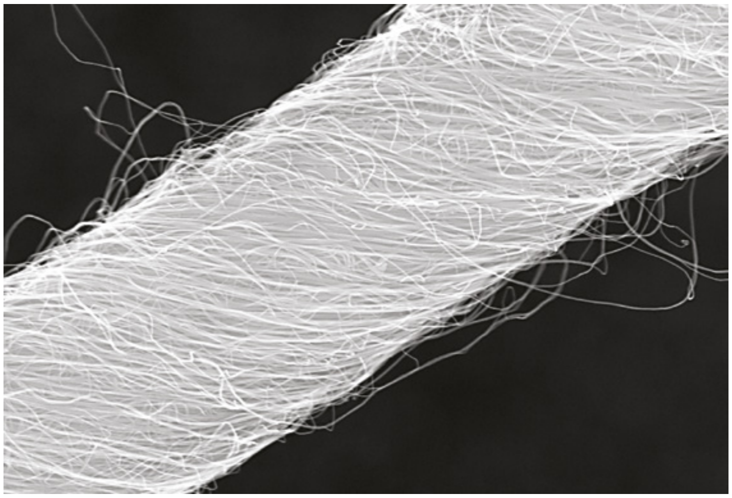


FIGURE 8.3.7 Graphite fibre can be used to reinforce plastics.

CARBON NANOMATERIALS

Imagine an elevator with a difference, an elevator into space. Rockets are costly and dangerous. Why not take an elevator ride instead? While this concept might sound like science fiction, scientists are considering this idea very seriously. Japanese company Obayashi has announced that they will have a space elevator up and running by the year 2050. The elevator would reach 96 000 km into space and use robotic cars powered by magnetic linear motors. The concept is only possible because of the properties of a recently discovered allotrope of carbon–carbon **nanotubes**. Figure 8.3.8 is an artist's impression of a space elevator, made of a carbon nanotube.

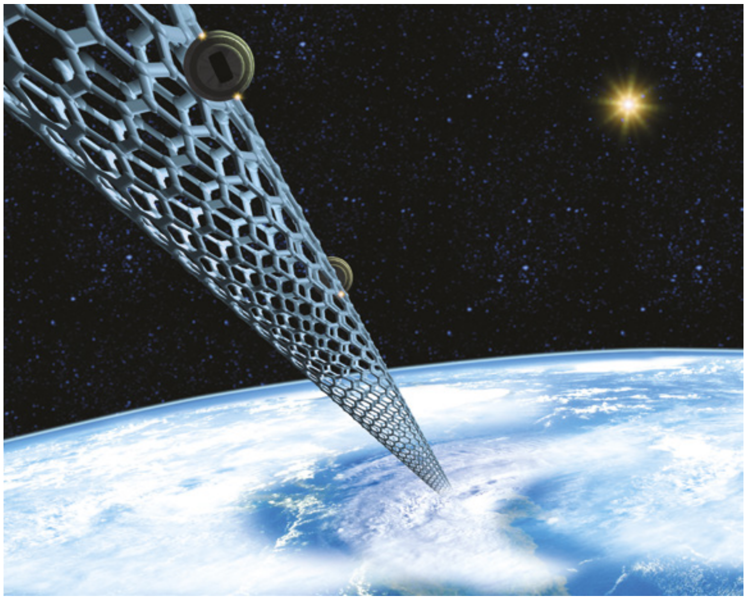


FIGURE 8.3.8 This is an artist's impression of a space elevator made of a carbon nanotube.

Diamond and graphite have long been recognised as allotropes of carbon. However, since the 1970s, scientists have discovered how to make a new range of carbon allotropes that are examples of **nanomaterials**. Nanomaterials are particularly interesting because they have a very high surface area to volume ratio, leading to some unique or enhanced properties.

Fullerenes

Fullerenes are an allotrope of carbon where the atoms are arranged in a series of pentagons and hexagons.

In the late 1970s, while working at the Australian National University in Canberra, Dr Bill Burch discovered a new allotrope of carbon. This allotrope was made up of molecules containing a roughly spherical group of carbon atoms arranged in a series of pentagons and hexagons, similar to the shape of a soccer ball, as you can see in Figure 8.3.9.

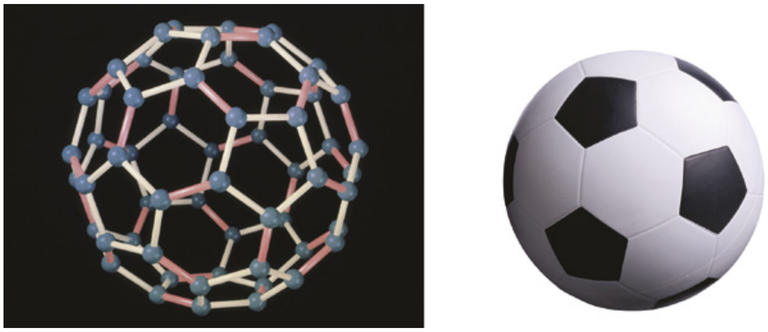


FIGURE 8.3.9 The structure of a fullerene has a similar pattern to the surface of some soccer balls.

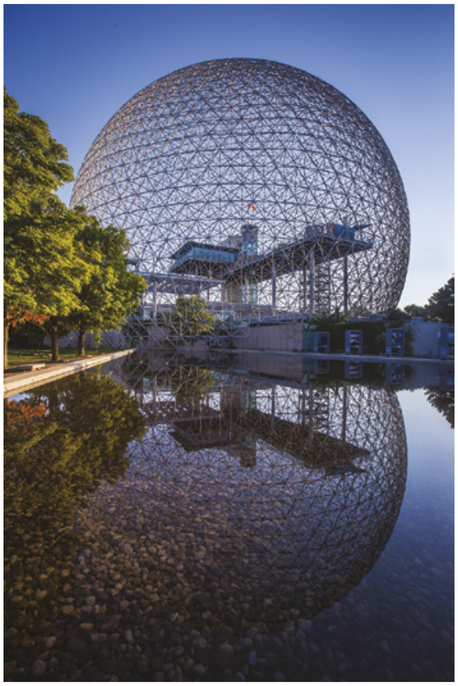


FIGURE 8.3.10 The Biosphere in Montreal, Canada, is a museum designed by Buckminster Fuller.

Scientists have since found further variations of this molecule. These molecules have similar structures to the geodesic designs of architect Richard Buckminster 'Bucky' Fuller. They are called fullerenes, although they are more commonly referred to as **buckyballs**. Figure 8.3.10 shows the Biosphere in Montreal, Canada—a museum designed by Buckminster Fuller.

Fullerenes have three covalent bonds to each carbon atom and in some ways appear to be similar to graphite. This leaves delocalised electrons in the structure and the possibility of electrical conductivity. Although fullerenes were initially just a curiosity, scientists predict that they have significant potential in a number of fields such as composite materials and **photovoltaic cells** (solar panels). The most stable fullerene molecule involves 60 carbon atoms bonded into an approximately spherical shape that is known as buckminsterfullerene or C_{60}.

Graphene

FIGURE 8.3.11 Graphene is a single-layer sheet with the same arrangement as those stacked in graphite.

You have seen that graphite has a layered structure. **Graphene** is best described as a single layer of graphite (Figure 8.3.11). Graphene is a single-layer sheet with the same arrangement as those stacked in graphite. It is a very new material and was first isolated in 2004.

Graphite is soft, due to the weak intermolecular forces between its layers. Graphene is only a single layer and retains the electrical conductivity of graphite but it is an extremely strong and tough material.

Graphene has many potential uses. Graphene could:

- replace silicon as the basis for computer chips and circuits due to its high electrical conductivity
- be used in desalination plants. Water under pressure can pass through the thin layer but dissolved impurities cannot
- be used to construct electrodes where it is an advantage for an electrode to be a non-metal
- be used in organic photovoltaic cells
- be used to reinforce composite materials because of its strength.

An interesting feature of graphene is that, because it is a single layer, every carbon atom is available for reaction from two sides at any instant during a chemical reaction.

Carbon nanotubes

'Nanotubes' are closely related to graphene. They are called nanotubes because they have a long, hollow structure with walls formed from graphene. The diameter of these cylinders is very small, around 1 nanometre (10^{-9} metre) wide, while they can be millions of times longer. They can be capped on the end of each cylinder by a half fullerene molecule as shown in Figure 8.3.12.

Nanotubes can be single-walled or multi-walled. A multi-walled nanotube has smaller tubes sitting inside larger tubes.

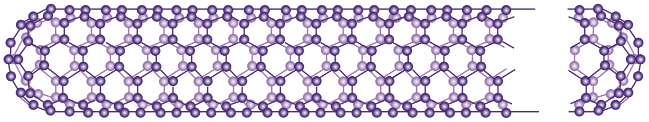


FIGURE 8.3.12 A carbon nanotube can be regarded as a sheet of graphene rolled into a cylinder and capped on the ends by half a fullerene molecule.

Scientists are interested in nanotubes because of their:

- unique strength
- electrical conductivity (the conductivity of nanotubes depends on their shape; some are conductors and other are semiconductors)
- thermal conductivity
- strong forces of attraction to each other.

Nanotubes hold great promise in fields such as optics, nanotechnology and electronics. Their extraordinary strength and thermal and electrical conductivity suggest they may be useful as additives in various structural materials.

POTENTIAL OF CARBON NANOMATERIALS

Carbon nanomaterials offer huge gains in performance and properties over some other materials in current use and have a broad range of potential applications. The carbon–carbon bonds in these structures are very strong and there are no weak points in a single layer of graphene or a nanotube.

Carbon nanotubes are:

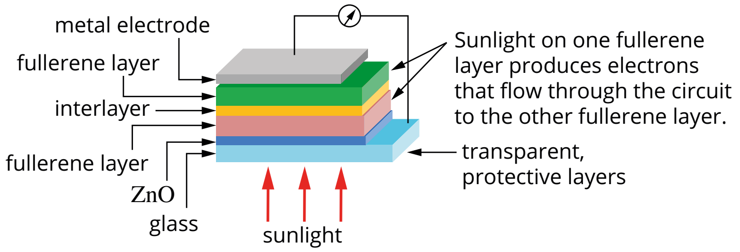
- up to 300 times stronger than steel. Rope made from nanotubes with a diameter of 1 cm could support a weight of over 1000 tonnes. Nanotubes are already being used in high-performance sporting equipment
- better conductors of electricity than silver. Since nanotubes are essentially 'wires' that are much narrower in diameter than metal wire, they offer the possibility for extreme miniaturisation of electrical circuits
- better thermal conductors than diamond. Nanotubes could be used to transfer heat away from electrical components
- stronger than Kevlar fibres. Stain-resistant nanofabrics that never require washing are already available. You could even carry water in the pockets of a vest made from this material
- capable of adsorbing more gas or impurities than activated charcoal.

Perhaps the best way to highlight the potential of nanomaterials is with three exciting examples.

Example 1: Flexible photovoltaic cells

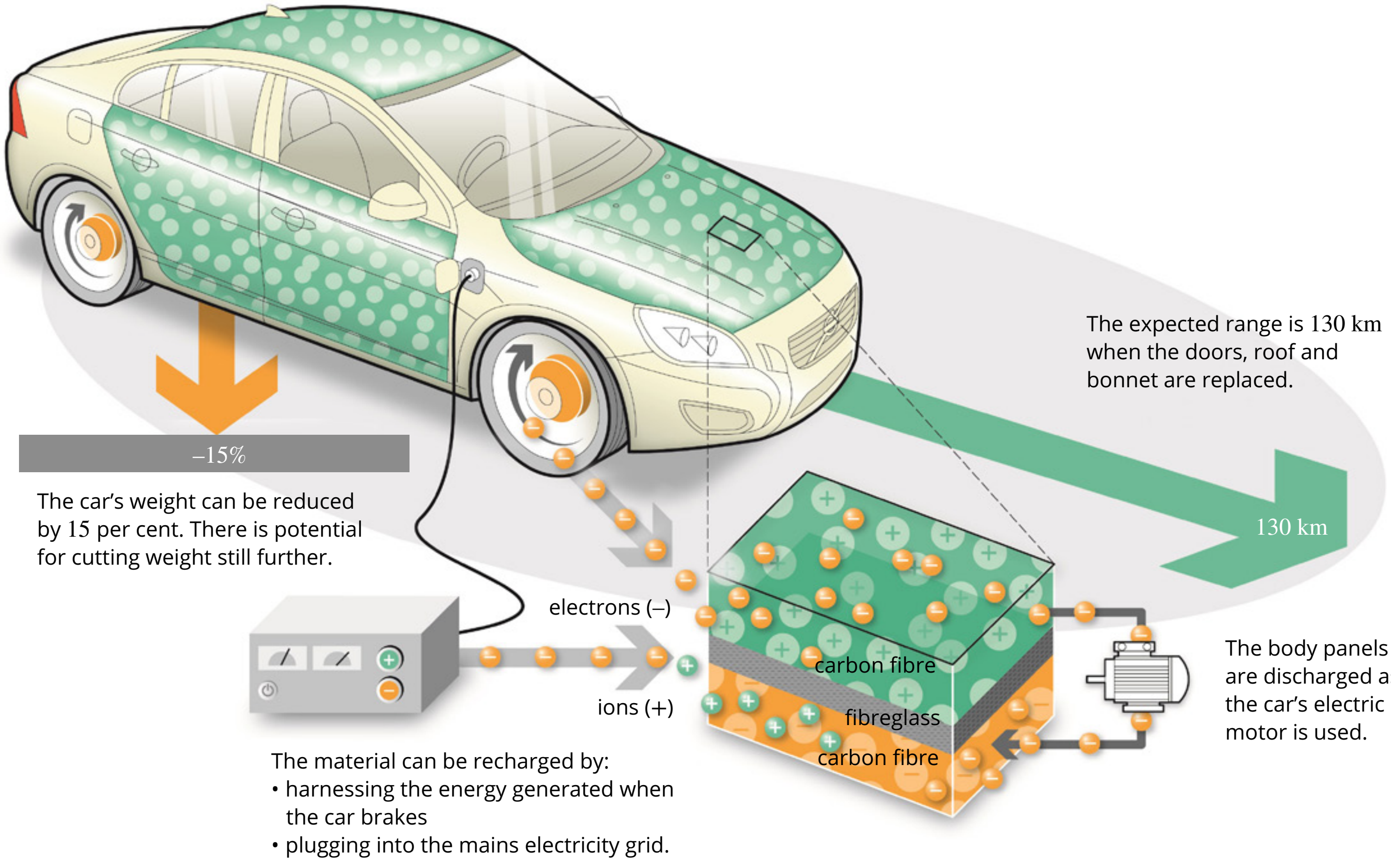
Traditional photovoltaic cells (the solar panels you see on many rooftops) are made from highly refined and purified silicon crystals. Manufacturing these crystals is complex and the cells produced are rigid and brittle.

Research into fullerenes has led to alternative types of photovoltaic cells, known as polymer solar cells. Polymer solar cells (Figure 8.3.13) use alternate layers of fullerene molecules instead of silicon. The cell produced is lighter than conventional cells and offers the advantage of being flexible. A flexible cell could match the curved shape of a caravan roof or the cabin of a boat. The initial problem with these cells was their low efficiency. Research into fullerenes and other aspects of the cells continues to improve their efficiency.



Sunlight on one fullerene layer produces electrons that flow through the circuit to the other fullerene layer.

FIGURE 8.3.13 Polymer solar cells are constructed from transparent covering layers and alternating layers of fullerenes.



The car’s weight can be reduced by 15 per cent. There is potential for cutting weight still further.

The expected range is 130 km when the doors, roof and bonnet are replaced.

The body panels are discharged as the car’s electric motor is used.

The material can be recharged by:

- harnessing the energy generated when the car brakes
- plugging into the mains electricity grid.

FIGURE 8.3.14 The Volvo concept car is made from the latest nanomaterials, which replace most of the steel in the car. The nanomaterials are made of extremely thin and strong carbon fibre and can be used in the car’s roof, doors, bonnet and floor and can also serve as the car’s battery.

Example 2: Volvo concept car

Volvo has embarked on a radical new design for an electric car that aims to harness the properties of nanomaterials. As shown in Figure 8.3.14, most of the steel in the car has been replaced with carbon nanotube sheets. These sheets have the advantage of lightweight strength but they can also serve as a giant battery for the car. Fullerenes are incorporated in the carbon sheets, allowing them to act as photovoltaic cells, supplying the energy needed to recharge and power the car.

Example 3: Solar aircraft

In March of 2015, the Swiss-designed plane *Solar Impulse-2* (Figure 8.3.15) set off from Abu Dhabi on a multi-stage journey to become the first solar-powered aeroplane to circumnavigate the globe. The plane is powered by solar cells and the lightweight strength for the structure comes from an assortment of carbon nanomaterials and composites. Over a 16-month period (which included time-off for critical repairs) the *Solar Impulse-2* completed the 42 000 km journey to become the first fixed-wing piloted aircraft to circumnavigate the Earth using only solar power.

FIGURE 8.3.15 The *Solar Impulse-2*

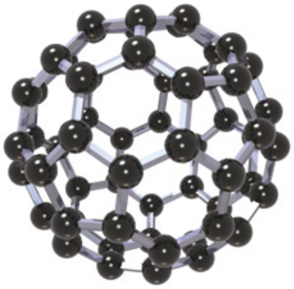
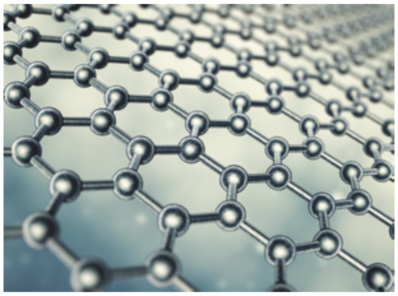
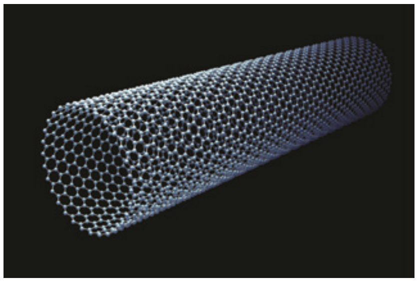
PA 1.2.2 WS 1.2.5 WS 1.2.9

8.3 Review

SUMMARY

- Carbon can be found in the Earth's crust in the form of diamond, graphite or charcoal. The structures and properties of these allotropes are very different.
- In diamond, each carbon atom is covalently bonded to another four carbon atoms in a tetrahedral shape, forming a covalent network lattice structure. Diamond sublimes at a high temperature, is extremely hard and has a sparkling, crystalline appearance.
- In graphite, each carbon atom is covalently bonded to three other carbon atoms. The layered network structure contains delocalised electrons. Bonds within the layers are strong but bonds between layers are weak dispersion forces. Graphite is slippery, conducts electricity and sublimes at a high temperature.
- Over the past 40 years, scientists have developed a range of new carbon allotropes called fullerenes.
- Fullerenes are examples of nanomaterials and are of interest to scientists because of the enhanced properties that their high surface area to volume ratio offers. Their high tensile strength and high electrical and heat conductivity are of particular interest.
- Spherical fullerenes are known as buckyballs. Fullerenes can also be tubular, as in nanotubes. Graphene is a single-layered form of graphite. In all these allotropic forms, the carbon atoms are bonded to three other carbon atoms as shown in Table 8.3.2.
- Potential applications of fullerenes include fibres and fabrics, electrical circuits, photovoltaic cells and filtration systems.

TABLE 8.3.2 Allotropic forms of carbon

Buckyball	Graphene	Nanotube
		

KEY QUESTIONS

Retrieval

1 State what is unusual about the electrical conductivity of some forms of carbon given that carbon is a non-metal.

Comprehension

2 Explain why carbon can form so many different compounds.

3 **a** Explain what is meant by the word 'sublime'.
  b Explain why diamond and graphite only sublime at temperatures over 3500°C.

4 Explain the following properties of diamond and graphite in terms of their respective structures.
  a hardness or softness
  b ability or inability to conduct electricity

5 Explain the following in terms of the structures of graphite and diamond.
  a Graphite is used as a lubricant.
  b Diamond is often used as an edge on saws and a tip on drills.

6 Describe the bonding within a C_{60} buckyball.

7 Describe the bonding within a graphene sheet.

Analysis

8 Compare and contrast the structures and properties of fullerenes and graphite.

9 Differentiate between graphene and graphite.

Chapter review

WS 1.2.6 WS 1.2.7

08

KEY TERMS

allotrope
ball-and-stick model
buckyball
covalent bond
covalent layer lattice
covalent network lattice
diamond
diatomic molecule
double covalent bond
electron dot diagram
fullerene
graphene
graphite
intermolecular bond
intramolecular bond
Lewis structure
lone pair
molecular compound
molecular formula
molecule
nanomaterial
nanotube
non-bonding electron
octet rule
photovoltaic cell
polyatomic molecule
single covalent bond
space-filling model
sublimation
tetrahedral shape
triple covalent bond

KEY QUESTIONS

Retrieval

1 Identify the statement that best describes the way hydrogen atoms bond to each other.
- **A** One hydrogen atom donates an electron to another hydrogen atom to form a molecule.
- **B** Hydrogen atoms form a lattice with delocalised electrons.
- **C** Hydrogen atoms share electrons to obtain a complete outer shell of eight electrons.
- **D** Two hydrogen atoms share an electron each to form a hydrogen molecule.

2 State how many valence electrons each of the following atoms has.
- **a** C
- **b** Se
- **c** Al
- **d** I
- **e** Si

3 Show the electron dot diagrams for the following atoms.
- **a** nitrogen
- **b** oxygen
- **c** hydrogen
- **d** chlorine
- **e** fluorine

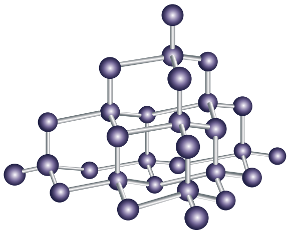
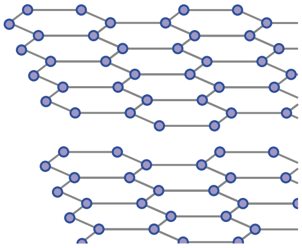
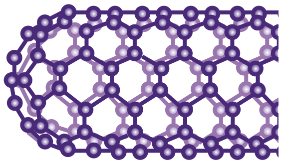
4 Show the formation of the following elemental molecules using electron dot diagrams for atoms and Lewis structures for the molecules formed.
- **a** the formation of the H_2 molecule from two H atoms
- **b** the formation of the O_2 molecule from two O atoms
- **c** the formation of the N_2 molecule from two N atoms

5 Show the formation of the following molecular compounds using electron dot diagrams for atoms and Lewis structures for the molecules formed.
- **a** the formation of the HCl molecule from an H atom and a Cl atom
- **b** the formation of the H_2O molecule from two H atoms and one O atom
- **c** the formation of the CH_4 molecule from one C atom and four H atoms

6 Indicate whether the following statements about carbon dioxide (CO_2) are true or false.
- **a** Carbon dioxide is a molecular compound.
- **b** A molecule of carbon dioxide contains three atoms.
- **c** There are 99 oxygen atoms in 33 molecules of carbon dioxide.
- **d** The bonds between the carbon and oxygen atoms in carbon dioxide are intramolecular bonds.
- **e** There are two single covalent bonds in a molecule of carbon dioxide.
- **f** There are four lone pairs of electrons in a molecule of carbon dioxide.

7 Identify the nature of the carbon–carbon bonding in diamond from the following list.
- **A** The carbon–carbon bond is a double covalent bond.
- **B** The carbon–carbon bond is a triple covalent bond.
- **C** A carbon atom is covalently bonded to four other carbon atoms to form the C_5 molecule.
- **D** The carbon–carbon bond is a single covalent bond that forms the C_2 molecule.
- **E** Each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral arrangement, forming a giant network lattice structure.

8 State the allotrope name, bonding feature and properties of the structures in the table below.

				
Allotrope name				
Bonding feature				
Properties				

Comprehension

9 Non-metallic substances generally do not contain free-moving charged particles. Explain what this indicates about the general physical properties of non-metallic substances.

10 Solid carbon dioxide sublimes when it is heated; that is, it goes straight from a solid to a gas. Explain what happens to the bonds in carbon dioxide when it sublimes. Use the terms 'intermolecular bonds' and 'intramolecular bonds' in your answer.

11 In a molecule of nitrogen (N_2), determine how many of the following there are.

- **a** bonding electrons
- **b** covalent bonds
- **c** non-bonding electrons

12 Explain why neon atoms do not form covalent bonds.

13 Draw Lewis structures for each of the following molecules and identify the number of bonding and non-bonding electrons in each molecule.

- **a** HBr
- **b** H_2O
- **c** CF_4
- **d** C_2H_6
- **e** PF_3
- **f** CH_4
- **g** H_2S

14 Describe the geometry of the bonds around carbon atoms in diamond and graphene.

15 Describe the structure of a carbon nanotube.

16 Draw the Lewis structures for the NO^+ ion and the NO_2^- ion, and state the number of bonding and non-bonding pairs of electrons in the structure.

17 Elements X and Y are represented by the following electron dot diagrams:

- **a** If X and Y were to form a molecular compound, determine what the formula would be.
- **b** The compound formed is one of three possible real compounds; name all three.

18 Draw the electron dot diagrams for nitrogen (N) and fluorine (F). Use the information from the electron dot symbols to predict the chemical formula for the molecular compound that forms between N and F. Draw the Lewis structure of the compound to confirm your prediction.

19 Draw the Lewis structures for PCl_5, SF_6 and IF_3. Explain what is unusual about these structures.

Analysis

20 Hydrogen chloride (HCl) exists as a gas at room temperature. Determine what conclusions you can draw about the strength of the intermolecular bonding in pure hydrogen chloride.

21 The formula of a molecule is XY_4. Identify the alternative that could match this formula.

- **A** OH_4
- **B** CH_4
- **C** HBr_4
- **D** CO_4

22 Determine what valence electrons are. If an atom has the electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^4$, determine how many valence electrons it has. Explain your answer.

23 Compare and contrast the electron dot diagram for neon with the Lewis structure for methane. Identify the ways in which the electron arrangements around neon and carbon are similar and in what ways are they different.

24 A substance XY is formed from two different elements. The substance has a boiling point of –44°C. Determine if XY is likely to be an ionic compound or a molecular compound. Explain your answer.

25 Determine how many lone pairs of electrons are in each of the following molecules.

a H_2
b NH_3
c HCl
d O_2

26 Deduce the number of bonding pairs and non-bonding pairs of electrons around the central atom in the following polyatomic molecules or ions.

a NF_3
b BF_4^-
c COF_2

27 Analyse the main differences between drawing Lewis structures for polyatomic ions compared with polyatomic molecules.

28 There are three possible Lewis structures for the nitrate ion (NO_3^-). Compare and contrast drawings of all three structures.

29 The atoms in molecules of nitrogen (N_2), oxygen (O_2) and fluorine (F_2) are held together by covalent bonds. Distinguish how the bonds in these molecules are:

a similar
b different.

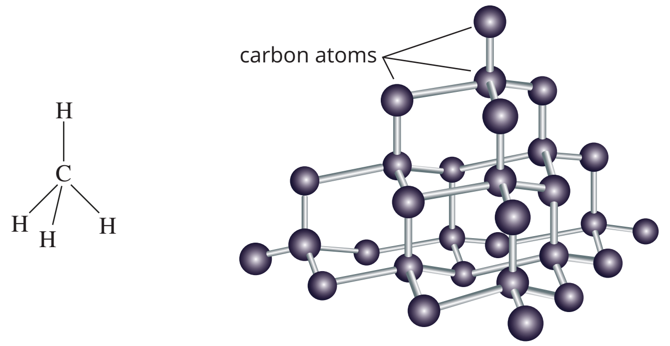
30 Categorise the following substances as metallic, ionic or molecular.

a $CuCl_2$
b Ag
c HCl
d H_2O
e Cu
f CaS
g NH_3

31 Compare and contrast the bonding in ethane, ethene and ethyne.

32 Compare and contrast the carbon–carbon bonding in ethene and benzene.

33 The structures of methane and diamond are shown in the following image. Each carbon atom in methane (CH_4) has a tetrahedral arrangement of atoms around it. A carbon atom in diamond also has a tetrahedral arrangement. However, the two substances have very different properties.

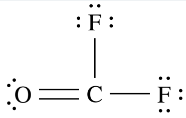


a Identify all of the types of bonding that would be present in each substance.
b Predict the different properties you would expect each to have based on the types of bonding present in each substance.

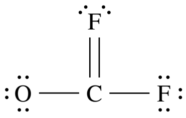
Knowledge utilisation

34 Decide which of the following structures of carbonyl fluoride (COF_2) is the most appropriate and justify your decision.

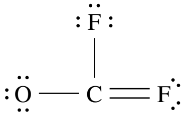
A



B

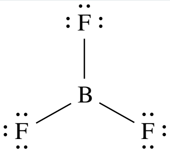


C

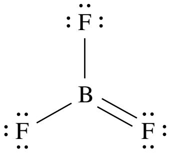


35 Evaluate which of the following is the better structure for boron trichloride (BF_3) and provide reasons for your choice.

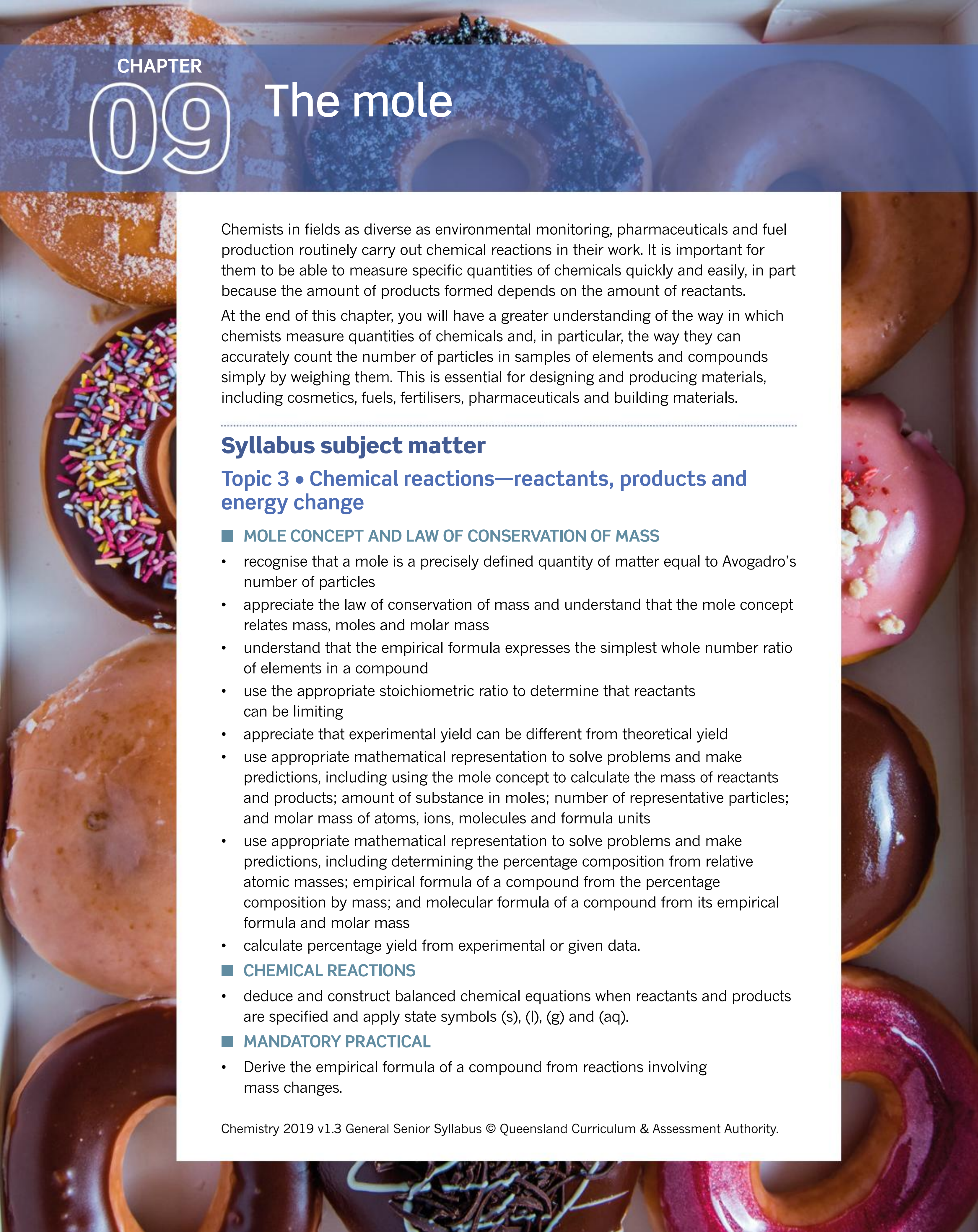
A



B



36 Silicon lies under carbon in the periodic table and may be expected to have similar chemical and physical properties. The compound carbon dioxide (CO_2) is a molecular compound that is a colourless gas at room temperature with a sublimation point of –78.5°C. The compound silicon dioxide (SiO_2) is a crystalline solid at room temperature with a melting point of 1710°C and a boiling point of 2230°C. Determine how you would reconcile these differences in properties between CO_2 and SiO_2.



CHAPTER

09 The mole

Chemists in fields as diverse as environmental monitoring, pharmaceuticals and fuel production routinely carry out chemical reactions in their work. It is important for them to be able to measure specific quantities of chemicals quickly and easily, in part because the amount of products formed depends on the amount of reactants.

At the end of this chapter, you will have a greater understanding of the way in which chemists measure quantities of chemicals and, in particular, the way they can accurately count the number of particles in samples of elements and compounds simply by weighing them. This is essential for designing and producing materials, including cosmetics, fuels, fertilisers, pharmaceuticals and building materials.

Syllabus subject matter

Topic 3 • Chemical reactions—reactants, products and energy change

MOLE CONCEPT AND LAW OF CONSERVATION OF MASS

- recognise that a mole is a precisely defined quantity of matter equal to Avogadro’s number of particles
- appreciate the law of conservation of mass and understand that the mole concept relates mass, moles and molar mass
- understand that the empirical formula expresses the simplest whole number ratio of elements in a compound
- use the appropriate stoichiometric ratio to determine that reactants can be limiting
- appreciate that experimental yield can be different from theoretical yield
- use appropriate mathematical representation to solve problems and make predictions, including using the mole concept to calculate the mass of reactants and products; amount of substance in moles; number of representative particles; and molar mass of atoms, ions, molecules and formula units
- use appropriate mathematical representation to solve problems and make predictions, including determining the percentage composition from relative atomic masses; empirical formula of a compound from the percentage composition by mass; and molecular formula of a compound from its empirical formula and molar mass
- calculate percentage yield from experimental or given data.

CHEMICAL REACTIONS

- deduce and construct balanced chemical equations when reactants and products are specified and apply state symbols (s), (l), (g) and (aq).

MANDATORY PRACTICAL

- Derive the empirical formula of a compound from reactions involving mass changes.

9.1 The mass of particles

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the concepts of atomic mass and relative atomic mass
- explain how 1 atomic mass unit is equal to $\frac{1}{12}$ the mass of a carbon-12 atom
- calculate relative atomic mass
- describe the concepts of relative molecular mass and relative formula mass
- calculate relative molecular mass and relative formula mass.

Atoms are particles that are so small, with such little mass, that they are impossible to count individually. For example, just one granule of table sugar can contain as many as 4.9×10^{19} atoms. That is 49 000 000 000 000 000 000 atoms!

The chemical name for table sugar is sucrose (Figure 9.1.1a). Sucrose is a type of sugar that is extracted and refined from sugarcane, which is grown in many areas of Queensland. Figure 9.1.1b shows a model of a sucrose molecule. One sucrose molecule contains 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms bonded together, and has a total mass of 5.7×10^{-22} g. This means that the 4 grams of sucrose crystals shown on the teaspoon in Figure 9.1.1a contains approximately 7×10^{21} sucrose molecules.

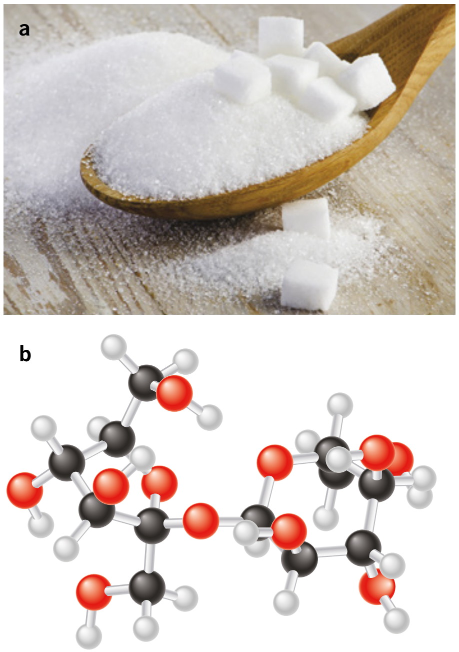


FIGURE 9.1.1 (a) A teaspoon of sucrose crystals contains an incredibly large number of extremely small sucrose molecules. (b) The single sucrose molecule pictured is made of 12 carbon atoms (black), 22 hydrogen atoms (white) and 11 oxygen atoms (red).

PROPERTIES OF ACIDS AND BASES

Such small masses are not easily measured and can be inconvenient to use in calculations. This module will introduce you to the ways scientists determine and use the masses of different particles.

ATOMIC MASSES

Chemistry is a quantitative science, so knowing the mass of atoms is essential. Due to the extremely small size of atoms, it is impossible to weigh an atom individually. What scientists can do is use **mass spectrometry** to compare the relative masses of atoms. Mass spectrometry is described in Chapter 5, beginning at page 104.

The first step in determining the relative mass of atoms is to give a value to the mass of a reference atom. This atom can then be used as a standard to compare the rest. Prior to 1961, oxygen was used as the standard. Unfortunately, physicists and chemists could not agree on a way of assigning a standard mass to oxygen. Chemists assigned a mass of exactly 16 to the average mass of oxygen atoms. Physicists assigned a mass of exactly 16 to the oxygen-16 isotope. This resulted in two different tables of slightly different atomic masses. In 1961, by international agreement, carbon-12 (or ^{12}C) was chosen and given an atomic mass of exactly 12. **Atomic mass** is therefore defined as the mass of an atom in **atomic mass units (amu)**. ^{12}C is the isotope of carbon that has exactly 6 protons and 6 neutrons and an amu of 12. Therefore, one atomic mass unit is defined as a mass exactly equal to $\frac{1}{12}$ the mass of a ^{12}C atom.

Giving ^{12}C an atomic mass of 12 amu then allows other atoms to be given a value in amu. When an atom of hydrogen is experimentally compared to ^{12}C using a mass spectrometer, it is found on average to be 8.400% the size. If hydrogen is 8.400% the size, it would have an atomic mass of 1.008 (0.08400×12 amu = 1.008 amu).

Atomic mass is the mass of the atom in atomic mass units. One atomic mass unit (amu) is defined as a mass exactly equal to $\frac{1}{12}$ the mass of a ^{12}C atom.

RELATIVE ATOMIC MASS

In Chapter 5, you looked at how mass spectra could identify different isotopes of an element and determine the isotopic abundance to calculate the **relative atomic mass** (A_r) of an element. This showed that most elements consist of a mixture of isotopes.

Isotopes, as you have seen, have the same number of protons but different numbers of neutrons. The atomic mass of an atom is mainly determined by its number of protons and neutrons. Since the masses of a proton and a neutron are similar to each other and close to 1 amu on the $^{12}C = 12$ scale, the **isotopic mass** of an isotope is almost, but not exactly, equal to the number of protons and neutrons in the nucleus.

ℹ The relative atomic mass of an element is the weighted average of the atomic masses of the isotopes of the element on the ^{12}C scale.

To calculate the average of the relative masses of the isotopes that exist in a naturally occurring mixture of an element, you must consider the relative abundances of each isotope. As shown in Figure 9.1.2, there are two isotopes for the element chlorine. Table 9.1.1. shows their relative abundance.

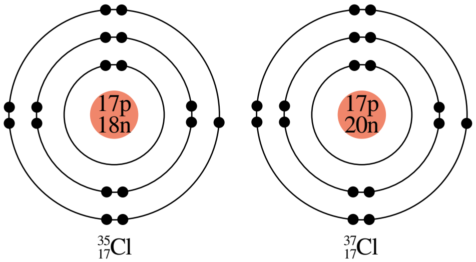


FIGURE 9.1.2 The two isotopes of the element chlorine. Different numbers of neutrons in the nuclei of these atoms give the atoms different masses.

TABLE 9.1.1 The isotopic composition of chlorine

Isotope	Isotopic mass	Relative abundance (%)
^{35}Cl	34.969	75.80
^{37}Cl	36.966	24.20

The data about the two isotopes of chlorine is used to calculate the relative atomic mass of chlorine. A weighted average mass is calculated by using the isotopic masses and abundances to find the total mass of 100 atoms. This mass is then divided by 100 to find the average mass of one atom.

The relative atomic mass (A_r) of an element with two isotopes can be calculated using the formula:

$$A_r = \frac{(\text{isotopic mass} \times \%\ \text{abundance}) + (\text{isotopic mass} \times \%\ \text{abundance})}{100}$$

The periodic table at the end of the book provides relative atomic masses, calculated by taking into account the relative abundances of all the natural isotopes of each element.

Worked example 9.1.1

CALCULATING RELATIVE ATOMIC MASS FROM ISOTOPIC MASSES AND PERCENTAGE ABUNDANCES

Determine the relative atomic mass of chlorine from the data obtained from its mass spectrum. (Refer to Table 9.1.1.)	
Thinking	**Working**
Determine the isotopic masses and abundances of each isotope.	first isotope: isotopic mass 34.969; abundance 75.80% second isotope: isotopic mass 36.966; abundance 24.20%
Substitute the isotopic masses and abundances into the formula for calculating relative atomic mass: $A_r = \frac{(\text{isotopic mass} \times \%\ \text{abundance}) + (\text{isotopic mass} \times \%\ \text{abundance})}{100}$	$A_r = \frac{(34.969 \times 75.80) + (36.966 \times 24.20)}{100}$
Calculate the relative atomic mass.	$A_r = \frac{2650.65 + 894.58}{100}$ $= 35.452$
Express the answer to 2 decimal places.	$A_r(Cl) = 35.45$

SKILLBUILDER 9.1.1

Significant figures

The accuracy of an answer in chemistry calculations is limited by the accuracy of the information given. An answer has the same number of significant figures as there are in the least accurate piece of information. In Worked example 9.1.1, isotopic masses have 5 significant figures, but the relative abundances have only 4 significant figures. Therefore, the answer to this question has 4 significant figures. (See Chapter 1 in your eBook for a more detailed treatment of significant figures.)

➤ Try yourself 9.1.1

CALCULATING RELATIVE ATOMIC MASS FROM ISOTOPIC MASSES AND PERCENTAGE ABUNDANCES

Boron has two isotopes. Their isotopic masses and percentage abundances are provided. Calculate the relative atomic mass of boron.

Isotope	Isotopic mass	Relative abundance (%)
^{10}B	10.013	19.91
^{11}B	11.009	80.09

RELATIVE MOLECULAR MASS

Some elements and compounds exist as **molecules**; for example oxygen (O_2), nitrogen (N_2) and carbon dioxide (CO_2). In Chapter 8, you learnt how non-metallic elements bond together to form covalent molecules. For these substances, a **relative molecular mass (*M_r*)** can be determined. The relative molecular mass is the mass of a molecule relative to the mass of an atom of ^{12}C taken as 12 units exactly. It is numerically equal to the sum of the relative atomic masses of the elements in the formula.

Worked example 9.1.2

CALCULATING THE RELATIVE MOLECULAR MASS OF MOLECULES

Calculate the relative molecular mass of carbon dioxide (CO_2).

Thinking	Working
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(C) = 12.01$ $A_r(O) = 16.00$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × C atom 2 × O atoms
Determine the relative molecular mass by adding the appropriate relative atomic masses.	$M_r = 1 \times A_r(C) + 2 \times A_r(O)$ $= 1 \times 12.01 + 2 \times 16.00$ $= 44.01$

➤ Try yourself 9.1.2

CALCULATING THE RELATIVE MOLECULAR MASS OF MOLECULES

Calculate the relative molecular mass of nitric acid (HNO_3).

RELATIVE FORMULA MASS

You learnt in Chapter 7 that some compounds, such as sodium chloride (NaCl) and magnesium oxide (MgO), do not exist as molecules but rather as ionic lattices. For ionic compounds, the term **relative formula mass** is used. Relative formula mass, like relative molecular mass, is calculated by taking the sum of the relative atomic masses of the elements in the formula.

The relative formula mass is the mass of a formula unit relative to the mass of an atom of ^{12}C taken as 12 amu exactly. It is numerically equal to the sum of the relative atomic masses of the elements in the formula.

Worked example 9.1.3

CALCULATING THE RELATIVE FORMULA MASS OF IONIC COMPOUNDS

Calculate the relative formula mass of magnesium hydroxide ($Mg(OH)_2$).	
Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(Mg) = 24.31$ $A_r(O) = 16.00$ $A_r(H) = 1.01$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × Mg atom 1 × 2 = 2 O atoms 1 × 2 = 2 H atoms
Determine the relative formula mass by adding the appropriate relative atomic masses.	Relative formula mass $= 1 \times A_r(Mg) + 2 \times A_r(O) + 2 \times A_r(H)$ $= 24.31 + 2 \times 16.00 + 2 \times 1.01$ $= 58.33$

➤ Try yourself 9.1.3

CALCULATING THE RELATIVE FORMULA MASS OF IONIC COMPOUNDS

Calculate the relative formula mass of copper(II) nitrate ($Cu(NO_3)_2$).

9.1 Review

SUMMARY

- Atomic mass is the mass of an atom in atomic mass units.
- Most elements consist of a mixture of isotopes.
- The most common isotope of carbon, carbon-12 (^{12}C), is used as the reference standard to compare the masses of atoms.
- The carbon-12 isotope is assigned a mass of exactly 12 amu.
- The isotopic mass and relative abundance of isotopes can be measured using a mass spectrometer.
- The relative atomic mass, A_r, of an element is a weighted average of its isotopic masses.
- The relative molecular mass of molecules, or formula mass of ionic compounds, is calculated from the sum of the relative atomic masses of its constituent elements.

KEY QUESTIONS

Retrieval

1 Describe the concept of atomic mass.

2 State the symbol for relative molecular mass.

3 Recall what the atomic mass of an atom is mainly determined by.

4 Recall how relative molecular mass and relative formula mass are calculated.

Comprehension

5 Explain how atomic mass is determined.

6 Determine why the ionic compound copper(II) sulfate ($CuSO_4$) does not have a relative molecular mass.

Analysis

7 Differentiate between relative atomic mass and atomic mass.

8 Determine if the relative atomic mass for chlorine, shown in Table 9.1.1 on page 225, could be calculated by adding the isotopic masses of the two isotopes together and dividing by two.

9 Determine the relative atomic mass of carbon. Consider that carbon has two isotopes: ^{12}C which has an isotopic mass of 12.000 amu and a relative isotopic abundance of 98.89% and ^{13}C with an isotopic mass of 13.003 amu and a relative isotopic abundance of 1.11%.

10 Consider the data for the isotopic composition of sulfur in the table below.

Element	**Isotope**	**Isotopic mass**	**Relative isotopic abundance (%)**
sulfur	^{32}S	31.972	94.99
	^{33}S	32.971	0.75
	^{34}S	33.967	4.25
	^{36}S	35.967	0.01

Calculate the relative atomic mass of sulfur.

11 Calculate the relative molecular masses of:

a sulfuric acid (H_2SO_4)
b ammonia (NH_3)
c ethane (C_2H_6)

12 Calculate the relative formula mass of:

a potassium chloride (KCl)
b sodium carbonate (Na_2CO_3)
c aluminium sulfate ($Al_2(SO_4)_3$)

9.2 Introducing the mole

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the concept of the mole
- use the mole concept to calculate:
  - number of particles
  - moles of a substance
  - molar mass
  - mass of a substance
- explain the relationship between the mass, moles and molar mass of a substance.

It is often essential for chemists to be able to measure an exact number of particles of an element or compound. However, the particles in elements and compounds are so small that it would be difficult to count atoms, ions or molecules individually or even by the thousands of millions. If it were possible to count individual particles, the numbers in even very small samples would be huge and very inconvenient to work with.

The ice cubes shown in Figure 9.2.1 each contain more than 10^{23} water molecules (H_2O). As each water molecule is composed of two hydrogen atoms and one oxygen atom, the number of individual atoms in each ice cube is greater than 10^{23}. A quantity that allows chemists to measure accurate amounts of extremely small particles is required. In this module, you will learn about the very convenient quantity used by chemists: the mole.

FIGURE 9.2.1 Each of these ice cubes contains more than 10^{23} water (H_2O) molecules. Molecules and atoms are so small, and the numbers of them in everyday samples are so large, that it would be very inconvenient to always count them individually.

WHAT IS A MOLE?

A dozen is a convenient quantity for buying the eggs shown in Figure 9.2.2. For atoms, ions and molecules, a quantity that describes a much larger number is needed. The accepted quantity for chemists is the **mole**, and the number of moles is given the symbol n.

One mole of any substance is defined as the same number of particles as there are atoms in exactly 12 grams of carbon-12. This number of atoms has been experimentally determined to be $6.022\,141 \times 10^{23}$ atoms. That's 602 214 100 000 000 000 000 000 atoms! This number is commonly rounded to 6.02×10^{23} and is referred to as **Avogadro's number** or **Avogadro's constant**. It is given the symbol N_A. You may have noticed that Avogadro's number is written in scientific notation. You can revise your understanding of scientific notation in Chapter 1 of your eBook.

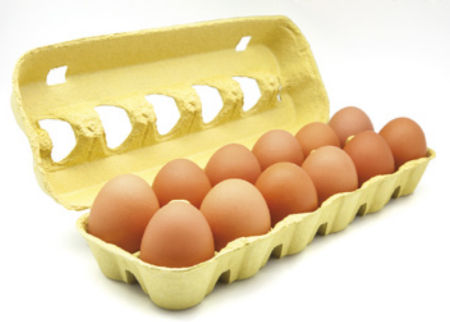


FIGURE 9.2.2 Twelve eggs is one dozen, 24 eggs is two dozen and six eggs is half a dozen.

Chemists use the mole as a counting measure. Figure 9.2.3 shows some quantities that you would be very familiar with, such as a pair, dozen or ream. One dozen is equal to 12, two dozen equals 24, 20 dozen equals 240, and half a dozen equals 6. In the same way, chemists know that 2 moles, 20 moles and half a mole are all multiples of 6.02×10^{23}. This means 2 moles would be 12.04×10^{23}, 20 moles would be 120.4×10^{23}, and so on. Later in this module you will learn how to use the mole to calculate the number of particles present in a substance.

A mole is the amount of substance that contains the same number of fundamental particles as there are atoms in 12 g of carbon-12. This number is 6.02×10^{23} and is referred to as Avogadro's number.

FIGURE 9.2.3 Convenient quantities that you would be very familiar with. A pair of shoes equals two, a dozen roses equals 12, a ream of paper equals 500 sheets.

FIGURE 9.2.4 One mole of sodium chloride and 1 mole of water contain the same number of particles, 6.02×10^{23}.

FIGURE 9.2.5 One mole of any substance contains 6.022×10^{23} particles.

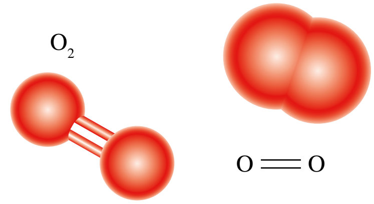


FIGURE 9.2.6 Four different ways chemists represent the oxygen molecule: a formula (O_2), a structural formula (O=O) and two coloured molecular models. Each formula or model shows that one oxygen molecule contains two oxygen atoms.

AVOGADRO'S NUMBER

Avogadro's number is named in honour of Italian scientist Amedeo Avogadro (1776–1856) who developed a hypothesis, now known as Avogadro's law, that states that equal volumes of gases at a given temperature and pressure contain the same number of molecules.

Avogadro's number is an enormous number, but the extremely small size of atoms, ions and molecules means that 1 mole of most elements and compounds does not take up a great deal of mass or volume. For example, in Figure 9.2.4, you can see that 1 mole of water has a volume of 18 mL.

Avogadro's number can be used in calculations. If you know that 1 mole of a substance contains 6.02×10^{23} particles, it follows that:

- 2 moles of a substance would contain $2 \times (6.02 \times 10^{23}) = 1.204 \times 10^{24}$ particles
- 0.3 moles of a substance would contain $0.3 \times (6.02 \times 10^{23}) = 1.81 \times 10^{23}$ particles.

Conversely, the number of moles can be calculated if the number of particles is known:

- 4.70×10^{23} particles $= \frac{4.70 \times 10^{23}}{6.02 \times 10^{23}} = 0.781\,\text{mol}$
- 7.35×10^{24} particles $= \frac{7.35 \times 10^{24}}{6.02 \times 10^{23}} = 12.2\,\text{mol}$

It is very difficult to imagine just how big Avogadro's number really is, especially when atoms, ions and molecules are so small. Here are some examples to help.

- 6.02×10^{23} grains of sand, placed side by side, would stretch from Earth to the Sun and back about 7 million times.
- A computer counting 10 billion times every second would take nearly 2 million years to reach 6.02×10^{23}.
- 6.02×10^{23} of the marshmallows shown in Figure 9.2.5 would cover Australia to a depth of 900 km!

INFORMATION PROVIDED BY FORMULAS

When referring to a mole of a substance, it is important to indicate which particle is being specified. In Chapter 8, you learnt that an oxygen molecule contains two oxygen atoms joined by covalent bonding. Figure 9.2.6 shows four different ways that chemists use to represent the oxygen molecule.

The expression, 'one mole of oxygen' is ambiguous because it could describe 1 mole of oxygen atoms (O) or 1 mole of oxygen molecules (O_2). As there are two atoms in each oxygen molecule, 1 mole of oxygen molecules will contain 2 moles of oxygen atoms. The **molecular formula** of a substance indicates the number of atoms of each element in 1 molecule of a substance. For example, in 1 mole of oxygen gas, there are 2 moles of oxygen atoms. Some other examples of the use of the mole as a counting unit are provided in Table 9.2.1.

TABLE 9.2.1 Examples of the use of the mole as a counting unit

Number of moles of element or compound	Information that can be obtained about numbers of particles
1 mole of hydrogen atoms (H)	1 mole of hydrogen atoms (H)
1 mole of hydrogen molecules (H_2)	1 mole of hydrogen molecules (H_2) 2 moles of hydrogen atoms (H)
2 moles of aluminium atoms (Al)	2 moles of aluminium atoms (Al)
2 moles of calcium fluoride (CaF_2)	2 moles of Ca^{2+} ions 4 moles of F^- ions
10 moles of glucose ($C_6H_{12}O_6$) molecules	10 moles of glucose ($C_6H_{12}O_6$) molecules 60 moles of carbon atoms 120 moles of hydrogen atoms 60 moles of oxygen atoms

CALCULATIONS USING THE MOLE CONCEPT

Four quantities have been introduced so far:

- the mole, which is given the unit **mol**
- the number of moles, which is given by the symbol n
- Avogadro's number, which is given the symbol N_A and has the value 6.02×10^{23}
- the actual number of particles (atoms, ions or molecules), which is given the symbol N.

As you can see from Figure 9.2.7, there is a mathematical relationship that exists between these quantities. This relationship can be written as the formula:

$$n = \frac{N}{N_A}$$

The mole–particle relationship can be used to perform chemical calculations. If the number of moles of a substance is known, then it allows the number of particles to be calculated. Conversely, if the number of particles is known, then the number of moles can be calculated.

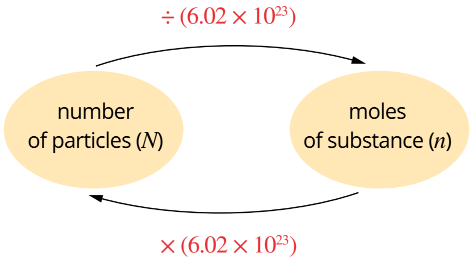


FIGURE 9.2.7 Relationship between number of particles and amount of substance in moles

Worked example 9.2.1

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 3.5 mol of water (H_2O).	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of water molecules is the unknown, so: $N(H_2O) = ?$ $n(H_2O) = 3.5\,\text{mol}$ $N_A = 6.02 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(H_2O) = n \times N_A$
Substitute in data and solve for the answer.	$N(H_2O) = n \times N_A$ $= 3.5 \times 6.02 \times 10^{23}$ $= 2.1 \times 10^{24}$ molecules

➤ Try yourself 9.2.1

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 1.6 mol of carbon dioxide (CO_2).

Worked example 9.2.2

CALCULATING THE NUMBER OF ATOMS

Calculate the number of oxygen atoms in 2.5 mol of oxygen gas (O_2).

Thinking	**Working**
Calculate the number of oxygen atoms. Include units.	The number of oxygen atoms is the unknown, so: $N(O) = ?$ $n(O_2) = 2.5\text{ mol}$ $N_A = 6.02 \times 10^{23}$
Calculate the number of moles of oxygen atoms from the number of moles of oxygen molecules and the molecular formula.	$n(O) = n(O_2) \times 2$ $= 2.5 \times 2$ $= 5.0\text{ mol}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N = n \times N_A$
Substitute in data and solve for the answer.	$N(O) = n \times N_A$ $= 5.0 \times 6.02 \times 10^{23}$ $= 3.0 \times 10^{24}$ atoms

➤ Try yourself 9.2.2

CALCULATING THE NUMBER OF ATOMS

Calculate the number of hydrogen atoms in 0.35 mol of methane (CH_4).

Worked example 9.2.3

CALCULATING THE NUMBER OF MOLES OF PARTICLES GIVEN THE NUMBER OF PARTICLES

Calculate the number of moles of ammonia molecules (NH_3) represented by 2.5×10^{22} ammonia molecules.

Thinking	**Working**
Calculate the number of moles of ammonia. Include units.	The number of moles of ammonia molecules is the unknown, so: $n(NH_3) = ?$ $N(NH_3) = 2.5 \times 10^{22}$ molecules $N_A = 6.02 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ n is the unknown so rearrangement is not required
Substitute in data and solve for the answer.	$n(NH_3) = \frac{N}{N_A}$ $= \frac{2.5 \times 10^{22}}{6.02 \times 10^{23}}$ $= 0.042\text{ mol}$

➤ Try yourself 9.2.3

CALCULATING THE NUMBER OF MOLES OF PARTICLES GIVEN THE NUMBER OF PARTICLES

Calculate the number of moles of magnesium atoms represented by 8.1×10^{20} magnesium atoms.

Worked example 9.2.4

CALCULATING THE NUMBER OF MOLES OF ATOMS GIVEN THE NUMBER OF MOLES OF MOLECULES

Calculate the number of moles of hydrogen atoms in 3.6 mol of sulfuric acid (H_2SO_4).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of moles of hydrogen atoms is the unknown, so: $n(\text{H}) = ?$ $n(\text{H}_2\text{SO}_4) = 3.6\,\text{mol}$
Calculate the number of moles of hydrogen atoms from the amount of sulfuric acid molecules and the molecular formula.	$n(\text{H}) = n(\text{H}_2\text{SO}_4) \times 2$ $= 3.6 \times 2$ $= 7.2\,\text{mol}$

➤ Try yourself 9.2.4

CALCULATING THE NUMBER OF MOLES OF ATOMS GIVEN THE NUMBER OF MOLES OF MOLECULES

Calculate the number of moles of hydrogen atoms in 0.75 mol of water (H_2O).

MOLAR MASS

Chemical laboratories always contain a balance like the one in Figure 9.2.8, which is used for weighing. Since a chemist knows that a mass of a specific substance always contains a specific number of particles, it is possible to easily weigh a sample of the substance and calculate the exact number of particles present in the sample.

Chemists have cleverly defined the mole so that you can determine the number of moles of a substance by simply measuring its mass. The particles of different elements and compounds have different masses. Therefore, the masses of 1 mole of different elements or compounds will also be different. This is like saying that the mass of one dozen oranges will be greater than the mass of one dozen mandarins because one orange is heavier than one mandarin.

The mass, in grams, of 1 mole of a particular element or compound is known as its **molar mass**. It is given the symbol M and the unit g mol^{-1}. Remember that a mole is defined as the amount of substance that contains the same number of specified particles as there are atoms in 12 g of carbon-12. This is a very convenient definition because:

- one atom of ^{12}C has a relative atomic mass of 12 amu exactly
- 1 mole of atoms of ^{12}C has a mass of 12 g exactly.

FIGURE 9.2.8 A digital balance is a simple piece of laboratory equipment used for weighing.

WS 1.3.7

In general, the molar mass of an element is the mass of 1 mole of the element. The molar mass of a compound is the mass of 1 mole of the compound. It is equal to the relative molecular or relative formula mass of the compound expressed in grams. The molar mass is given the symbol, *M*, and the unit, g mol^{-1}.

Naturally occurring carbon is mainly composed of the ^{12}C isotope, so the molar mass of carbon is 12.01 g mol^{-1}. Consider an atom of ^{12}C and an atom of ^{24}Mg. One atom of ^{12}C has been assigned an isotopic mass of 12 exactly. On that scale, the isotopic mass of ^{24}Mg is approximately 24. Since 1 mole of ^{12}C atoms weighs exactly 12 g, then 1 mole of ^{24}Mg must weigh approximately twice as much, 24 g.

Table 9.2.2 shows you how to calculate the molar masses of some common substances. From this table, and the photograph in Figure 9.2.9 of 1 mole of some common substances, you can see that 1 mole of each substance has a different mass.

TABLE 9.2.2 Calculating the molar mass of a substance by adding the relative atomic masses for each atom present in the substance, based on the molecular or ionic formula

Substance	Relative atomic masses	Molar mass of substance
Na	Na: 22.99	= 23.0 g mol^{-1}
O_2	O: 16.00	= 2 × 16.00 = 32.0 g mol^{-1}
H_2O	H: 1.01 O: 16.00	= (2 × 1.01) + 16.00 = 18.0 g mol^{-1}
CO_2	C: 12.01 O: 16.00	= 12.01 + (2 × 16.00) = 44.0 g mol^{-1}
$NaNO_3$	Na: 22.99 N: 14.01 O: 16.00	= 22.99 + 14.01 + (3 × 16.00) = 85.0 g mol^{-1}

FIGURE 9.2.9 One mole of each substance has a different mass.

MOLE–MASS CALCULATIONS

It is possible to calculate the mass of a sample if the moles and molar mass are known. This relationship between the number of moles of a substance (n), its molar mass (M) and the given mass of the substance (m) can be summarised with the following formula:

$$m = n \times M$$

This formula is rearranged in Figure 9.2.10 to solve for an unknown number of moles:

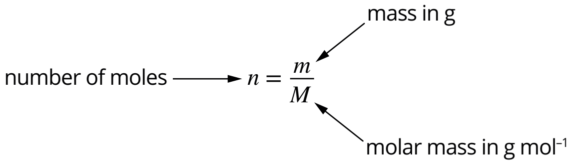


FIGURE 9.2.10 One mole of each substance has a different mass.

Worked example 9.2.5

CALCULATING THE MASS OF A SUBSTANCE

<table>
<tr><td colspan="2">Calculate the mass of 0.35 mol of magnesium nitrate ($Mg(NO_3)_2$), to 2 significant figures.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>List the data given to you in the question.
Remember that whenever you are given a formula, you can calculate the molar mass.</td><td>$m(Mg(NO_3)_2) = ?\ g$
$n(Mg(NO_3)_2) = 0.35\ mol$
$M(Mg(NO_3)_2) = 24.30 + (2 \times 14.01) + (6 \times 16.00)$
$= 148.33\ g\ mol^{-1}$
$= 150\ g\ mol^{-1}$ to 2 significant figures</td></tr>
<tr><td>Calculate the mass of magnesium nitrate using:
$n = \frac{m}{M}$</td><td>$n = \frac{m}{M}$ so $m = n \times M$
$m(Mg(NO_3)_2) = 0.35 \times 148.33$
$= 52.00\ g$</td></tr>
</table>

➤ Try yourself 9.2.5

CALCULATING THE MASS OF A SUBSTANCE

Calculate the mass of 4.68 mol of sodium carbonate (Na_2CO_3).

The relationship between the number of moles and molar mass of a substance allows you to determine the number of moles of the substance or quantity of atoms present by weighing the substance. This is because the number of particles present in a substance is equal to the number of moles of the substance multiplied by 6.02×10^{23}. Some calculations require you to use both the formulas $n = \frac{m}{M}$ and $n = \frac{N}{N_A}$.

Worked example 9.2.6

CALCULATING THE NUMBER OF MOLECULES

<table>
<tr><td colspan="2">Calculate the number of CO_2 molecules present in 22 g of carbon dioxide.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>List the data given to you in the question.
Convert mass to grams if required. Remember that whenever you are given a formula, you can calculate the molar mass.</td><td>$N(CO_2) = ?$
$M(CO_2) = 12.01 + (2 \times 16.00)$
$= 44.01\ g\ mol^{-1}$
$m(CO_2) = 22\ g$</td></tr>
<tr><td>Calculate the number of moles of CO_2 using:
$n = \frac{m}{M}$</td><td>$n(CO_2) = \frac{m}{M}$
$= \frac{22}{44.01}$
$= 0.50\ mol$</td></tr>
<tr><td>Calculate the number of CO_2 molecules using:
$n = \frac{N}{N_A}$</td><td>$n = \frac{N}{N_A}$ so $N = n \times N_A$
$N(CO_2) = 0.50 \times 6.02 \times 10^{23}$
$= 3.0 \times 10^{23}$ molecules</td></tr>
</table>

➤ Try yourself 9.2.6

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of sucrose molecules in a teaspoon (4.2 g) of sucrose ($C_{12}H_{22}O_{11}$).

9.2 Review

SUMMARY

- A mole is a convenient quantity for counting particles. The number of moles is given the symbol *n* and expressed in units of mol.
- One mole is defined as the amount of substance that contains the same number of 'specified' particles as there are atoms in 12 g of carbon-12.
- The number of particles in 1 mole is given the symbol N_A. This is known as Avogadro's constant and has the numerical value of 6.02×10^{23}.
- The formula $n = \frac{N}{N_A}$ can be used or rearranged to calculate the amount or number of specified particles in a sample.
- The molar mass of an element or compound is the mass, in grams, of 1 mole of that element or compound. Molar mass is given the symbol *M* and the unit $g\,mol^{-1}$.
- The molar mass of an element or compound has the same numerical value as the relative mass of the element or compound.
- The formula $n = \frac{m}{M}$ can be used or rearranged to calculate the mass, amount or molar mass of an element or compound.

KEY QUESTIONS

Retrieval

1 State what the symbol N_A represents.

2 Describe the concept of the mole.

3 Indicate if the following statement is true or false. Molar mass is given the symbol *m* and the unit $g\ mol^{-1}$.

4 State the formula used to calculate the moles of a substance from a given mass.

Comprehension

5 Explain why the expression 'one mole of oxygen' is ambiguous.

6 Describe how molar mass of a compound is calculated.

Analysis

7 Calculate the following quantities using the relationship between moles and the number of particles.
  - **a** number of atoms in 2.0 moles of sodium atoms (Na)
  - **b** number of molecules in 0.10 moles of nitrogen molecules (N_2)
  - **c** number of atoms in 1.0×10^{-2} moles of iron atoms (Fe)
  - **d** number of molecules in 4.62×10^{-5} moles of CO_2 molecules

8 Calculate the number of moles represented by the following using the relationship between moles and the number of particles:
  - **a** 3.0×10^{23} molecules of water (H_2O)
  - **b** 1.5×10^{23} atoms of neon (Ne)
  - **c** 4.2×10^{25} atoms of iron (Fe)

9 **a** Calculate the number of moles of the following using the relationship between the number of moles and the number of particles.
  - **i** sodium atoms represented by 1.0×10^{20} sodium atoms
  - **ii** aluminium atoms represented by 1.0×10^{20} aluminium atoms
  - **iii** chlorine molecules represented by 1.0×10^{20} chlorine molecules

  b The answers to the above show that the different substances contain the same number of moles. Determine what else is different about them.

  Remember: Avogadro's number, N_A, is 6.02×10^{23}.

10 Calculate the number of moles of the following using the relationship between moles and the number of particles.
  - **a** chlorine atoms in 0.4 mol of chlorine (Cl_2)
  - **b** hydrogen atoms in 1.2 mol of methane (CH_4)
  - **c** hydrogen atoms in 0.12 mol of ethane (C_2H_6)
  - **d** oxygen atoms in 1.5 mol of sodium sulfate (Na_2SO_4)

11 Calculate the molar mass of the following substances.
  - **a** nitrogen (N_2)
  - **b** ammonia (NH_3)
  - **c** sulfuric acid (H_2SO_4)
  - **d** iron(III) nitrate ($Fe(NO_3)_3$)
  - **e** ethanoic acid (CH_3COOH)
  - **f** sulfur atoms (S)
  - **g** vitamin C (ascorbic acid, $C_6H_8O_6$)
  - **h** hydrated copper(II) sulfate ($CuSO_4{\cdot}5H_2O$)

12 Calculate the following masses using the relationship between mass, moles and molar mass.

- **a** 1.0 mol of sodium atoms (Na)
- **b** 2.0 mol of oxygen molecules (O_2)
- **c** 0.10 mol of methane molecules (CH_4)
- **d** 0.25 mol of aluminium oxide (Al_2O_3)

13 Calculate the number of moles of these substances using the relationship between mass, moles and molar mass.

- **a** H atoms in 5 g of hydrogen (H_2)
- **b** H_2 molecules in 5 g of hydrogen (H_2)
- **c** Al atoms in 2.7 g of aluminium (Al)
- **d** CH_4 molecules in 0.4 g of methane (CH_4)
- **e** O_2 molecules in 0.10 g of oxygen (O_2)
- **f** O atoms in 0.10 g of oxygen (O_2)
- **g** P_4 molecules in 1.2×10^{-3} g of phosphorus (P_4)
- **h** P atoms in 1.2×10^{-3} g of phosphorus (P_4)

14 Apply the relationship between number of particles and moles with the relationship between mass, moles and molar mass to calculate the number of atoms in the following substances.

- **a** 23 g of sodium (Na)
- **b** 4.0 g of argon (Ar)
- **c** 0.243 g of magnesium (Mg)
- **d** 10.0 g of gold (Au)

15 Apply the relationship between number of particles and moles with the relationship between mass, moles and molar mass to calculate the:

- **a** number of molecules in:
  - **i** 16 g of oxygen (O_2)
  - **ii** 2.8 g of nitrogen (N_2)
- **b** number of oxygen atoms in 3.2 g of sulfur dioxide (SO_2)
- **c** total number of atoms in 288 g of ammonia (NH_3)

9.3 Percentage composition and empirical formulas

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the concept of percentage composition
- calculate percentage composition for a compound
- explain why empirical formulas express the simplest whole number ratio of elements in a compound
- explain the difference between the empirical and molecular formulas for a molecular compound
- determine the empirical formula using data on percentage composition
- determine the molecular formula for a molecular compound based on its molar mass and empirical formula.

A mole is a convenient quantity for counting particles. The number of moles is given the symbol *n* and expressed in units of mol.

Compounds are substances that contain two or more different elements. Consider the example of the compound carbon dioxide, depicted in Figure 9.3.1. A single molecule of carbon dioxide consists of one carbon atom bonded to two oxygen atoms. Likewise, 1 mole of carbon dioxide is made up of 1 mole of carbon atoms and 2 moles of oxygen atoms. The relative proportions of the atoms that make up the compound are expressed in its formula, CO_2.

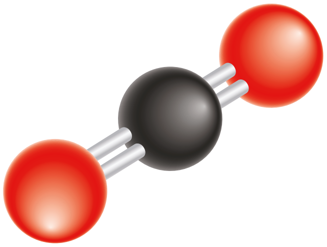


FIGURE 9.3.1 This is a molecular model of a single molecule of carbon dioxide (CO_2), with one carbon atom (black) bonded to two oxygen atoms (red).

The formulas of compounds are an important part of the language of chemistry. You would be familiar with some of the formulas shown in Figure 9.3.2. Formulas are used to represent chemicals in equations and on numerous other occasions.

While formulas show the proportions of the atoms, in the laboratory, chemicals are commonly measured by their mass. The formula does not directly tell you the proportional mass of the elements in a compound. If you consider the above example of CO_2, the ratio of carbon to oxygen atoms is 1:2. However, carbon and oxygen have different atomic masses, so the ratio by mass will not be in the ratio 1:2. The ratio by mass is called the **percentage composition**.

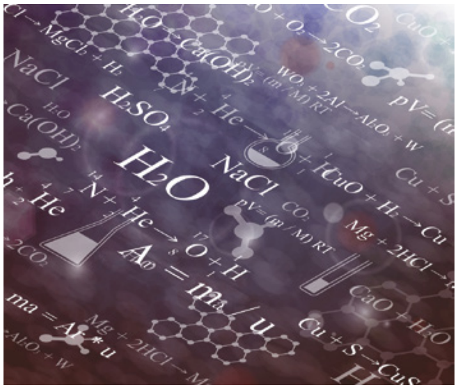


FIGURE 9.3.2 Chemical formulas are part of the language of chemistry.

PERCENTAGE COMPOSITION

The percentage composition of a given compound tells you the proportion by mass of the different elements in that compound. The proportion of each element is expressed as a percentage of the total mass of the compound.

Being able to determine percentages by mass is important in chemistry. For example, the bauxite ore which is mined in Weipa, Queensland, is refined into aluminium. Bauxite is mostly composed of the chemical aluminium hydroxide, which has the formula $Al(OH)_3$. During the refining process, aluminium hydroxide is reacted to produce aluminium oxide (alumina), which is finally converted to aluminium metal.

A company that is producing aluminium from aluminium hydroxide would want to know how much metal can potentially be extracted from a given quantity of ore. The pie chart in Figure 9.3.3 shows that by mass the percentage of aluminium in aluminium hydroxide is approximately 35%. This means 35 g can potentially be extracted from every 100 g of ore.

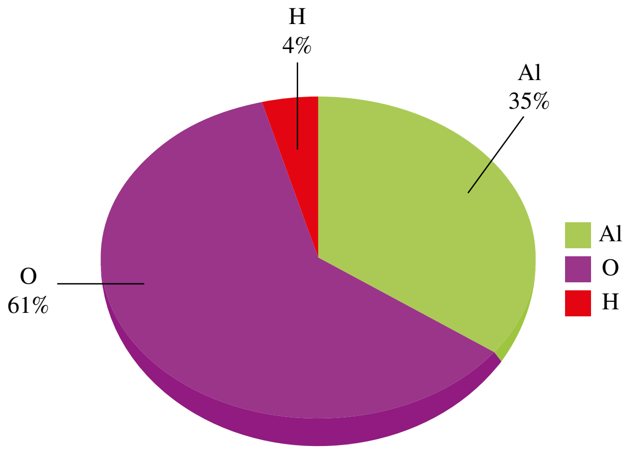


FIGURE 9.3.3 This pie chart shows the percentage composition, by mass, of aluminium hydroxide.

If the chemical formula of a compound is known, the percentage composition can be determined using the molar masses of the elements compared to the molar mass of the compound.

$$\% \text{ by mass of an element in a compound} = \frac{\text{mass of the element in 1 mol of the compound}}{\text{molar mass of the compound}} \times 100$$

Worked example 9.3.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of aluminium in alumina (Al_2O_3).	
Thinking	**Working**
Find the molar mass of the compound.	$M(Al_2O_3) = (2 \times 27.00) + (3 \times 16.00)$ $= 101.96\ \text{g mol}^{-1}$
Find the total mass of the element in 1 mol of the compound.	mass of Al in 1 mol $= 2 \times M(\text{Al})$ $= 2 \times 26.98$ $= 53.96\ \text{g}$
Find the percentage by mass of the element in the compound.	% by mass of Al in Al_2O_3 $= \frac{\text{mass of Al in 1 mol of } Al_2O_3}{\text{molar mass of } Al_2O_3} \times 100$ $= \frac{53.96}{101.96} \times 100$ $= 52.92\%$

➤ Try yourself 9.3.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3).

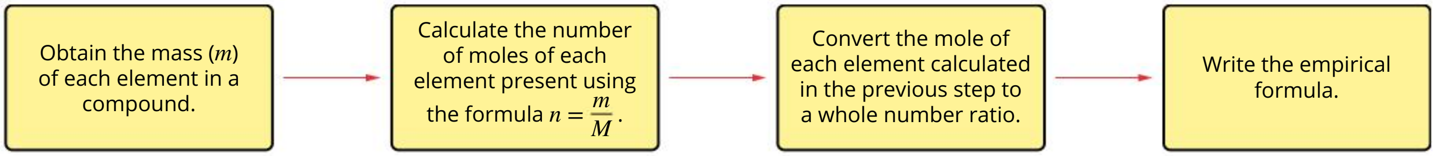
EMPIRICAL FORMULAS

Atoms or ions are present in compounds in fixed whole number ratios. The **empirical formula** of a compound gives the simplest whole number ratio of elements in that compound. Some examples are given in Table 9.3.1.

TABLE 9.3.1 Empirical formulas of some common compounds

Compound	Empirical formula	Simplest whole number ratio of elements in the compound
water	H_2O	H : O 2 : 1
ethene	CH_2	C : H 1 : 2
calcium carbonate	$CaCO_3$	Ca : C : O 1 : 1 : 3

The empirical formula for a compound is determined from the mass of each element present in a given mass of the compound. These masses can be determined experimentally. Once the masses of elements in a compound are known, the steps in Figure 9.3.4 are followed to convert these masses to a mole ratio—that is, a ratio by number of atoms—and then to an empirical formula.



Obtain the mass (m) of each element in a compound.

Calculate the number of moles of each element present using the formula $n = \frac{m}{M}$.

Convert the mole of each element calculated in the previous step to a whole number ratio.

Write the empirical formula.

FIGURE 9.3.4 Follow these steps to calculate an empirical formula.

Worked example 9.3.2

DETERMINING THE EMPIRICAL FORMULA FROM PERCENTAGES

A compound of carbon and oxygen contains 27.3% carbon and 72.7% oxygen by mass. Calculate the empirical formula of the compound.

Thinking	**Working**
Write down the mass, in g, of all elements present in the compound. If masses are given as percentages, assume that the sample weighs 100 g, then the percentages become masses in grams.	$m(\text{C}) = 27.3\,\text{g} \quad m(\text{O}) = 72.7\,\text{g}$
Calculate the number of moles of each element in the compound using: $n = \frac{m}{M}$	$n(\text{C}) = \frac{27.3}{12.01} = 2.28\,\text{mol} \quad n(\text{O}) = \frac{72.7}{16.00} = 4.54\,\text{mol}$
Simplify the ratio by dividing each number of moles by the smallest number of moles calculated in the previous step. This gives you a ratio of the number of atoms of each element.	$\frac{2.28}{2.28} = 1 \quad \frac{4.54}{2.28} = 2$
Find the simplest whole number ratio.	1 : 2
Write the empirical formula.	CO_2

➤ Try yourself 9.3.2

DETERMINING THE EMPIRICAL FORMULA FROM PERCENTAGES

A compound contains 60.3% magnesium and 39.7% oxygen. Calculate the empirical formula of the compound.

Worked example 9.3.3

DETERMINING THE EMPIRICAL FORMULA FROM MASS

A 0.50 g sample of magnesium is heated and allowed to completely react with chlorine. A 1.96 g deposit of white powder is formed. Determine the empirical formula of the compound.

Thinking	**Working**
Write down the mass, in g, of all elements present in the compound. If masses are given as percentages, assume that the sample weighs 100 g, then the percentages become masses in grams.	$m(\text{Mg}) = 0.50\,\text{g}$ $m(\text{Cl}) = \text{mass of compound} - \text{mass of Mg}$ $= 1.96 - 0.50$ $= 1.46\,\text{g}$
Calculate the number of moles of each element in the compound using: $n = \frac{m}{M}$	$n(\text{Mg}) = \frac{0.50}{24.31} = 0.0206 \qquad n(\text{Cl}) = \frac{1.46}{35.5} = 0.0411$
Simplify the ratio by dividing each number of moles by the smallest number of moles calculated in the previous step. This gives you a ratio of the number of atoms of each element.	$\frac{0.0206}{0.0206} = 1 \qquad \frac{0.0411}{0.0206} = 2$
Find the simplest whole number ratio.	1 : 2
Write the empirical formula.	$MgCl_2$

➤ Try yourself 9.3.3

DETERMINING THE EMPIRICAL FORMULA FROM MASS

A 16.06 g sample of nitrogen reacts with oxygen to form 62.00 g of a compound containing only nitrogen and oxygen atoms. Determine the empirical formula of the compound.

MOLECULAR FORMULAS

Molecular compounds have a molecular formula in addition to an empirical formula. The molecular formula gives the actual number of atoms of each element present in a molecule, rather than the simplest whole number ratio. The molecular formula can be the same as or different from the empirical formula. The empirical and molecular formulas of some common molecular compounds are shown in Table 9.3.2.

TABLE 9.3.2 Empirical and molecular formulas of some common molecular compounds

Molecule	Molecular formula	Empirical formula
water	H_2O	H_2O
ethane	C_2H_6	CH_3
carbon dioxide	CO_2	CO_2
glucose	$C_6H_{12}O_6$	CH_2O

Ionic compounds do not have molecular formulas because they do not exist as molecules. However, they do have empirical formulas that describe the fixed ratio of ions that exist in their lattices. The formula for calcium chloride ($CaCl_2$) is an example of an empirical formula of an ionic compound.

A molecular formula can be determined from the empirical formula of a compound if the molar mass of the compound is also known. The molecular formula of a molecule is always a whole number multiple of the empirical formula. The number of the multiple is determined by the following formula.

$$\text{Number of empirical formula units in a molecule} = \frac{\text{molar mass of the compound}}{\text{molar mass of one empirical formula unit}}$$

The general steps in the determination of a molecular formula are shown in Figure 9.3.5.

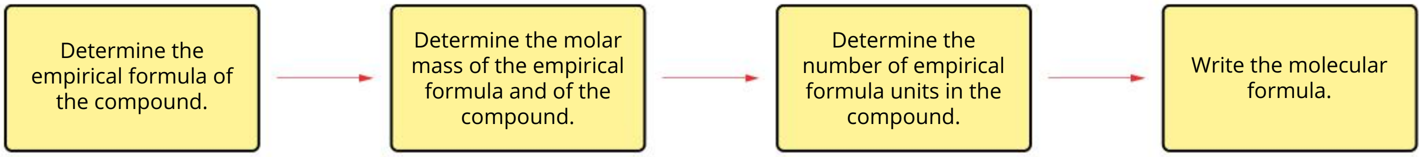


FIGURE 9.3.5 Steps for calculating a molecular formula

Worked example 9.3.4

DETERMINING MOLECULAR FORMULA

A compound has the empirical formula CH. The molar mass of this compound is 78 g mol^{-1}. Determine the molecular formula of the compound.

Thinking	**Working**
Calculate the molar mass of one unit of the empirical formula.	Molar mass of a CH unit = 12.01 + 1.01 = 13.03 g mol^{-1}
Determine the number of empirical formula units in the molecular formula.	Number of CH units = $\frac{78}{13.03}$ = 6
Determine the molecular formula of the compound.	Molecular formula = 6 × CH = C_6H_6

➤ Try yourself 9.3.4

DETERMINING MOLECULAR FORMULA

WS 1.3.9

A compound has the empirical formula C_2H_5. The molar mass of this compound was determined to be 58 g mol^{-1}. Determine the molecular formula of the compound.

9.3 Review

SUMMARY

- The percentage composition of a given compound tells you the proportion by mass of the different elements in that compound.
- The percentage, by mass, of an element in a compound can be calculated from the mass of the element in 1 mole of the compound and the molar mass of the compound.
- The empirical formula of a compound gives the simplest whole number ratio of elements in that compound.
- Molecular compounds have a molecular formula that gives the actual number of atoms of each element in the molecule. It may be the same as, or different from, the empirical formula.
- A molecular formula can be determined from the empirical formula of a compound if the molar mass of the compound is also known.

KEY QUESTIONS

Retrieval

1 Describe the concept of percentage composition.

2 Identify in which of the following compounds the molecular formula is the same as the empirical formula.

A H_2SO_4
B H_2O_2
C $C_6H_{12}O_6$

3 Indicate whether the following statement is true or false. If the chemical formula of a compound is known, then the percentage composition can be determined using the molar masses of the elements compared to the molar mass of the compound.

Comprehension

4 Explain why chemical formulas do not directly reflect the percentage composition of a compound.

5 Explain why ionic compounds only have empirical formulas.

Analysis

6 Calculate the percentage by mass of the following substances.

a iron in iron(III) oxide (Fe_2O_3)
b uranium in uranium oxide (U_3O_8)
c nitrogen in ammonium chloride (NH_4Cl)
d oxygen in copper(II) nitrate ($Cu(NO_3)_2$)

7 Determine the empirical formulas of the compounds with these percentage compositions.

a 2.74% hydrogen, 97.26% chlorine
b 42.9% carbon, 57.1% oxygen

8 Determine the empirical formulas of the compounds with the following masses.

a 10.0 g of a compound of magnesium and oxygen that contains 6.03 g of magnesium
b 3.2 g of a hydrocarbon that contains 2.4 g of carbon

9 Determine the molecular formula of the following compounds.

	Empirical formula	**Relative molecular mass**
a	HO	34
b	CH_2O	90
c	NO_2	46
d	CH_2	154

10 A hydrocarbon contains 85.7% carbon by mass. Its relative molecular mass is 70. Determine the hydrocarbon's:

a empirical formula
b molecular formula

11 A sample of the carbohydrate glucose contains 1.8 g carbon, 0.3 g hydrogen and 2.4 g oxygen.

a Calculate the empirical formula of the compound.
b Deduce its molecular formula given that its relative molecular mass is 180 g mol^{-1}.

9.4 Chemical equations and the mole

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- write chemical equations using words and formulas
- understand that mass is conserved in chemical reactions and physical transformations
- use coefficients to balance chemical equations
- perform stoichiometric calculations using mole ratios
- determine the mass of an unknown reactant or product using a balanced chemical equation
- calculate limiting and excess reagents
- compare actual yield to theoretical yield.

If you think about it, **chemical reactions** are taking place all the time. They occur in the world around you, and even inside your body. In preparing a meal, chemical reactions occur as a result of the cooking process, and after eating the meal your body digests the food and releases its energy through chemical reactions. Plants use chemical reactions to take carbon from the air and turn it into food, and the combustion in a car engine releases the stored energy in fuel.

What do all chemical reactions have in common? The answer is they involve one or more substances changing into a new substance. Figure 9.4.1 shows the ingredients needed to bake a simple cake: eggs, sugar, butter, milk and self-raising flour. Chemical reactions take place when the ingredients are combined and heated, with the result being a cake! What would you need to do if you wanted to double the mixture? You would need to double every ingredient to maintain the correct proportion. Can you predict from the amounts of ingredients that you have, how much cake you will make? Or, if you need to make three cakes, could you determine whether you have enough ingredients in your refrigerator and pantry? That is what this section is about. You will learn how chemical reactions are written, described, balanced and used in calculations.

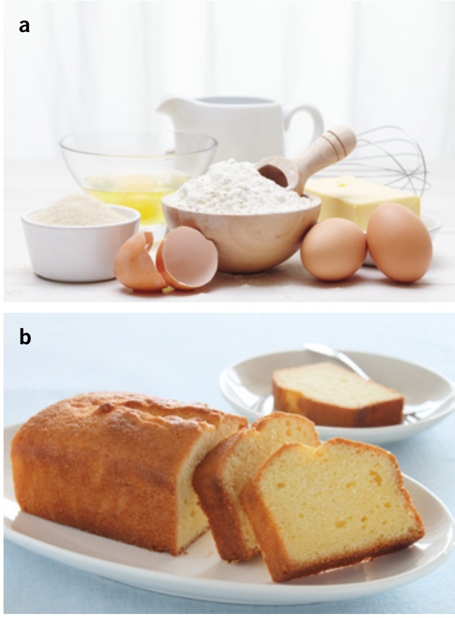


FIGURE 9.4.1 The ingredients needed to bake a simple cake: eggs, sugar, butter, milk and self-raising flour. Chemical changes occur when a cake is baked.

WRITING CHEMICAL EQUATIONS

Chemical reactions are a change in which one or more **reactants** change into one or more **products**. This change is characterised by the breaking of bonds in the reactants and the formation of bonds in the products. As their names suggest, reactants are the substances present at the start of a reaction, and products are the substances produced in a chemical reaction. Using the example of the cake depicted in Figure 9.4.1, the ingredients would be the reactants and the product is the cake.

Chemical equations are visual representations that chemists use to describe chemical reactions. To write a chemical equation, you write the names of the reactants on the left and the products on the right. An arrow (→) is used to separate the reactants and products, and if you are reading the equation you read the arrow as 'yields'. If there is more than one reactant or product they are separated with a plus (+) sign.

Consider the chemical reaction shown in Figure 9.4.2, which shows the **decomposition** of hydrogen peroxide into water and oxygen. In the chemical equation for this reaction, hydrogen peroxide would be on the left because it is the reactant. Water and oxygen would be on the right because they are the products. The equation would be read 'hydrogen peroxide yields water and oxygen'.

FIGURE 9.4.2 Hydrogen peroxide decomposes to form water and oxygen. In the picture, manganese oxide has been added as a catalyst to speed up the reaction.

Some reactions are complicated, involving many chemicals that can have long names. For convenience, chemists usually write chemical equations using formulas. The equation from Figure 9.4.2 when written with formulas would be:

$$H_2O_2 \rightarrow H_2O + O_2$$

Additional information can be added to chemical equations to indicate physical states, the presence of **catalysts** or the need to supply heat. A catalyst is a substance that speeds up a reaction but is not used up in the reaction. The decomposition of hydrogen peroxide can be catalysed with a chemical called manganese oxide (MnO_2). Table 9.4.1 shows some of the commonly used symbols in chemical equations and their meaning.

Chemical reactions occur when one or more reactants change into one or more products. Chemical equations are used to describe chemical reactions. In a chemical equation, the reactants are written on the left and the products on the right.

TABLE 9.4.1 Symbols used in chemical equations

Symbol	Meaning
(s)	substance is in a solid state; placed after the formula
(l)	substance is in a liquid state; placed after the formula
(g)	substance is in a gaseous state; placed after the formula
(aq)	substance is dissolved in water, placed after the formula
$\rightarrow$	yields
+	used to separate two reactants or two products
$\rightleftharpoons$	used in place of $\rightarrow$ for reversible reactions
$\xrightarrow{\text{heat}}$ or $\xrightarrow{\Delta}$	heat is supplied to reaction
$\xrightarrow{MnO_2}$	catalyst used; formula of catalyst written above the yield sign

If you were to write the equation for the breakdown of hydrogen peroxide with all the additional symbols it would be:

$$H_2O_2(aq) \xrightarrow{MnO_2} H_2O(l) + O_2(g)$$

BALANCING CHEMICAL EQUATIONS

In a chemical reaction, the atoms are rearranged to form new substances, but the total number of each atom does not change. If there are a total of two hydrogen atoms that go into a reaction, two hydrogen atoms must come out of the reaction as products. This is because chemical reactions obey the **law of conservation of mass**. The law of conservation of mass states that in any physical change, or chemical reaction, matter can neither be created nor destroyed.

The total mass of all products is equal to the total mass of all reactants in a chemical reaction.

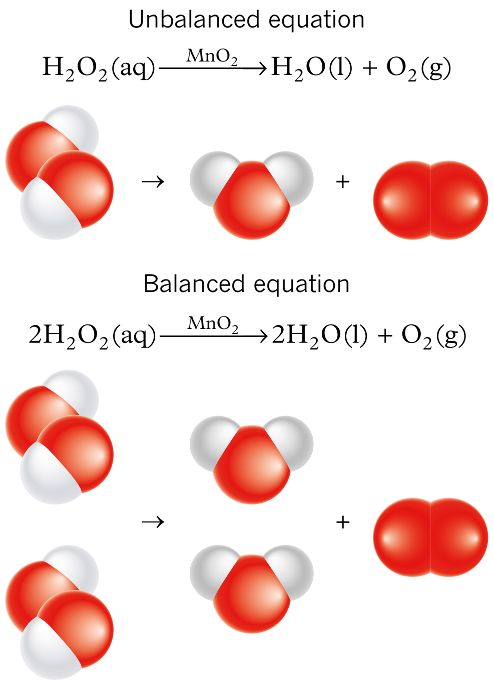
If reactions obey the law of conservation of mass, it follows that chemical equations must also. This means for a chemical equation to be correct, the number of each type of atom on the reactant side must equal the number of each type of atom on the product side. Therefore it is a **balanced equation**. Let's again consider the decomposition of hydrogen peroxide:

$$H_2O_2(aq) \xrightarrow{MnO_2} H_2O(l) + O_2(g)$$

If you tally the atoms, there are two hydrogen atoms on each side of the arrow but as the equation is written there are two oxygen atoms on the reactant side but three oxygen atoms on the product side of the equation. The equation is not balanced. For this equation to be correct, two molecules of hydrogen peroxide would need to react to produce two water molecules and one oxygen molecule.

To write a balanced chemical equation, use coefficients so that the number of each atom on the reactant side equals the the number of each atom on the product side.

To write a balanced equation that reflects the correct proportions of reactant and product particles, **coefficients** must be added. Coefficients are whole numbers placed in front of formulas to make the equation balance. Figure 9.4.3 shows the balanced equation for the breakdown of hydrogen peroxide.



Unbalanced equation

$$H_2O_2(aq) \xrightarrow{MnO_2} H_2O(l) + O_2(g)$$

Balanced equation

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

FIGURE 9.4.3 For the equation to be balanced, and obey the law of conservation of mass, two molecules of H_2O_2 must react to form two molecules of H_2O and 1 molecule of O_2. The balanced equation shows the reactant and product molecules in the correct proportion.

SKILLBUILDER 9.4.1

Writing balanced equations

To write a balanced equation, follow these steps.

1. Find the correct formulas for all the reactants and products.
2. Write a 'skeleton' of the equation—place the formulas for reactants on the left and the formulas for products on the right, separated by a yield (→) sign. If there is more than one reactant or product, separate them with a plus sign.
3. Determine the number of atoms of each element in the reactants and products. Remember that polyatomic ions work as a single unit and appear on both sides of the equation unchanged.
4. Balance the elements one at a time using coefficients. When no coefficient is written, it is 1. Start by balancing elements that appear only once on each side of the equation. Remember, you cannot change the subscripts to balance an equation. Doing so would change the formula to a new substance.
5. Check the numbers of each type of atom on each side to make sure they are equal.
6. Make sure all coefficients are the lowest ratio and add other symbols (such as the state) if known.

Worked example 9.4.1

WRITING A BALANCED EQUATION

Hydrogen and oxygen gas react to form water vapour. The reaction releases energy and is used to fuel rockets. Write a balanced equation for this reaction.

Thinking	**Working**	
Write the correct formulas to give a skeleton.	$H_2 + O_2 \rightarrow H_2O$	
Tally the number of each type of atom.	Reactants: 2 hydrogen 2 oxygen	Products: 2 hydrogen 1 oxygen
Use coefficients to balance the number of atoms. You may need to use trial and error. If the H_2O is doubled to 2, then the O_2 will balance. This unbalances the hydrogen, but that can be corrected by putting a 2 in front of the H_2.	$2H_2 + O_2 \rightarrow 2H_2O$ Reactants: 4 hydrogen 2 oxygen	Products: 4 hydrogen 2 oxygen
Check to make sure both sides balance and write the balanced equation with any known symbols.	$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$	

➤ Try yourself 9.4.1

WRITING A BALANCED EQUATION

Iron metal and chlorine gas react to form solid iron(III) chloride. Write a balanced equation for this reaction.

STOICHIOMETRY AND THE MOLE RATIO

The coefficients used to balance the equations also show the ratios between the reactants and products involved in the reaction. The study of ratios of moles of substances is called **stoichiometry**. Stoichiometric calculations are based on the law of conservation of mass.

Consider the equation for the precipitation reaction that occurs when a solution of calcium chloride reacts with a solution of silver nitrate:

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

The equation indicates that 1 mole of $CaCl_2$ reacts with 2 moles of $AgNO_3$ to form 2 moles of solid AgCl and 1 mole of $Ca(NO_3)_2$.

In more general terms, the number of moles of $AgNO_3$ that reacts will always be double the number of moles of $CaCl_2$ that reacts. The number of moles of AgCl produced will be equal to the number of moles of $AgNO_3$ used and double the number of moles of $Ca(NO_3)_2$ produced.

You can use the coefficients of this reaction to write relationships that show the **mole ratios** of any two chemicals involved in the reaction:

$$\frac{n(AgNO_3)}{n(CaCl_2)} = \frac{2}{1}, \quad \frac{n(CaCl_2)}{n(AgCl)} = \frac{1}{2} \quad \text{and} \quad \frac{n(AgCl)}{n(AgNO_3)} = \frac{2}{2} = \frac{1}{1}$$

ℹ Stoichiometric calculations allow you to use the mole ratio established in a chemical equation to predict the amount of a product that will be formed or how much reactant will be used.

When carrying out any stoichiometric calculation, you must always clearly state the mole ratio you are working with.

In the stoichiometric calculations that you will perform in this module, the number of moles or mass of one of the reactants or products will always be known (called the ‘known chemical’) and is used to determine the number of moles or mass of one of the other reactants or products involved in the reaction (called the ‘unknown chemical’). You can write the relationship between the known and the unknown chemicals using ratios:

$$\frac{n(\text{unknown chemical})}{n(\text{known chemical})} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$

Worked example 9.4.2

USING MOLE RATIOS

Solutions of copper(II) sulfate and sodium hydroxide are mixed together, forming 0.40 mol of solid copper(II) hydroxide. Calculate how many moles of sodium hydroxide reacted in this reaction.

Thinking	**Working**
Write a balanced equation for the reaction.	$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
Determine the number of moles of the ‘known substance’. The known substance is the one you are provided information about in the question.	$n(Cu(OH)_2) = 0.40\text{ mol}$
Find the mole ratio: $\text{mole ratio} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$ The ‘unknown’ substance is the one whose number of moles you are required to calculate.	$\frac{n(NaOH)}{n(Cu(OH)_2)} = \frac{2}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = n(\text{known}) \times (\text{mole ratio})$	$n(NaOH) = 0.40 \times \frac{2}{1}$ $= 0.80$ moles

➤ Try yourself 9.4.2

USING MOLE RATIOS

Solutions of lead(II) nitrate and potassium iodide are mixed together, forming 0.30 mol of solid lead(II) iodide. Calculate how many moles of potassium iodide reacted in this reaction.

MASS–MASS STOICHIOMETRY

When you carry out a reaction in the laboratory, you will typically measure quantities of chemicals in grams, not moles. For this reason, most calculations will require you to start and finish with mass rather than moles of a substance. As you learnt in Module 9.2, the number of moles of a substance can be calculated from a given mass using the relationship:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1}\text{)}}$$

This can be written as:

$$n = \frac{m}{M}$$

To calculate a final answer as a mass, this formula is rearranged:

$$m = n \times M$$

There are several steps involved in calculating the mass of an unknown chemical.

1 Write a balanced equation for the reaction.
2 Calculate the number of moles using the mass of the known substance and the formula $n = \frac{m}{M}$.
3 Use the mole ratios in the equation to calculate the number of moles of the second substance.
4 Calculate the mass of the unknown substance using $m = n \times M$.

Figure 9.4.4 provides a flow chart that summarises this process. Worked Example 9.4.3 will help you to understand these steps.

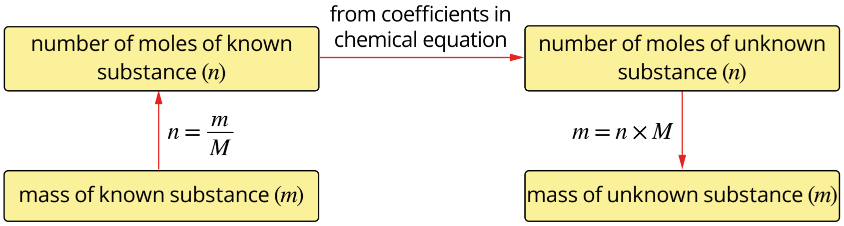


FIGURE 9.4.4 A flow chart for mass–mass stoichiometric calculations is helpful when trying to solve these problems.

Worked example 9.4.3

MASS–MASS STOICHIOMETRIC PROBLEMS

Solid magnesium burns in the presence of oxygen gas to form magnesium oxide. Calculate the mass of oxygen that would need to burn to produce 24.0 g of magnesium oxide.

Thinking	Working
Write a balanced equation for the reaction.	$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
Calculate the number of moles of the known substance (the precipitate): $n = \frac{\text{mass } (m)}{\text{molar mass } (M)}$	$n(MgO) = \frac{24.0\text{ g}}{40.31\text{ g}}$ $= 0.60\text{ mol}$
Calculate the mole ratio: $\text{mole ratio} = \frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	Mole ratio = $\frac{1}{2}$
Calculate the number of moles of the known substance: *n*(unknown) = *n*(known) × mole ratio	*n*(unknown) = *n*(MgO) × mole ratio $= 0.60 \times \frac{1}{2} = 0.30\text{ mol}$
Calculate the mass of the unknown substance: *m* = *n*(unknown) × molar mass	$m(O_2) = 0.30 \times 32.0 = 9.6\text{ g}$

➤ Try yourself 9.4.3

MASS–MASS STOICHIOMETRIC PROBLEMS

A reaction between solutions of sodium sulfate and barium nitrate produces a precipitate of barium sulfate with a mass of 2.440 g. Calculate the mass of sodium sulfate required to produce this precipitate.

LIMITING REAGENT

When two reactants are mixed in a chemical reaction, it is possible that they will be combined in just the right mole ratio as indicated in the equation for each to react completely. However, it is more likely that they are not present in exactly the right mole ratio. In this case, one of the reactants will be used up before the other and some of the other reactant will be left over at the end of the reaction.

To illustrate this situation simply, first work on a problem that does not involve chemicals. Suppose that you have been given some skateboard decks and wheels and you want to make as many complete skateboards as you can. As you can see in Figure 9.4.5, a complete skateboard is made up of one deck and four wheels.

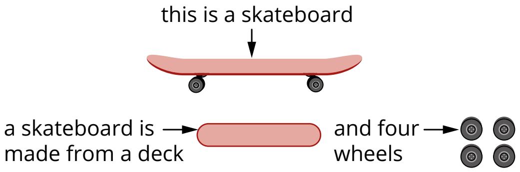


FIGURE 9.4.5 In order to make a complete skateboard, you must always use one deck and four wheels. This sets up the basic formula for a skateboard.

Now consider the situation shown in Figure 9.4.6. If you were given two decks and ten wheels, how many complete skateboards could you make from these materials?

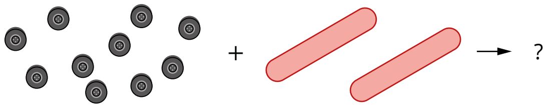


FIGURE 9.4.6 When provided with ten wheels and two skateboard decks, how many complete skateboards can be made?

The answer is that you could make two complete skateboards and there would be two wheels left over (Figure 9.4.7).

FIGURE 9.4.7 When supplied with two decks and ten wheels the maximum number of skateboards that can be made is two. There will be two wheels that are not used.

In this example, you can say that the two wheels left at the end were in excess. Also, with the number of decks and wheels that you were given, the total number of complete skateboards you could make was limited by the number of decks available.

A similar situation arises in chemical reactions if the quantities of reactants supplied are not in the exact same ratio as that shown in the equation for the reaction.

When this happens, the reactant that is:

- completely consumed is the **limiting reactant**
- not completely consumed is the **excess reactant**.

Figure 9.4.8 shows three different scenarios for the reaction in which hydrogen gas and oxygen gas combine to form water, according to the equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

ℹ In a chemical reaction, the limiting reactant is completely used up.

Each of the diagrams provides examples to illustrate the concepts of limiting and excess reactants.

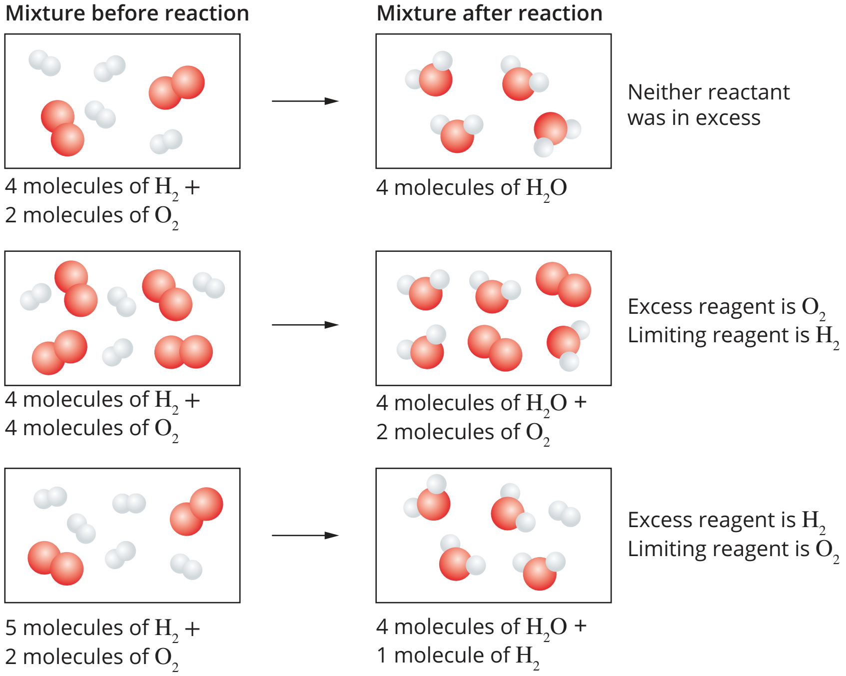


FIGURE 9.4.8 Different scenarios showing the concept of a limiting reactant for the reaction $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Note that in each of the examples shown in Figure 9.4.8, the amount of product:

- formed in these types of reactions is determined by the amount of the limiting reactant present in the reaction mixture
- cannot be determined from the amount of excess reactant.

In the skateboard example, it was the number of decks, not the number of wheels, which determined how many complete skateboards could be made. You must always use the amount of the limiting reactant to determine the amount of product that will be formed.

When attempting to solve a limiting reactant problem in which you are required to work out the amount of product, there are three main steps.

1 Calculate the number of moles of each reactant.
2 Determine which reactant is in excess and, therefore, which is the limiting reactant.
3 Use the amount of limiting reactant to work out the amount of product formed.

Worked example 9.4.4

MASS–MASS STOICHIOMETRY CALCULATION WITH ONE REACTANT IN EXCESS

A solution containing 20.0 g of dissolved sodium hydroxide is added to a solution containing 25.0 g aluminium nitrate. The balanced equation for the reaction is:

$$3NaOH(aq) + Al(NO_3)_3(aq) \rightarrow Al(OH)_3(s) + 3NaNO_3(aq)$$

a Determine which reactant is the limiting reactant.

Thinking	**Working**
Calculate the number of moles of each of the reactants in the equation using $n = \frac{m}{M}$.	Use the equation $n = \frac{m}{M}$. For NaOH: $n(NaOH) = \frac{20.0}{39.98} = 0.500\,mol$ For $Al(NO_3)_3$: $n(Al(NO_3)_3 = \frac{25.0}{213.01} = 0.117\,mol$
Use the coefficients of the equation to find the limiting reactant.	The equation shows that 3 mol of NaOH reacts with 1 mol of $Al(NO_3)_3$. So to react all of the $Al(NO_3)$ you will require $\frac{3}{1} \times n\left(Al(NO_3)_3\right)$ of NaOH $\frac{3}{1} \times 0.117 = 0.351\,mol$ As there is 0.500 mol available, the NaOH is in excess. So, $Al(NO_3)_3$ is the limiting reactant (it will be completely consumed).

b Determine the mass of precipitate formed.

Thinking	**Working**
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients: $\text{mole ratio} = \frac{\text{coefficient of unknown}}{\text{coefficient of limiting reactant}}$	From the equation coefficients: $\frac{\text{coefficient of } Al(OH)_3}{\text{coefficient of } Al(NO_3)_3} = \frac{1}{1}$
Calculate the number of moles of the unknown substance using the number of moles of limiting reactant: n(unknown) = n(limiting reactant) × mole ratio	$n(Al(OH)_3) = n(Al(NO_3)_3) \times \frac{1}{1}$ $= 0.117 \times \frac{1}{1}$ $= 0.117$ mol
Calculate the mass of the unknown substance using: m(unknown) = n(unknown) × molar mass	Molar mass of $Al(OH)_3$ = 78.01 g mol^{-1} $m(Al(OH)_3) = 0.117 \times 78.01$ $= 9.13\,g$

➤ Try yourself 9.4.4

MASS–MASS STOICHIOMETRY CALCULATION WITH ONE REACTANT IN EXCESS

WS 1.3.8

A solution containing 0.600 g of lead(II) nitrate is added to a solution containing 2.60 g of potassium iodide. The balanced equation for the reaction is:

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

a Determine which reactant is the limiting reactant.

b Determine the mass of precipitate formed.

YIELD

A major consideration in industrial chemistry is the amount of product that can be produced from given amounts of reactants in a given time. Many industrial processes involve a number of steps to make the final product. At each step, the conversion from reactants to products is usually less than complete. For every step in a reaction pathway, the amount of final product diminishes. Industrial chemists must consider the efficiency of a reaction pathway and the wastes that are produced (Figure 9.4.9).

FIGURE 9.4.9 Most chemical reactions carried out in industrial processes are not 100% efficient and so waste chemicals are produced. The reduction or elimination of waste chemicals is a major concern for industrial chemists.

The mass of product that can be formed if all reactants react to produce products according to the reaction equation is known as the **theoretical yield**. The theoretical yield is calculated using the mole ratios of the equation and assumes 100% conversion of the reactants. When reactants are mixed together in the correct mole ratio, the amount of products will not always be exactly as predicted from stoichiometric calculations.

A number of factors can influence the amount of product that will be produced for a given reaction:

- Some reactions do not go to completion. They reach a state of equilibrium, which will be covered in detail in Unit 3. When a reaction reaches equilibrium rather than continuing on to completion, the **actual yield** will be less than the theoretical yield.
- If the reaction rate is slow, the reaction may not proceed to completion in the time available. This will reduce the actual yield so that the theoretical yield is not obtained.
- Loss of reactants and products during transfers between reaction vessels and in separation and purification stages such as filtration will all result in smaller amounts of the product being obtained than expected.

ℹ Theoretical yield is the maximum amount of product that can be formed using stoichiometric ratios of the limiting reactant and assumes 100% conversion.

PERCENTAGE YIELD

The **percentage yield** compares the actual yield to the theoretical yield. It is a measure of the efficiency of a production process, for the particular conditions and method used for the synthesis. The higher the value of the percentage yield, the greater the degree of conversion from reactants to products for the reaction.

Percentage yield can be calculated using the formula:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

Worked example 9.4.5

CALCULATING THE PERCENTAGE YIELD OF A REACTION

Aspirin ($C_9H_8O_4$) is made when ethanoic anhydride ($C_4H_6O_3$) reacts with 2-hydroxybenzoic acid ($C_7H_6O_3$). A 12.50 g sample of $C_4H_6O_3$ is reacted with excess 2-hydroxybenzoic acid in the lab to produce 37.80 g of aspirin. Calculate the percentage yield of this reaction.

Thinking	Working
Write an equation for the reaction.	$2C_7H_6O_3 + C_4H_6O_3 \rightarrow 2C_9H_8O_4 + H_2O$
Use the formula $n = \frac{m}{M}$ to determine the amount of reactant.	$n(C_4H_6O_3) = \frac{12.50}{102.10} = 0.1224\text{ mol}$
Use the mole ratio for the reaction to determine the number of moles of the product that would be made if all of the reactant reacted.	$\text{mole ratio} = \frac{\text{coefficient of unknown}}{\text{coefficient of known}} = \frac{2}{1}$ $n(C_9H_8O_4) = \frac{2}{1} \times n(C_4H_6O_3) = 0.2448\text{ mol}$
Use the formula $m = n \times M$ to determine the mass of the product if all of the reactant reacts. This is the theoretical yield of the product.	$m(C_9H_8O_4) = n \times M = 0.2448 \times 180.17$ $= 44.11\text{ g}$
Calculate the percentage yield for this reaction from the formula: $\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times \frac{100}{1}$	$\text{percentage yield} = \frac{37.80}{44.11} \times \frac{100}{1} = 85.69\%$

➤ Try yourself 9.4.5

CALCULATING THE PERCENTAGE YIELD OF A REACTION

An 80.0 g sample of potassium is reacted with excess chlorine gas to produce 120.2 g of potassium chloride. Calculate the percentage yield of this reaction.

9.4 Review

SUMMARY

- Chemical equations are visual representations that chemists use to describe chemical reactions.
- The total mass of all products is equal to the total mass of all reactants in a chemical reaction.
- To write a balanced chemical equation, use coefficients so that the number of each atom on the reactant side equals the product side.
- A balanced equation shows the ratio of the number of moles of reactants used and products formed in the reaction.
- Given the quantity of one of the reactants or products of a chemical reaction, the quantity of all other reactants and products can be predicted by working through the following steps.
  1. Write a balanced equation for the reaction.
  2. Calculate the number of moles of the given substance.
  3. Use the mole ratios of reactants and products in the balanced equation to calculate the number of moles of the required substance.
  4. Use the appropriate formula to determine the required quantities of the required substance. The formula is:

$$m = n \times M$$

- The limiting reactant is the reactant that is consumed completely.
- The limiting reactant is used to predict the amount of product formed and the amount of the other reactant in excess.
- The theoretical yield of a chemical reaction is the mass of the product that would be formed if the limiting reactant reacted completely.
- To calculate the percentage yield, divide the actual yield obtained by the theoretical yield that would be obtained if all of the limiting reactant reacted completely, and multiply by 100:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

KEY QUESTIONS

Retrieval

1 Define the term 'chemical equation'.

2 Identify if the following reaction is balanced:

$$Zn(OH)_2 + H_3PO_4 \rightarrow Zn_3(PO_4)_2 + H_2O$$

3 Define the term 'stoichiometry'.

4 Identify if the following statement is true or false. The reactant that is completely consumed in a chemical reaction is the excess reagent.

5 State the formula for calculating percentage yield.

6 State three reasons why actual yield varies from theoretical yield.

Comprehension

7 Explain why when chemical equations are written, they must be balanced.

8 Represent the following reactions with balanced equations.

a Nitrogen and oxygen react in the cylinders of car engines to produce nitrogen monoxide gas.

b Heating tin(IV) hydroxide powder gives solid tin(IV) oxide and water.

c Solid magnesium reacts with aqueous silver nitrate to form solid silver and aqueous magnesium nitrate.

9 Explain how stoichiometric calculations use the mole ratio.

10 Explain how percentage yield is calculated.

Analysis

11 A 5.64 g precipitate of calcium phosphate is obtained when a solution of sodium phosphate is added to a solution of calcium nitrate. The balanced equation for the reaction is:

$$3Ca(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NaNO_3(aq)$$

Calculate the mass of calcium nitrate required to produce this precipitate.

9.4 Review *continued*

12 In three different experiments, different amounts of nitrogen and hydrogen reacted to form ammonia, according to the equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

This table shows the amount of reactants and products in each experiment. Determine the amount of each product remaining at the end of the reaction.

Nitrogen molecules available	Hydrogen molecules available	Ammonia molecules produced	Nitrogen molecules in excess	Hydrogen molecules in excess
2	10			
879	477			
9 mol	6 mol			

13 Sodium metal can react with chlorine gas to form sodium chloride. A 25.0 g piece of sodium is reacted with 50.0 g of chlorine gas.

a Write the balanced equation for the reaction between sodium and chlorine.

b Calculate the mass of sodium chloride that will form in the reaction.

14 Calculate the percentage yield for the reaction in which 20.0 g of ethanol is oxidised to produce 21.5 g of ethanoic acid according to the equation:

$$C_2H_5OH \xrightarrow{H^+/Cr_2O_7^{2-}} CH_3COOH$$

15 During a trial experiment for the industrial production of ethanol from ethene, 150 g of ethene was reacted with water to produce ethanol according to the balanced equation:

$$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$$

Calculate the theoretical yield, in grams, of ethanol for this reaction.

16 The mass of aluminium nitrate in a solution is determined by adding sodium carbonate solution to precipitate the aluminium as aluminium carbonate. The equation for the reaction occurring is:

$$2Al(NO_3)_3(aq) + 3Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + 6NaNO_3(aq)$$

In a particular reaction, 4.68 g of precipitate is obtained.

a Calculate the moles of aluminium carbonate produced.

b Determine the required mole ratio for the reaction.

c Calculate the mass of aluminium nitrate that reacted.

17 Potassium iodide and lead(II) nitrate solutions react together to form a precipitate of lead(II) iodide:

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

In each of the following cases, calculate the quantities required.

a If 1.0 mol of potassium iodide reacts with 1.0 mol of lead(II) nitrate, determine which reactant is in excess and by how many moles.

b If 0.50 mol of potassium iodide reacts with 2.0 mol of lead(II) nitrate, determine which reactant is in excess and by how many moles.

c If 1.00 g of lead(II) nitrate reacts with 1.50 g of potassium iodide, determine which reactant is in excess and the mass of lead(II) iodide that forms.

MANDATORY PRACTICAL 1

Derive the empirical formula of a compound from reactions involving mass changes

Research and planning

Aim

To derive the empirical formula of magnesium oxide through experimentation

To calculate the percentage by mass of magnesium in magnesium oxide

Rationale (scientific background to the experiment)

By definition a compound consists of elements chemically combined in fixed proportions by mass. When magnesium is heated in oxygen, the compound magnesium oxide forms. By finding the mass of the original magnesium and that of the magnesium oxide, using the law of conservation of mass, the mole ratio of magnesium to oxygen can be determined so that the empirical formula can be derived. The percentage composition of magnesium in magnesium oxide can also be calculated.

Timing

50 minutes

Materials

- magnesium ribbon 20–40 cm long
- steel wool
- crucible and lid
- pipeclay triangle
- Bunsen burner
- bench mat
- tripod
- tongs
- electronic balance
- safety glasses

PRE-LAB SAFETY INFORMATION

Material used	Hazard	Control
crucible	gets hot when heated	Use tongs to handle crucible when hot.
magnesium strips	creates a bright light when reacting with oxygen	Do not stare directly at magnesium when heating it.

Please indicate that you have understood the information in the safety table.

Name (print): ______________________________________

I understand the safety information (signature): ____________

Method

Risk assessment

Consideration of risks include chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and Assessment book or download from your eBook.

1. Your teacher will give you a 20–40 cm length of magnesium ribbon. Clean the magnesium ribbon thoroughly using steel wool.
2. Weigh the mass of a clean, dry crucible and lid. Place the magnesium loosely in the crucible. Weigh the total mass of crucible, lid and magnesium. Record your data in Table 9.5.1 on page 258.
3. Set up your experiment as shown in Figure 9.5.1.

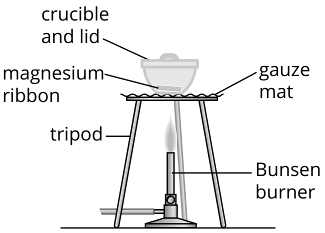


FIGURE 9.5.1 Experimental set-up

4. Heat the crucible strongly. Using tongs, occasionally lift the lid to allow air to enter the crucible but replace the lid quickly to avoid loss of magnesium oxide ash.
5. When the reaction appears complete, allow the equipment to cool to room temperature. Weigh the total mass of crucible, lid and magnesium oxide.
6. Calculate the original mass of magnesium used and the mass of magnesium oxide formed. Record the results in Table 9.5.1.
7. Assuming all the original magnesium is now in the form of magnesium oxide, calculate the percentage of magnesium in magnesium oxide. Record the result in Table 9.5.2. (Round off your result to the nearest per cent.)
8. Calculate:
   - **a** the mass of oxygen that reacted with the magnesium
   - **b** the number of moles of magnesium reacted
   - **c** the number of moles of oxygen reacted
   - **d** the simplest whole number mole ratio of magnesium to oxygen in magnesium oxide
   - **e** the empirical formula of magnesium oxide

Variables

i Independent: the mass of magnesium
ii Dependent: the mass of magnesium oxide formed
iii Controlled: the crucible mass, the equipment and balance used

Analysing

Raw data

1 Record your data in Table 9.5.1.

TABLE 9.5.1 Mass results

Material	**Mass (g)**
empty crucible and lid (1)	
crucible, lid and magnesium (2)	
crucible, lid and magnesium oxide (3)	

Processed data

2 Calculate the mass of magnesium that reacted.
3 Calculate the mass of magnesium oxide that formed.

Reflect and check that your data analysis demonstrates these characteristics

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.
☐ Accurate application of algorithms, visual and graphical representations of data is demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

Analysis

4 Calculate the percentage by mass of magnesium in magnesium oxide. Record your results in Table 9.5.2.

5 a Find out the result for percentage by mass obtained by three other groups that had different starting masses of magnesium. Record the results in Table 9.5.2.

TABLE 9.5.2 Group results

Group	**Starting mass of magnesium (g)**	**Percentage by mass of magnesium (%)**
your group		
other group 1		
other group 2		
other group 3		

b Explain why these results should be similar for each group, despite each group using a different mass of magnesium at the beginning of the experiment.

c If a group did not obtain a result similar to that of the other groups, suggest a reason for this occurring.

6 Calculate:
a the number of moles of magnesium that reacted
b the mass of oxygen, in grams, in your sample of magnesium oxide
c the number of moles of oxygen that reacted
d the ratio of moles of magnesium and oxygen

7 Using the result you obtained in Question **6d**, derive the empirical formula of magnesium oxide.

8 Determine how many atoms of magnesium were actually involved in your reaction. Use Avogadro's number to determine this.

9 a List each piece of equipment used in this experiment that had error associated with it. State the error value associated with each.
b Explain how this value may have had an effect on the determination of the empirical formula of magnesium oxide.

10 a Collect and process the data from your experiment and the results of three other groups. What is the most scientifically appropriate way to display the comparison between the initial mass of the magnesium used and the four mole ratios of magnesium : oxygen that were obtained, so that a clear and valid trend can be seen?
b Create this display and explain what you observe.
c Assess if there any points that do not fit the trend. Explain why this might be so.

Reflect and check that your analysis demonstrates these characteristics

☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.
☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.
☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Interpreting and communicating

Conclusion

1 **a** State your conclusion.

b State the evidence that you have collected that leads you to this conclusion.

Evaluation

2 **a** Considering your analysis and conclusion, decide if the experiment provided an effective and efficient method of determining the empirical formula of magnesium oxide.

b Judge if the level of uncertainty that you calculated in Analysis Question **9** was reasonable. (Provide values that were close to whole number ratios.)

Improvements

3 If you were to repeat the experiment, identify the steps that you would do differently. You should include these points in your answer.

a Explain how you would change the methodology and how this might improve the results.

b Consider how well you performed the tasks and the skills that you need to improve on in your technique.

c Explain how the collection of data could be improved or uncertainty reduced.

Extension

4 **a** Evaluate whether this methodology could be applied to determine the empirical formula of any metal oxide.

b Identify any limitations and possible problems that might restrict the use of this methodology to determine the empirical formula of other metal oxides.

Reflect and check that your evaluation demonstrates these characteristics

☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis).

☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respect to the error analysis and limitations or sufficiency of the data.

☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment which are logically derived from the analysis of the evidence.

Enhancement questions

5 The percentage by mass of aluminium in alumina (aluminium oxide) is 52.9%. Calculate the mass of aluminium that could theoretically be extracted from 800 tonnes of alumina.

6 A 10 g sample of a certain oxide of tellurium (Te) contains 8.0 g of tellurium. Find its empirical formula.

Chapter review

09

KEY TERMS

actual yield
atomic mass
atomic mass unit (amu)
Avogadro's constant
Avogadro's number
balanced equation
catalyst
chemical equation
chemical reaction
coefficient
decomposition
empirical formula
excess reactant
isotopic mass
law of conservation of mass
limiting reactant
mass spectrometry
mol
molar mass
mole
mole ratio
molecular formula
molecule
percentage composition
percentage yield
product
reactant
relative atomic mass
relative formula mass
relative molecular mass
stoichiometry
theoretical yield

KEY QUESTIONS

Retrieval

1 Select the correct definition for relative atomic mass from the options below.

- **A** The relative atomic mass is the mass of 1 mol of the compound. It is equal to the relative molecular or relative formula mass of the compound expressed in grams.
- **B** The relative atomic mass is the mass of one molecule of that substance relative to the mass of a ^{12}C atom taken as 12 units exactly.
- **C** The relative atomic mass is the weighted average of the atomic masses of isotopes of the element on the ^{12}C scale.
- **D** The relative atomic mass is the percentage abundance of an isotope in the natural environment.

2 Define the term 'molar mass'.

3 Define the term 'empirical formula'.

4 State the law of conservation of mass.

Comprehension

5 The standard on which all relative masses are based is the ^{12}C isotope, which is given a mass of 12 exactly. Explain why in the table of relative atomic masses in Appendix B at the end of the book, the relative atomic mass of carbon is listed as 12.01.

6 Write balanced equations, including state symbols if required, for the following reactions.

- **a** Methane gas reacts with oxygen gas to produce carbon dioxide gas and water vapour.
- **b** Iron metal reacts with oxygen to form solid iron(III) oxide.
- **c** Zinc hydroxide powder reacts with phosphoric acid solution to produce the precipitate zinc phosphate and water.

7 Explain why the actual yield of a chemical reaction is less than would be predicted from the balanced equation.

Analysis

8 When a sample of palladium is placed in a mass spectrometer, the following peaks are recorded as the isotopic masses and corresponding percentage abundances given in the table below.

Isotopic mass	Abundance (%)
101.9049	0.9600
103.9036	10.97
104.9046	22.23
105.9032	27.33
107.9039	26.71
109.9044	11.80

Calculate the relative atomic mass of palladium. Give your answer to 2 significant figures.

9 The table below gives isotopic composition data for argon and potassium.

Element	Atomic number	Atomic mass	Relative abundance (%)
argon	18	35.978	0.307
		37.974	0.060
		39.974	99.633
potassium	19	38.975	93.3
		39.976	0.011
		40.974	6.69

- **a** Determine the relative atomic masses of argon and potassium.
- **b** Explain why the relative atomic mass of argon is greater than that of potassium, even though potassium has a higher atomic number.

10 For each of the following numbers of molecules, calculate the number of moles of the following quantities.

- **a** 4.50×10^{23} molecules of water (H_2O)
- **b** 9.00×10^{24} molecules of methane (CH_4)
- **c** 2.3×10^{28} molecules of chlorine (Cl_2)
- **d** 1 molecule of sucrose ($C_{12}H_{22}O_{11}$)

11 For each of the following amounts of molecular substances, calculate the:

- **i** number of molecules
- **ii** total number of atoms

- **a** 1.45 mol of ammonia (NH_3)
- **b** 0.576 mol of hydrogen sulfide (H_2S)
- **c** 0.0153 mol of hydrogen nitrate (HNO_3)

12 Calculate the mass of each of the following.

- **a** 0.060 mol of ethane (C_2H_6)
- **b** 0.32 mol of glucose ($C_6H_{12}O_6$)
- **c** 6.8×10^{-3} mol of urea ($(NH_2)_2CO$)
- **d** 6.12 mol of copper atoms (Cu)

13 Calculate the number of moles for each of the following.

- **a** carbon atoms in 1.201 g carbon
- **b** sulfur molecules (S_8) in 10.0 g sulfur
- **c** methane molecules (CH_4) in 20.0 g methane
- **d** aspirin molecules ($C_6H_4(OCOCH_3)COOH$) in 300 mg aspirin

14 For each of the following molecular substances, calculate the:

- **i** number of moles of the substance
- **ii** number of molecules
- **iii** total number of atoms

- **a** 4.2 g of phosphorus (P_4)
- **b** 75.0 g of sulfur (S_8)
- **c** 0.32 g of hydrogen chloride (HCl)

15 Calculate the mass of iron (Fe) that would contain as many iron atoms as there are molecules in 20.0 g water (H_2O).

16 If 6.0×10^{22} molecules of a substance have a mass of 10 g, calculate the molar mass of the substance.

17 Calculate the molar mass of a substance if:

- **a** 2.0 mol of the substance has a mass of 80 g
- **b** 0.1 mol of the substance has a mass of 9.8 g
- **c** 1.7 mol of the substance has a mass of 74.8 g
- **d** 3.5 mol of the substance has a mass of 371 g

18 A new antibiotic has been isolated and only 2.0 mg is available. The molar mass is found to be 12.5 kg mol^{-1}.

- **a** Express the molar mass in g mol^{-1}.
- **b** Calculate the number of moles of antibiotic.
- **c** Determine how many molecules of antibiotic have been isolated.

19 Determine the percentage by mass of carbon in the following substances.

- **a** naphthalene ($C_{10}H_8$)
- **b** ethanoic acid (CH_3COOH)
- **c** urea (NH_2CONH_2)
- **d** aspirin ($C_6H_4(OCOCH_3)COOH$)

20 Determine the empirical formulas of the compounds with the following compositions.

- **a** 27.2% carbon, 72.8% oxygen
- **b** 54.5% carbon, 9.1% hydrogen, 36.4% oxygen
- **c** 9.6 g carbon, 0.67 g hydrogen, 4.7 g chlorine
- **d** 4.42 g carbon, 0.842 g hydrogen

21 A clear liquid extracted from fermented lemons was found to consist of carbon, hydrogen and oxygen. Analysis showed it to be 52.2% carbon and 34.8% oxygen.

- **a** Find the empirical formula of the substance.
- **b** If 2.17 mol of the compound has a mass of 100 g, find the molecular formula of the compound.

22 Find the relative atomic mass of nickel if 3.370 g nickel was obtained by reduction of 4.286 g of the oxide (NiO).

23 A 3.72 g sample of element X exactly reacts with 4.80 g of oxygen to form a compound whose molecular formula is shown, from other experiments, to be X_4O_{10}. Determine the relative atomic mass of X.

24 When solutions of iron(II) sulfate and potassium hydroxide are mixed, the reaction that occurs can be represented as:

$$FeSO_4(aq) + 2KOH(aq) \rightarrow Fe(OH)_2(s) + K_2SO_4(aq)$$

Complete the following expressions based on this equation. The first one has been done for you.

- **a** $n(KOH) = \frac{2}{1} \times n(Fe(OH)_2)$
- **b** $n(FeSO_4) = \frac{(\ )}{(\ )} \times n(KOH)$
- **c** $n(KOH) = \frac{(\ )}{(\ )} \times n(K_2SO_4)$
- **d** $n(Fe(OH)_2) = \frac{(\ )}{(\ )} \times n(FeSO_4)$

25 Carbon disulfide is prepared by reacting coke with excess sulfur dioxide. The balanced equation for this reaction is:

$$5C(s) + 2SO_2(g) \rightarrow CS_2(l) + 4CO(g)$$

Determine how many grams of carbon monoxide will be produced if 200 g of carbon reacts with excess sulfur dioxide.

26 In three separate experiments, different amounts of carbon and oxygen were reacted together to form carbon dioxide, according to the equation:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

The table lists the amount of reactants and products in each of the three experiments. Complete the table to indicate the amount of product formed and the remaining reactant at the end of the reaction.

Carbon atoms available	Oxygen molecules available	Carbon dioxide molecules produced	Carbon atoms in excess	Oxygen molecules in excess
8	20			
1000	3000			
9 mol	6 mol			

27 Sodium can react with oxygen gas to form sodium oxide. The equation for this reaction is:

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

A 3.0 mol sample of sodium is reacted with 0.8 mol of oxygen gas.

a Determine which reactant is in excess.

b Determine how many moles of sodium oxide is produced in the reaction.

28 When 4.40 g of P_4O_6 is mixed with 3.00 g of I_2, they react according to this equation:

$$5P_4O_6(s) + 8I_2(g) \rightarrow 4P_2I_4(s) + 3P_4O_{10}(s)$$

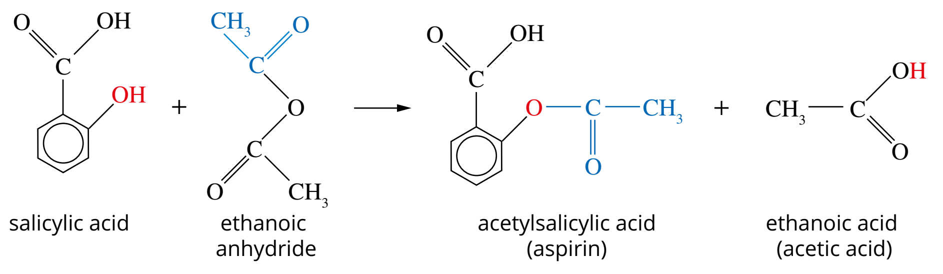
a Determine which reactant is in excess and by how much, in g.

b Determine the mass of P_2I_4 formed.

c Determine the mass of P_4O_{10} formed.

d Determine the total mass of all the products. (Hint: compare this with the mass of the reactants.)

29 Aspirin can be synthesised by an esterification reaction according to the pathway shown below.



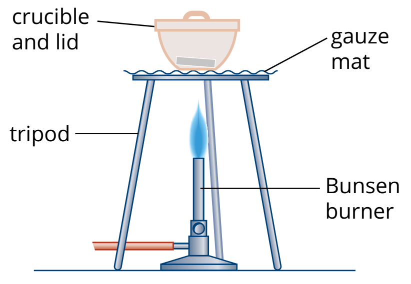
A student reacted a 2.50 g sample of salicylic acid with an excess of ethanoic anhydride, using sulfuric acid as a catalyst. After purification, a mass of 2.35 g of pure aspirin was obtained.

a Calculate the theoretical yield of aspirin for the reaction.

b Calculate the percentage yield of aspirin for the reaction.

Knowledge utilisation

30 The empirical formula of a metal oxide can be found by experimentation using the equipment set-up as shown.



The mass of the oxygen that reacts with the mass of the metal must be determined. Steps A–F form the experimental method.

a Sort the steps in the correct order by letter.

- **A** Ignite a burner and heat the metal.
- **B** Allow the crucible to cool, then weigh it.
- **C** Continue the reaction until no further change occurs.
- **D** Clean a piece of metal with emery paper to remove any oxide layer.
- **E** Place the metal in a clean, weighed crucible and cover with a lid.
- **F** Weigh the metal and record its mass.

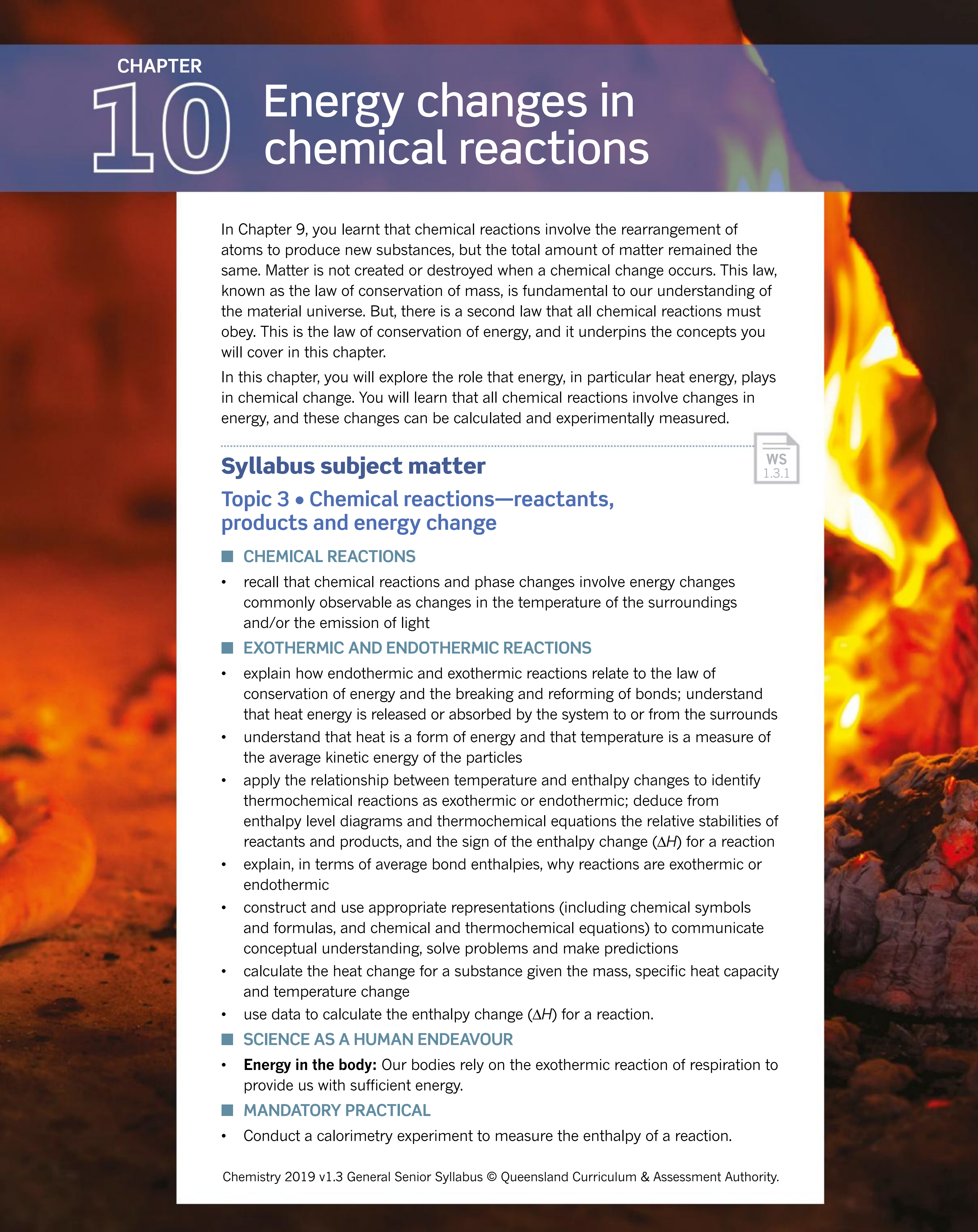
b Wan and Eric collected the following data:

mass of the metal = 0.542 g
mass of the empty crucible = 20.310 g
mass of the crucible and metal oxide = 21.068 g

They found from this data that the metal oxide had a 1 : 1 formula, i.e. MO, where M = metal. Copy and complete the table below, using the data given.

	Metal	Oxygen
mass (g)		
relative atomic mass		16.0
moles		
ratio		

c Determine the metal that was used in the experiment.



CHAPTER 10

Energy changes in chemical reactions

In Chapter 9, you learnt that chemical reactions involve the rearrangement of atoms to produce new substances, but the total amount of matter remained the same. Matter is not created or destroyed when a chemical change occurs. This law, known as the law of conservation of mass, is fundamental to our understanding of the material universe. But, there is a second law that all chemical reactions must obey. This is the law of conservation of energy, and it underpins the concepts you will cover in this chapter.

In this chapter, you will explore the role that energy, in particular heat energy, plays in chemical change. You will learn that all chemical reactions involve changes in energy, and these changes can be calculated and experimentally measured.

WS 1.3.1

Syllabus subject matter

Topic 3 • Chemical reactions—reactants, products and energy change

CHEMICAL REACTIONS

- recall that chemical reactions and phase changes involve energy changes commonly observable as changes in the temperature of the surroundings and/or the emission of light

EXOTHERMIC AND ENDOTHERMIC REACTIONS

- explain how endothermic and exothermic reactions relate to the law of conservation of energy and the breaking and reforming of bonds; understand that heat energy is released or absorbed by the system to or from the surrounds
- understand that heat is a form of energy and that temperature is a measure of the average kinetic energy of the particles
- apply the relationship between temperature and enthalpy changes to identify thermochemical reactions as exothermic or endothermic; deduce from enthalpy level diagrams and thermochemical equations the relative stabilities of reactants and products, and the sign of the enthalpy change (ΔH) for a reaction
- explain, in terms of average bond enthalpies, why reactions are exothermic or endothermic
- construct and use appropriate representations (including chemical symbols and formulas, and chemical and thermochemical equations) to communicate conceptual understanding, solve problems and make predictions
- calculate the heat change for a substance given the mass, specific heat capacity and temperature change
- use data to calculate the enthalpy change (ΔH) for a reaction.

SCIENCE AS A HUMAN ENDEAVOUR

- **Energy in the body:** Our bodies rely on the exothermic reaction of respiration to provide us with sufficient energy.

MANDATORY PRACTICAL

- Conduct a calorimetry experiment to measure the enthalpy of a reaction.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

10.1 Energy change

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the concept of energy and know that the SI unit for energy is the joule
- compare the concepts of system and surrounds
- understand that heat is a form of energy and that temperature is a measure of the average kinetic energy of particles
- recall that chemical changes and phase changes involve energy changes
- describe and name the phase changes: melting, freezing, vaporisation, condensation, sublimation and deposition.

Before the invention of agriculture, a person in an early hunter–gatherer society used approximately 10 MJ of energy per day. Despite being less active than hunter–gatherers, a person in our modern industrial society uses, on average, nearly 1000 MJ per day. This figure is 100 times greater than the body's basic requirement. The bulk of the energy used, directly or indirectly, comes from the combustion of fuels such as coal, natural gas or petrol. Coal is burnt in coal-fired power stations, like the one in Figure 10.1.1, where the heat released is used to generate electricity. Some activities use small amounts of energy, whereas others such as launching a space shuttle (Figure 10.1.2), use enormous amounts of energy.

FIGURE 10.1.2 The space shuttle *Endeavour* taking off from Kennedy Space Center in Cape Canaveral, Florida. The solid rocket boosters, which helped propel the shuttle into space, consumed 5000 kg of fuel per second, which releases the same amount of energy as 14 700 diesel-powered locomotives.

FIGURE 10.1.1 There are several coal-fired power stations in Queensland, like this one near Gladstone, which generate electricity by burning coal. The heat comes from the combustion of the coal that power turbines to generate electricity.

In this module, you will define what energy is, and the role it plays in chemical reactions and physical changes. You will look specifically at heat and temperature, and see how these concepts help us to understand changes of state.

CHEMICAL ENERGY

Chemical reactions occur when particles (atoms, molecules or ions) collide and are rearranged to form new particles. All chemical reactions involve energy changes. Unlike matter, you know the effects of energy without being able to directly see, touch, smell, or weigh it. **Energy** is usually defined as the capacity to do work, that is, the capacity to move an object against an opposing force. It comes in many forms, and includes heat, light, sound, electricity and chemical energy.

Chemical energy is the energy stored in the bonds between the atoms and molecules of a substance. It results from factors such as the attractions between electrons and protons in atoms, repulsions between nuclei, repulsions between electrons, movement of electrons, and the vibration and rotations around bonds.

Chemical energy is the energy stored in the bonds between the atoms and molecules of a substance.

During a chemical reaction, a **reactant** substance is transformed into a **product** substance. This new substance may have a different amount of chemical energy stored in its bonds, so the reaction will either absorb or release energy as it progresses from reactants to products. For example when you eat a meal (Figure 10.1.3), the bonds within the food molecules have stored energy that you can access to provide energy for other chemical and physical activities that take place in your body. When wood in Figure 10.1.4 is burned the chemical energy stored in the chemical bonds is released as heat and light. The study of energy changes that occur during chemical reactions is called **thermochemistry**.

FIGURE 10.1.3 When you eat food, you access the chemical energy stored in the food. This energy powers all of the chemical reactions and physical activities that take place in your body.

FIGURE 10.1.4 The combustion of wood in a fire releases energy in the form of heat and light.

The SI unit for energy is the **joule**, J. Other units also used for quantifying energy are kilojoules, kJ, and megajoules, MJ.

The relationship between joules, kilojoules and megajoules is:

$$1\,\text{J} = 10^{-3}\,\text{kJ} = 10^{-6}\,\text{MJ}$$

This can also be expressed as:

$$1\,\text{MJ} = 1000\,\text{kJ} = 1\,000\,000\,\text{J}$$

Figure 10.1.5 shows how you can convert between different units of energy.

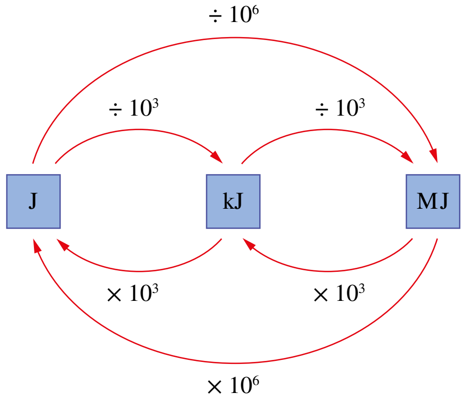


FIGURE 10.1.5 Converting between different energy units

SYSTEMS AND SURROUNDINGS

Energy changes in chemical reactions often refer to a **system** and its **surroundings**. In chemistry, the system is usually the chemical reaction. When energy is released or absorbed by a system, it means that energy changes occur as bonds are broken and formed between the atoms of the elements involved in the reaction.

The surroundings are usually regarded as everything else. For example, the walls of a container in which a reaction takes place in the gas phase, or the water in a solution in which a reaction takes place in the aqueous phase, can be regarded as the surroundings for the reaction. Energy leaves the system (the reaction) and enters the surroundings, or leaves the surroundings and enters the system.

HEAT AND TEMPERATURE

The reactants in a chemical reaction have a certain amount of chemical energy stored in their bonds. The products that form as a result of the rearrangement of particles during the chemical reaction have different bonds and so have a different amount of chemical energy. Energy will be released or absorbed during the reaction depending on the relative energies of the bonds within the reactants and products. When fuels, for example, undergo **combustion**, they release energy that can be used. This is what makes them fuels.

Often the energy released to or absorbed from the surroundings is in the form of **heat** (thermal energy). Heat is represented with the symbol Q, and is defined as the energy that flows from one object to another because of a difference in **temperature**.

Heat (Q) is the energy that flows from one object to another because of a difference in temperature.

If you add heat to an object you will increase its temperature, and heat flows spontaneously from warmer objects to cooler objects. If two objects remain in contact, heat will flow until the two objects are the same temperature.

While heat is the flow of energy, temperature is a measure of the average kinetic energy of the particles of an object. If the same amount of heat energy flows into two different objects, the temperature change will not be the same. One object may be smaller, with less particles, so it would experience a larger temperature increase (Figure 10.1.6). The increase in temperature of an object depends on the mass of the object, the amount of heat added, and the nature of the substance. Different substances, with different types of bonds between their atoms and molecules, have differing capacities to absorb heat.

FIGURE 10.1.6 (a) The water in the kettle has a higher temperature but the water in the swimming pool has more heat energy. (b) It would take more heat to raise the temperature of the swimming pool by 1°C than the kettle because there are more water molecules in the swimming pool.

Temperature is a measure of the average kinetic energy of the particles that make up an object. Temperature is measured with the Celsius (°C) or Kelvin (K) scale.

Temperature can be measured with a thermometer on the **Celsius** (°C) **scale**, where 0°C is the temperature at which water freezes at a pressure of 101.3 kPa. In chemistry, temperature is often measured on the **Kelvin scale** (K). The Kelvin scale is named after Lord Kelvin (Figure 10.1.7) who proposed the scale in 1848. Zero on the Kelvin scale is the temperature at which particles have no energy, and is equivalent to –273°C. Temperatures can be converted from Celsius to the Kelvin scale by the relationship:

$$T(\text{K}) = T(^\circ\text{C}) + 273$$

One degree Celsius is equal to one kelvin, so when changes in temperature are measured, they can equally be expressed in degrees C or K.

$$1^\circ\text{C} = 1\,\text{K}$$

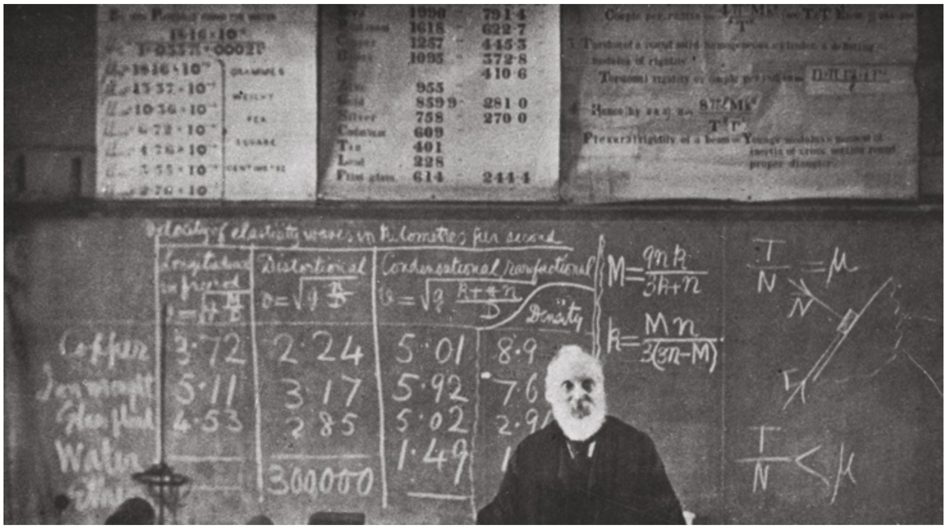


FIGURE 10.1.7 Kelvin (1824–1907) proposed the existence of absolute zero in 1848 and the use of an absolute temperature scale.

STATES OF MATTER

If you consider the example of the chemical water molecule (H_2O), three different words can be used to describe it—water, ice or steam. Imagine if somebody dropped a 25 kg block of ice on your head. It would have a very different effect to dropping 25 kg of water on your head! Methane, like most substances, can also exist in three different physical **states** (or **phases**) **of matter**, which are: solid, liquid and gas (Figure 10.1.8). Changes of state, such as a solid melting to form a liquid, are a type of physical change, rather than a chemical change. These changes also involve energy being absorbed or released.

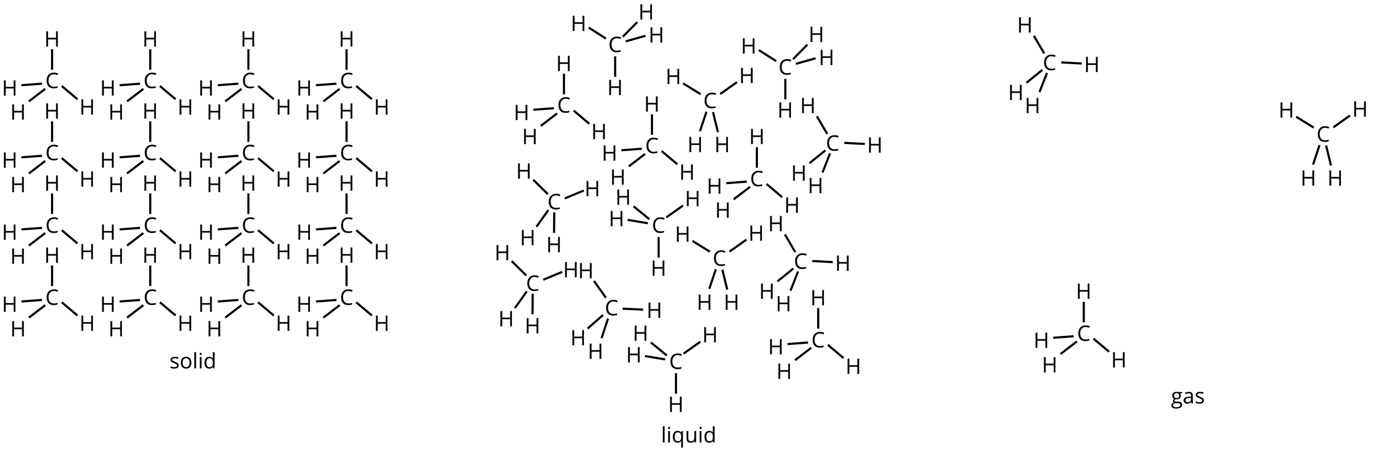


FIGURE 10.1.8 The chemical methane (CH_4) can exist as a solid, liquid or gas.

Matter in a **solid** state has a definite shape and volume, and the shape does not depend on the shape of the container. The particles in a solid are packed tightly together, often in an orderly arrangement. Due to the arrangement of their particles, solids are effectively incompressible.

Matter in a **liquid** state does not have a defined shape, and will take the shape of the container it is placed in. The particles in a liquid are in close contact with one another, but the arrangement is not fixed or orderly. The particles are free to move from one location to another, but the volume is fixed and like solids, liquids are almost incompressible.

There are three states of matter, which can be characterised in terms of the arrangement and movement of the particles and the forces between them: solid, liquid and gas.

Matter in a **gaseous** state takes the shape and volume of its container. This is because the particles in gases are no longer in contact with one another and are free to move. Compared to solids and liquids, the particles in **gases** are much further apart, so gases are much more easily compressed. Sometimes the word **vapour** is used interchangeably with the word ‘gas’.

Kinetic theory explains the differences in the physical properties of the different states of matter. Kinetic theory states that all matter consists of particles in motion, and as the temperature increases the movement of the particles increases. Most substances can exist in all three states, and the state at a given temperature and pressure is determined by the strength of the inter-particle forces.

Kinetic theory states that all matter consists of particles in motion, and as the temperature increases, the movement of the particles increases.

Heat flow results from the microscopic movement and collisions of particles (Figure 10.1.9). When heat moves between two systems at different temperatures, the kinetic energy of the warmer system's particles is transferred to the cooler system's particles. When the temperature increases enough, and the particles have sufficient energy to overcome inter-particle forces, a **change of state** occurs. When the vibrations of the particles of a solid have sufficient energy, they will move away from their fixed positions and take liquid form. This process is called **melting**, and the reverse process (liquid to solid) is called **freezing**. If the particles of a substance in a liquid state absorb enough energy, some may have sufficient energy to break away from the surface of the liquid. These particles are said to **evaporate** (Figure 10.1.10). When there is sufficient energy for all the particles to break apart, the substance is said to have reached **boiling point**. **Evaporation** and boiling are both types of **vaporisation**. The reverse of vaporisation (gas to liquid) is **condensation**.

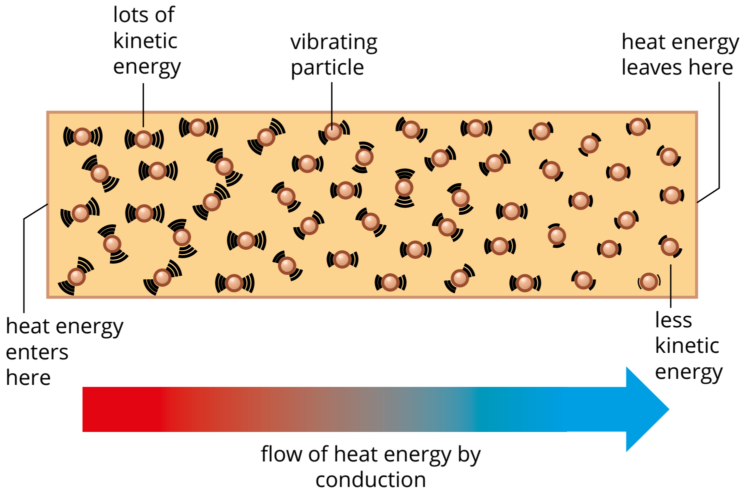


FIGURE 10.1.9 Heat flow, in this case by conduction, results from the microscopic movement and collisions of particles.

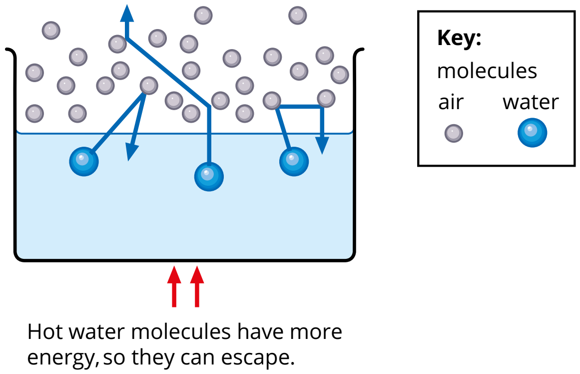


FIGURE 10.1.10 When water molecules have enough heat energy to overcome inter-particle forces, they will break away from the liquid. This process is called evaporation.

FIGURE 10.1.11 Dry ice, or carbon dioxide (CO_2) in its solid state, will sublime at room temperature.

It is possible, under certain temperatures and pressures, for solids to go directly to a gaseous state without passing through an intermediate liquid state. This type of state change is known as **sublimation**. In Figure 10.1.11, solid carbon dioxide (CO_2), or dry ice, is sublimating at room temperature. The reverse of sublimation is **deposition**. Figure 10.1.12 summarises the different changes of state.

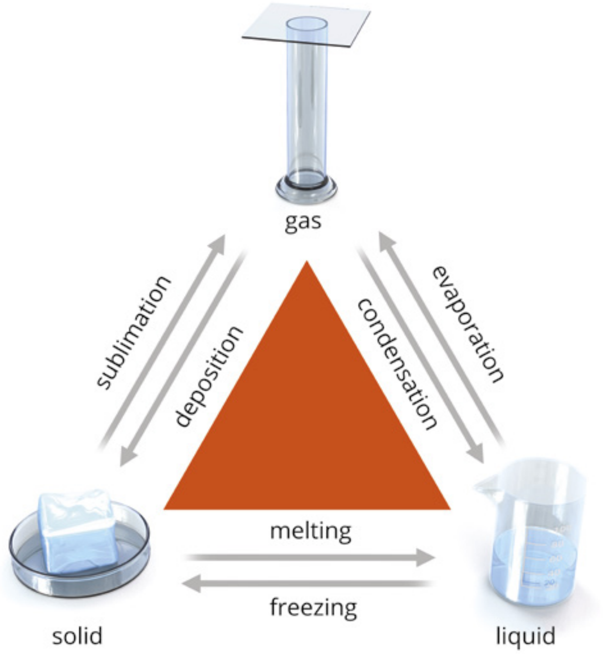


FIGURE 10.1.12 A summary of the possible state changes that can occur between solid, liquid and gas

Consider the heating curve for water below in Figure 10.1.13. You will notice that while the water is undergoing a change of state the temperature does not increase. This is because the heat energy being absorbed by the substance is being used to break the intermolecular bonds. As a result, the average kinetic energy or temperature of the substance does not increase. Once all the particles have changed state, and the intermolecular bonds have been overcome, the temperature starts increasing again.

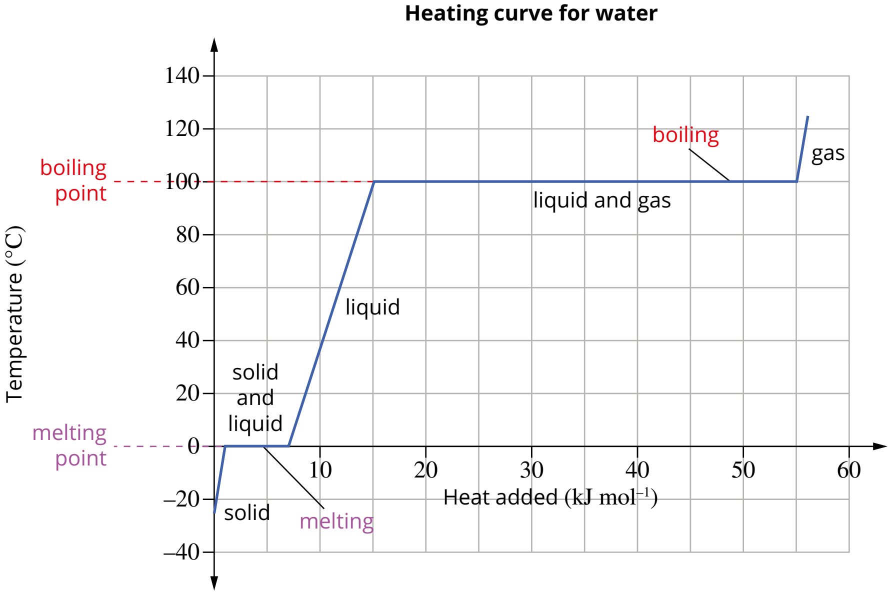


FIGURE 10.1.13 The heating curve for water. While the water is undergoing a change of state the temperature does not increase. Notice the state change from liquid to gas needs more energy than solid to liquid as the intermolecular bonds are broken during this process.

10.1 Review

SUMMARY

- Chemical energy is the energy stored in the bonds between the atoms and molecules of a substance. The SI unit for energy is the joule, J.
- The system is the chemical reaction and the surroundings are everything else.
- Heat (*q*), is defined as the energy that flows from one object to another because of a difference in temperature.
- Temperature is a measure of the average kinetic energy of the particles of an object. Temperature is measured on the Celsius (°C) or Kelvin (K) scale.
- Kinetic theory states that all matter consists of particles in motion, and as the temperature increases, the movement of the particles increases.
- Matter can exist in three different physical states (or phases): solid, liquid and gas.
- When the temperature increases enough for the particles to have sufficient energy to overcome inter-particle forces, a change of state occurs.

KEY QUESTIONS

Retrieval

1 Define chemical energy.

2 State the symbol used to represent heat.

3 Define sublimation.

Comprehension

4 Explain the difference between the terms 'system' and 'surroundings' in relation to a chemical reaction.

5 A student reacts an aqueous solution of an acid and an aqueous solution of a base together in a beaker. Describe the system in this scenario.

6 Explain how the Celsius and Kelvin temperature scales are similar and how they are different.

7 Explain evaporation in terms of kinetic theory.

Analysis

8 Compare the concepts of heat and temperature.

9 Consider the heating curve for water depicted in Figure 10.1.13 on page 269. Determine why the temperature does not rise immediately after melting and boiling begins.

10 Determine the following energy values as kJ.

a 0.180 MJ
b 1.5×10^6 J
c 10.0 J
d 2.0×10^{-3} J

11 Determine the following temperatures as K.

a 25.0°C
b –183.2°C
c 1450°C
d –272.8°C

12 Determine the following temperatures as °C.

a 45 K
b 199.2 K
c 312 K
d 455.5 K

10.2 Exothermic and endothermic reactions

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- state the law of conservation of energy and explain how it relates to endothermic and exothermic reactions
- describe exothermic reactions and endothermic reactions in terms of energy change
- understand the concept of enthalpy
- write thermochemical equations and use them to perform calculations
- use average bond enthalpy to explain whether a reaction is exothermic or endothermic
- interpret enthalpy level diagrams.

Chemical reactions always involve energy changes. As the reactant particles are rearranged, the chemical energy of the reactants will be different to that of the products. In some chemical reactions, for example the combustion of fuels, the rearrangement of atoms causes energy to be released to the surroundings. In other chemical reactions energy is absorbed from the surroundings as the chemical reaction takes place. The energy change in some reactions is very small and can only be measured with specialised equipment.

In this module, you will learn about the energy changes that occur during chemical reactions. You will also learn how to classify chemical reactions based on their energy changes.

FIGURE 10.2.1 The reaction between sodium and chlorine gas releases energy into the surroundings.

EXOTHERMIC AND ENDOTHERMIC PROCESSES

Chemical reactions and changes in physical state generally involve either an absorption or release of heat. Figure 10.2.1 shows the reaction between sodium and chlorine gas. In this reaction, a change in energy can be seen with the release of heat and light. The **law of conservation of energy** states, that during any chemical or physical process, energy is neither created nor destroyed. In thermochemistry, this means that if the energy of the system increases during a chemical reaction, the energy of the surroundings must decrease by the same amount.

When the total chemical energy of the products of a chemical reaction is less than the total chemical energy of the reactants, the excess energy is released to the surroundings. In thermochemical calculations, heat flow is described from the point of view of the system. If energy 'exits' the reaction system then the chemical reaction is called an **exothermic** reaction (Figure 10.2.2). The released energy can be shown in a chemical equation by writing 'energy' on the product side of the arrow.

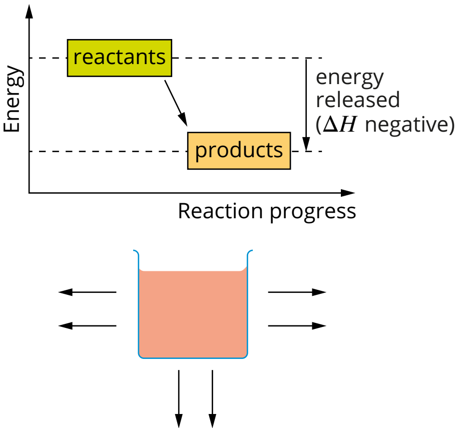


FIGURE 10.2.2 For an exothermic reaction, the enthalpy of the reactants is greater than the enthalpy of the products, so energy is released to the surroundings during the reaction.

For example, the production of water from the reaction between hydrogen and oxygen gas is an exothermic reaction. This can be represented by the balanced equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) + \text{energy}$$

Another example of an exothermic reaction is the combustion of methane gas:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) + \text{energy}$$

When the total chemical energy of the products of a chemical reaction is greater than the total chemical energy of the reactants, energy is absorbed from the surrounding environment. Energy 'enters' the reaction system and the chemical reaction is called an **endothermic** reaction (Figure 10.2.3). If an endothermic reaction takes place in a container, the container may feel cold to the touch. This is because the reaction system is absorbing heat from the surroundings, leaving the environment, including the container, cooler.

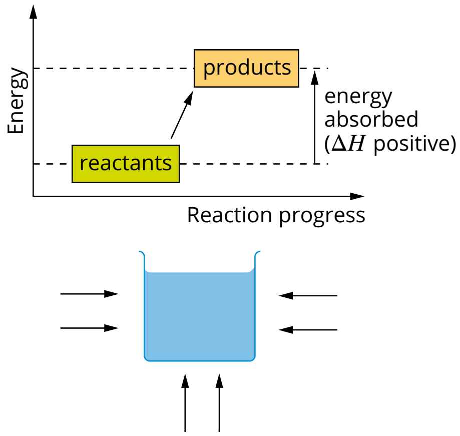


FIGURE 10.2.3 For an endothermic reaction, the enthalpy of the reactants is less than the enthalpy of the products so energy is absorbed during the reaction.

In a chemical equation of an endothermic reaction, the energy that is required can be written on the reactant side of the equation arrow. For example, the decomposition of calcium carbonate is an endothermic process. This can be represented by the balanced equation:

$$CaCO_3(s) + \text{energy} \rightarrow CaO(s) + CO_2(g)$$

Endothermic reactions require the constant input of energy.

Enthalpy

The chemical energy of a substance is sometimes called its **heat content** or **enthalpy**. It is given the symbol H. The enthalpy of the reactants in a chemical reaction is given the symbol H_r and the enthalpy of the products is given the symbol H_p.

Most chemical processes take place in open systems under a constant pressure (usually atmospheric). The exchange of heat energy between the system and its surroundings under constant pressure is referred to as the **enthalpy change**, or **heat of reaction**, and is given the symbol ΔH. The capital delta symbol (Δ) is commonly used in chemistry to represent 'change in'. For example, ΔT is the symbol for change in temperature.

For the general reaction, reactants → products, the enthalpy change (ΔH) is calculated by:

$$\Delta H = H_{(\text{products})} - H_{(\text{reactants})}$$

When $H_{(\text{products})}$ is less than $H_{(\text{reactants})}$, energy is released from the system into the surroundings, so the reaction is exothermic. The system has lost energy, so ΔH has a negative value. Therefore, for exothermic reactions, $\Delta H < 0$. When H_p is greater than H_r, energy must be absorbed from the surroundings, so the reaction is endothermic. The system has gained energy, so ΔH has a positive value, i.e. $\Delta H > 0$.

- If the total chemical energy of the products is less than the total energy of the reactants, energy will be released from the system into the surroundings. This is called an exothermic reaction.
- If the total chemical energy of the products is greater than the total energy of the reactants, energy will be absorbed from the surroundings. This is called an endothermic reaction.

Enthalpy change is a measure of the amount of energy absorbed or released during chemical reactions. It is given by the symbol ΔH and is determined by subtracting the enthalpy of the reactants from the enthalpy of the products.

THERMOCHEMICAL EQUATIONS

The enthalpy change can be shown by writing the ΔH value to the right of the chemical equation. Such an equation is called a **thermochemical equation**. The ΔH value in a thermochemical equation usually has the units kJ mol^{-1}. This means that the amount of energy (in kJ) signified by the ΔH value corresponds to the mole amounts specified by the coefficients in the equation.

For example, consider the cellular respiration of glucose:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \qquad \Delta H = -2803\,\text{kJ mol}^{-1}$$

This thermochemical equation tells you that when 1 mole of glucose reacts with 6 moles of oxygen to produce 6 moles of carbon dioxide and 6 moles of water, 2803 kJ of energy is released. The ΔH is negative because the reaction is exothermic.

Enthalpy changes also occur during physical changes, so thermochemical equations can be written for physical changes. Melting ice is an example of a physical change. It is an endothermic process, because heat must be applied to solid ice to convert it into liquid water. The thermochemical equation for this change is:

$$H_2O(s) \rightarrow H_2O(l) \qquad \Delta H = +6\,\text{kJ mol}^{-1}$$

The ΔH value is positive because the process is endothermic.

Thermochemical equations and mole ratios

The ΔH value in a thermochemical equation corresponds to the mole amounts specified by the equation. If the coefficients in the equation are changed, the ΔH value will also change.

For example, the balanced thermochemical equation for the combustion of methanol can be written as:

$$CH_3OH(l) + \tfrac{3}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -725\,kJ\,mol^{-1}$$

This means that 726 kJ of energy is released when 1 mole of methanol reacts with 1.5 moles of oxygen gas, to produce 1 mole of carbon dioxide and 2 moles of water. If twice as much methanol were to react, then twice as much energy would be released. So, if the coefficients of the equation are doubled, the ΔH value is also doubled:

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l) \qquad \Delta H = -1450\,kJ$$

If the mole amounts are tripled, the ΔH value is also tripled:

$$3CH_3OH(l) + \tfrac{9}{2}O_2(g) \rightarrow 3CO_2(g) + 6H_2O(l) \qquad \Delta H = -2175\,kJ$$

It is very important to always include **state symbols** in thermochemical equations. Physical changes involve an enthalpy change, so the state of the species in a chemical reaction affects the enthalpy change of the reaction. For example, both of the following equations represent physical changes involving water. They have different ΔH values because the states are different.

$$H_2O(s) \rightarrow H_2O(l) \qquad \Delta H = +6\,kJ\,mol^{-1}$$

$$H_2O(l) \rightarrow H_2O(g) \qquad \Delta H = +44\,kJ\,mol^{-1}$$

You can see that it requires more energy to boil water than it does to melt ice.

Reversing a chemical equation changes the sign but not the magnitude of ΔH. For example, methane (CH_4) reacts with oxygen gas to produce carbon dioxide gas and water in an exothermic reaction:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -890\,kJ\,mol^{-1}$$

If this reaction is reversed, the magnitude of ΔH remains the same because the enthalpies of the individual chemicals have not changed, but the sign changes to indicate that energy must be absorbed for this reaction to proceed.

$$CO_2(g) + 2H_2O(l) \rightarrow CH_4(g) + 2O_2(g) \qquad \Delta H = +890\,kJ\,mol^{-1}$$

Worked example 10.2.1

CALCULATING ΔH FOR A REACTION

Iron reacts with oxygen according to the equation:

$$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s) \qquad \Delta H = -1121\,kJ\,mol^{-1}$$

Calculate ΔH for the reaction represented by the equation:

$$2Fe_3O_4(s) \rightarrow 6Fe(s) + 4O_2(g)$$

Thinking	**Working**
The reaction has been reversed in the second equation, so the sign for ΔH is changed to the opposite sign.	ΔH for the second equation is positive.
Identify how the mole amounts in the equation have changed.	The mole amount of Fe_3O_4 has changed from 1 to 2, O_2 has changed from 2 to 4 and Fe has changed from 3 to 6. They have all doubled.
Identify how the magnitude of ΔH will have changed for the second equation.	The mole amounts of the chemicals have all doubled, so ΔH will also have doubled.
Calculate the new *magnitude* of ΔH. (You will write the sign of ΔH in the next step.)	2×1121 $= 2242$
Write ΔH for the second equation, including the sign.	$\Delta H = +2242\,kJ\,mol^{-1}$

➤ Try yourself 10.2.1

CALCULATING ∆H FOR A REACTION

Carbon reacts with hydrogen according to the equation:

$$6C(s) + 3H_2(g) \rightarrow C_6H_6(g) \qquad \Delta H = +49\,\text{kJ mol}^{-1}$$

Calculate ∆H for the reaction represented by the equation:

$$3C_6H_6(g) \rightarrow 18C(s) + 9H_2(g)$$

ENTHALPY LEVEL DIAGRAMS

ℹ A reaction cannot proceed until the bonds in the reactants are broken. The energy required to do this is the activation energy.

The energy required to break the bonds of reactants so that a reaction can proceed is called the **activation energy**. The activation energy is an energy barrier that must be overcome before a reaction can get started.

An activation energy barrier exists for both exothermic and endothermic reactions. If the activation energy for a reaction is very low, the chemical reaction can be initiated as soon as the reactants come into contact because the reactants already have sufficient energy for a reaction to take place. Once an exothermic reaction has started the energy released is often sufficient to overcome the activation energy and so the reaction continues. An example of this can be seen in the reaction between zinc and hydrochloric acid.

The reaction between zinc and hydrochloric acid produces hydrogen gas:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

As you can see in Figure 10.2.4, bubbles of hydrogen gas are vigorously produced as soon as zinc is added to the acid.

FIGURE 10.2.4 When zinc comes into contact with hydrochloric acid, it reacts almost immediately. The reactants have sufficient energy to 'overcome' the activation energy barrier.

The energy changes that occur during a chemical reaction can be shown on an **enthalpy level diagram**. The enthalpy level diagram for an exothermic combustion reaction, like the one shown in Figure 10.2.5, indicates that the enthalpy of the products is always less than the enthalpy of the reactants. Overall, energy is released and so the ∆H value is negative. The enthalpy level diagram also shows that, even in exothermic reactions, the activation energy must first be absorbed to start the reaction.

The enthalpy level diagram for an endothermic reaction (Figure 10.2.6) shows that the enthalpy of the products is greater than the enthalpy of the reactants. Overall, energy is absorbed and so the ∆H value is positive. The diagram also shows the absorption of the activation energy before the release of energy as bonds form in the products.

PA 1.3.1

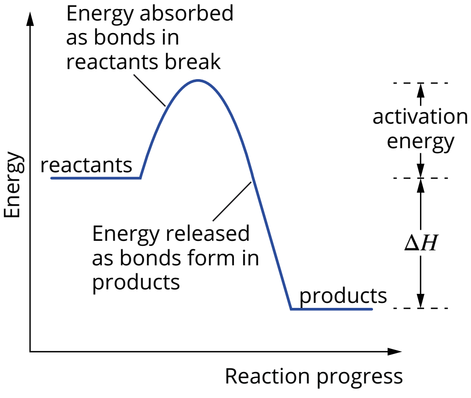


FIGURE 10.2.5 The characteristic shape of an enthalpy level diagram for an exothermic reaction

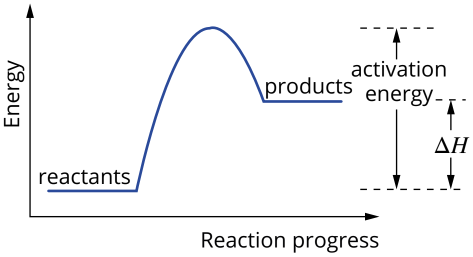


FIGURE 10.2.6 The characteristic shape of an enthalpy level diagram for an endothermic reaction

Predicting relative stability

The enthalpy level diagram and the sign (+ve or –ve) of the enthalpy change can be used to predict the comparative stability of the reactant and product chemicals. A reaction is exothermic overall if the bonds that are formed are stronger than the bonds that are broken. A reaction is endothermic when the broken bonds are stronger than the bonds being formed.

BOND ENTHALPIES

Chemical reactions involve making and breaking chemical bonds. The energy changes in a chemical reaction are due to the difference between the energy used to break the reactant bonds and the energy released when new bonds are formed in the products.

An ionic bond is a type of chemical bond that involves the electrostatic attraction between oppositely charged ions. Likewise, a covalent bond is due to the electrostatic attraction between the shared pair of electrons and the positive nuclei of the bonded atoms. Since energy is required to break bonds, the reaction is endothermic.

Forming bonds involves the bringing together of atoms that are attracted by an electrostatic force. Since this is the reverse reaction of breaking bonds, it stands to reason that the formation of bonds releases energy. That is bond formation is exothermic. The same amount of energy is absorbed when a bond is broken as is given out when the bond is made.

The energy needed to break one mole of bonds of gaseous molecules under standard conditions (25°C, 1 atm pressure) is known as the **bond enthalpy**. For example, 436 kJ of energy is required to break 1 mole of bonds in hydrogen molecules, forming 2 moles of hydrogen atoms. This process can be written as:

$$H_2(g) \rightarrow 2H(g) \qquad \Delta H = +436\,\text{kJ mol}^{-1}$$

Breaking chemical bonds is an endothermic process, and making chemical bonds is an exothermic process.

The bond enthalpies of diatomic molecules can be accurately measured. However, in larger molecules it is often difficult to successively break bonds one at a time to measure bond enthalpies. In these larger molecules, it is usual to quote the **average bond enthalpy**.

Consider the complete dissociation of methane:

$$CH_4(g) \rightarrow C(g) + 4H(g) \qquad \Delta H = +1662\ \text{kJ mol}^{-1}$$

This is made up of the four processes:

$$CH_4(g) \rightarrow CH_3(g) + H(g) \qquad \Delta H = +435\ \text{kJ mol}^{-1}$$
$$CH_3(g) \rightarrow CH_2(g) + H(g) \qquad \Delta H = +444\ \text{kJ mol}^{-1}$$
$$CH_2(g) \rightarrow CH(g) + H(g) \qquad \Delta H = +444\ \text{kJ mol}^{-1}$$
$$CH(g) \rightarrow C(g) + H(g) \qquad \Delta H = +339\ \text{kJ mol}^{-1}$$

These equations represent the breaking of a C–H bond. However, they are not all identical because the bond is being broken from a different reactant particle each time.

The average bond enthalpy for the C–H bond could be determined by averaging the above values.

$$\begin{aligned}\text{Average bond enthalpy for C–H bond} &= \frac{+435 + 444 + 444 + 339}{4}\\ &= \frac{1662}{4}\\ &= 415.5\,\text{kJ mol}^{-1}\end{aligned}$$

Table 10.2.1 shows the average bond enthalpies for various single and multiple bonds. The values in this table make the assumption that the bond enthalpy is the same in different molecules. As the values are derived from experimental observation, there may be slight variation in values between different publications.

TABLE 10.2.1 Average bond enthalpies (kJ mol^{-1})

Single bonds							
C–H	414	N–H	391	O–H	463	F–F	159
C–C	346	N–N	158	O–O	144		
C–N	286	N–O	214	O–F	191	Cl–F	255
C–O	358	N–FI	278	O–Cl	206	Cl–Cl	242
C–F	492	N–Cl	192	O–I	201		
C–Cl	324	N–Br	243				
C–Br	285			S–H	364	Br–F	249
C–I	228	H–H	436	S–F	327	Br–Cl	219
C–S	289	H–F	567	S–Cl	271	Br–Br	193
		H–Cl	431	S–Br	218		
Si–H	323	H–Br	366	S–S	266	I–Cl	211
Si–Si	226	H–I	298			I–Br	178
Si–C	307					I–I	151
Si–O	466						
Multiple bonds							
C=C	602	N=N	418	O_2	494		
C≡C	835	N≡N	942				
C=N	615			S=O	523		
C≡N	887			S=S	429		
C=O	799						
C≡O	1072						

Enthalpy changes associated with a chemical reaction can be understood in terms of the energy associated with making and breaking the chemical bonds during the reaction.

Consider the formation of hydrogen chloride:

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

In this reaction, the covalent bonds in hydrogen (H–H) and chlorine (Cl–Cl) are broken and two new covalent bonds (H–Cl, H–Cl) are formed.

Using data from Table 10.2.1, the enthalpy change of this reaction can be calculated. Note that bond breaking requires energy; the process is endothermic, so the sign attached to the bond enthalpy value is positive. Conversely, bond formation is an exothermic process, so the sign attached to the bond enthalpy will be negative.

		Average bond enthalpy (kJ mol^{-1})
bonds broken	1 × H–H	+436
	1 × Cl–Cl	+243
bonds formed	2 × H–Cl	2 × –431
enthalpy change		–183

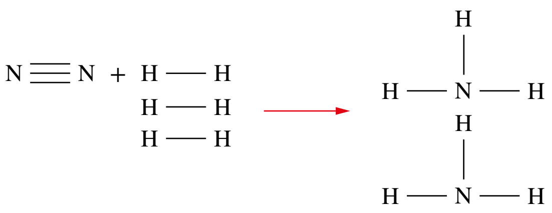
Therefore, the thermochemical equation for this reaction can be written as:

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g) \qquad \Delta H = -183 \text{ kJ}$$

The enthalpy change for a reaction is the overall difference in energy between the amount of energy absorbed to break bonds in the reactants and the amount of energy released when bonds are formed in making the products. This is given the following formula: $\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$

Worked example 10.2.2

CALCULATING THE ENTHALPY CHANGE OF A REACTION USING AVERAGE BOND ENTHALPIES

<table>
<tr><td colspan="2">Use data from Table 10.2.1 to determine if the reaction below is exothermic or endothermic and calculate the enthalpy change:
$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Draw molecules and identify the number and type of bonds broken and made.</td><td>
<table>
<tr><th>bonds broken</th><th>bonds formed</th></tr>
<tr><td>1 × N≡N</td><td>6 × N–H</td></tr>
<tr><td>3 × H–H</td><td></td></tr>
</table>

</td></tr>
<tr><td>Use Table 10.2.1 and record average bond enthalpy for each bond. Remember breaking bonds is endothermic, so average enthalpies will be +ve, while forming bonds is exothermic so average enthalpies will be –ve.</td><td>
<table>
<tr><th></th><th></th><th>Average bond enthalpy ($kJ\,mol^{-1}$)</th></tr>
<tr><td>bonds broken</td><td>1 × N≡N</td><td>+942</td></tr>
<tr><td></td><td>3 × H–H</td><td>3 × +436</td></tr>
<tr><td>bonds formed</td><td>6 × N–H</td><td>6 × –391</td></tr>
</table>
</td></tr>
<tr><td>Sum the values for bonds broken and bonds formed. Pay special attention to the signs of the bond enthalpy for each bond.</td><td>
<table>
<tr><th></th><th></th><th>Average bond enthalpy ($kJ\,mol^{-1}$)</th></tr>
<tr><td>bonds broken</td><td>1 × N≡N</td><td>+942</td></tr>
<tr><td></td><td>3 × H–H</td><td>+1308</td></tr>
<tr><td>bonds formed</td><td>6 × N–H</td><td>–2346</td></tr>
<tr><td>enthalpy change</td><td></td><td>–97</td></tr>
</table>
</td></tr>
<tr><td>Write Δ<i>H</i>, including the sign, and state the answer.</td><td>$\Delta H = -96\,kJ$
The reaction will be exothermic.</td></tr>
</table>

➤ Try yourself 10.2.2

CALCULATING THE ENTHALPY CHANGE OF A REACTION USING AVERAGE BOND ENTHALPIES

Use data from Table 10.2.1 to determine if the reaction below is exothermic or endothermic and estimate the enthalpy change:

$$H_2(g) + F_2(g) \rightarrow 2HF(g)$$

It is important to note that average bond enthalpies can be used in calculations, but the results they give are only an approximation. They are average values, and may differ from the specific molecules involved in your calculation. Also, average bond enthalpies are given for reactions of molecules in a gaseous state, so the enthalpy changes caused by breaking **intermolecular forces** are not accounted for. The average bond enthalpies in Table 10.2.1 are given for reactions of molecules in a gaseous state. If the molecules are not in the gaseous state, there will be differences in the average bond enthalpies.

WS 1.3.2

WS 1.3.4

SCIENCE AS A HUMAN ENDEAVOUR

Cellular respiration

Living organisms are comprised of tiny membrane bound compartments called cells. Most of the reactions that allow the cell to grow and function occur inside the cell. Cells contain, and are made up of, a diverse assortment of chemicals. These range from simple ions to highly complicated molecules, all in an aqueous solution known as the cytoplasm.

All of the chemical process that occur within the cells of an organism constitute the metabolism. Metabolism involves many types of chemical reactions that are essential to maintain life, but broadly speaking they can be classified into condensation, hydrolysis and oxidation–reduction reactions.

Complex molecules are formed through condensation reactions between simpler molecules. For example, starch is formed when simple glucose molecules react, with water molecules being formed (Figure 10.2.7). This is how cells assemble large and complicated molecules.

Hydrolysis reactions occur when water is added in the presence of a catalyst, such as enzymes, to allow the breakdown of complex molecules into simpler ones. This is how cells break down complex carbohydrates into sugars, or fats into fatty acids.

When foods such as carbohydrates are broken down, O_2 reacts with the atoms of glucose to form oxides of carbon and hydrogen. The energy stored in the C–H and C–C bonds is released for the cell to use. This process is known as cellular respiration.

In cells which have a nucleus, respiration occurs on the inner membranes of cell organelles called mitochondria (Figure 10.2.8). The process is complex and involves dozens of intermediate reactions. This is because some of the bonds, in particular the C–C bond, does not oxidise easily. Also, energy must be released in a controlled way that does

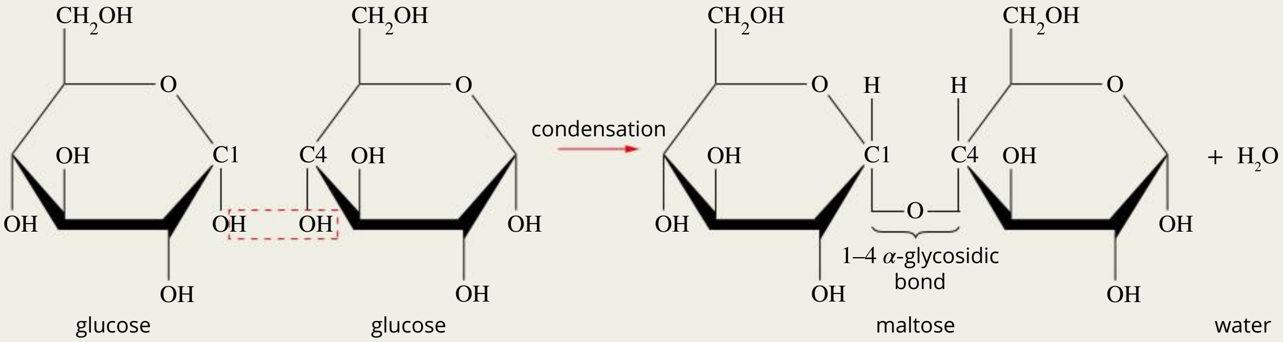


FIGURE 10.2.7 Condensation reaction between glucose molecules. Starch is formed when simple glucose molecules react.

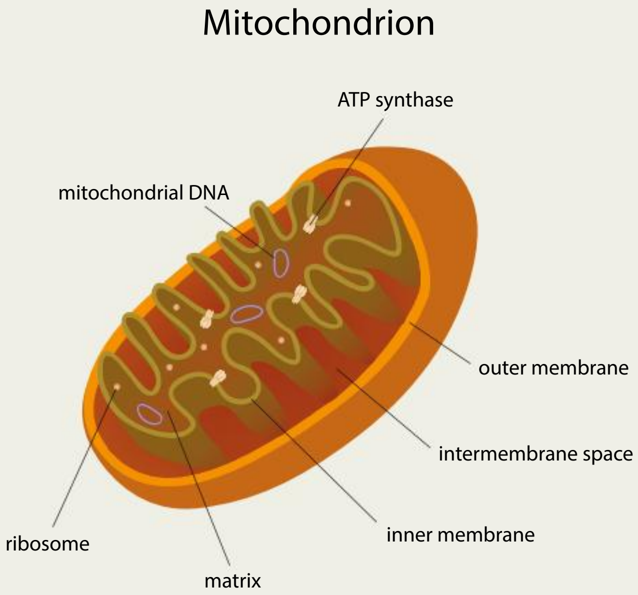


FIGURE 10.2.8 Cellular respiration occurs on the inner membranes of cell organelles known as mitochondria.

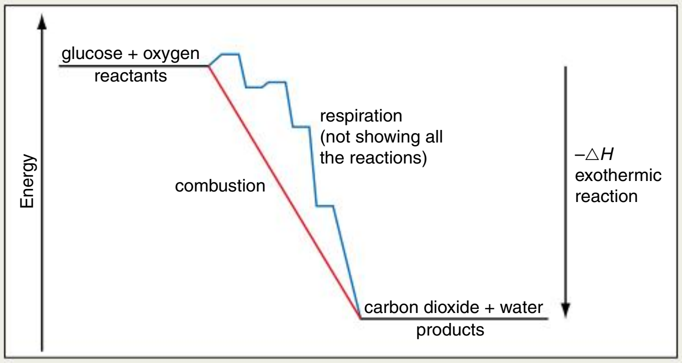


FIGURE 10.2.9 Enthalpy level diagram for the cellular respiration of glucose

not damage the cell. Food molecules are very energetic and if they released their energy rapidly, as in combustion, the cell would likely be destroyed. Figure 10.2.9 is an enthalpy level diagram for the cellular respiration of sucrose, a type of simple carbohydrate molecule.

There are two main types of cellular respiration.

- Aerobic respiration requires oxygen and is the main source of energy for the human body.
- Anaerobic respiration does not require oxygen and yields less energy.

In aerobic respiration, glucose is oxidised to carbon dioxide and water through a sequence of reactions. The overall equation for aerobic respiration is:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$
$$\Delta H = -2803\ \text{kJ mol}^{-1}$$

In humans, an alternative form of respiration, called anaerobic respiration, can occur in muscles during prolonged and vigorous exercise, when the supply of oxygen is limited. It can result in the build-up of lactate ions, $CH_3CH(OH)COO^-$, which can cause painful muscular cramping (Figure 10.2.10). The overall equation for anaerobic respiration in humans is:

$$C_6H_{12}O_6(aq) \rightarrow 2CH_3CH(OH)COO^-(aq) + 2H^+(aq)$$
$$\Delta H = -120\ \text{kJ mol}^{-1}$$

In yeast, anaerobic respiration produces ethanol and carbon dioxide—this process is widely used to produce alcoholic beverages. The equation for this process is:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$
$$\Delta H = -69\ \text{kJ mol}^{-1}$$

These reactions for anaerobic and aerobic respiration are all exothermic. Although anaerobic respiration yields less energy per mole of glucose, it is often a faster process than aerobic respiration. Short bursts of exercise, such as sprinting, rely on anaerobic processes for energy because the individual steps in the reactions involved in anaerobic respiration occur more rapidly.

REVIEW

1. Determine whether cellular respiration is an endothermic or exothermic process.
2. Explain why cellular respiration involves dozens of intermediate steps.
3. Research the respiration of glucose in yeast. Design an experiment that could be used to study the effect of temperature on respiration rate.

FIGURE 10.2.10 The fatigue experienced by an athlete is partly due to formation of lactate ions produced by the anaerobic breakdown of glucose molecules in muscle cells.

10.2 Review

SUMMARY

- The law of conservation of energy states that during any chemical or physical process, energy is neither created nor destroyed.
- In an exothermic reaction, energy will be released from the system into the surroundings because the total chemical energy of the products is less than the total energy of the reactants.
- In an endothermic reaction, energy will be absorbed from the surroundings because the total chemical energy of the products is greater than the total energy of the reactants.
- Enthalpy (*H*) is the chemical energy of a substance.
- Enthalpy change is given the symbol ΔH and is. calculated using the equation $\Delta H = H_p - H_r$.
- A thermochemical equation shows enthalpy change by writing the ΔH value to the right of the chemical equation. The ΔH value usually has the units kJ mol^{-1}.
- Breaking chemical bonds is an endothermic process, and making chemical bonds is an exothermic process.
- Bond enthalpy is the energy needed to break one mole of bonds of gaseous molecules under standard conditions (25°C, 1 atm pressure).

KEY QUESTIONS

Retrieval

1 State the law of conservation of energy.

2 Define enthalpy.

3 Define thermochemical equation.

Comprehension

4 Explain the term 'endothermic' in relation to the total amount of chemical energy of the reactants and products.

5 Explain why both exothermic and endothermic reactions require activation energy to proceed.

6 Explain why average bond enthalpies only give an approximation of the enthalpy change of a reaction.

7 When methane (CH_4) combusts, it reacts with oxygen to form carbon dioxide and water. The heat of combustion of methane is –890 kJ mol^{-1}. Represent this reaction with a thermochemical equation.

8 Draw an enthalpy level diagram for the reaction described in Question **7**.

Analysis

9 The combustion of octane to form carbon dioxide and liquid water can be written as:

$$C_8H_{18}(g) + 12\tfrac{1}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l)$$

$$\Delta H = -5450\text{ kJ mol}^{-1}$$

The combustion of octane to form carbon dioxide and steam can be written as:

$$C_8H_{18}(g) + 12\tfrac{1}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g)$$

Compare the energy released by the combustion of 1 mole of octane to form steam with the energy released by 1 mole of octane to form liquid water.

10 The thermochemical equation for photosynthesis can be represented as:

$$6CO_2(g) + 6H_2O(l) \xrightarrow[\text{chlorophyll}]{\text{light}} C_6H_{12}O_6(aq) + 6O_2(g)$$

$$\Delta H = +2802.5\text{ kJ}$$

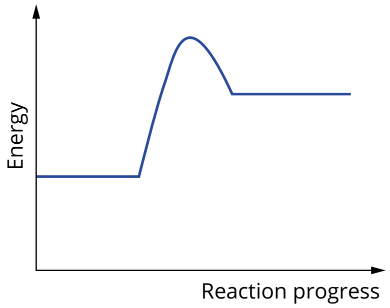
a Determine if this reaction is endothermic or exothermic.

b Consider a new reaction that reversed and halved the above reaction. It is written below. Calculate the energy change of the new reaction.

$$\tfrac{1}{2}C_6H_{12}O_6(aq) + 3O_2(g) \rightarrow 3CO_2(g) + 3H_2O(l)$$

11 The enthalpy level diagram below shows the energy changes during the reaction of hydrogen and iodine to form hydrogen iodide. The equation for the reaction is:

$$H_2(g) + I_2(g) \rightarrow 2HI(g)$$



a Determine from this diagram if this reaction is exothermic or endothermic.

b Describe the relative enthalpies of the reactants and products.

c Compare the size of the activation energy with ΔH.

12 Refer to the equations **A**, **B** or **C** to answer the questions below.

A $Br_2(g) \rightarrow Br_2(l)$

B $Br_2(g) \rightarrow 2Br(g)$

C $\frac{1}{2}Br_2(g) \rightarrow Br(g)$

a Determine which equation describes the bond enthalpy of bromine.

b Determine which equation describes the condensation of bromine.

c Determine which change is exothermic.

13 Identify the atoms involved in each bond and list the number of each of these bonds in the following molecule.

Name	**Formula**
hydrogen	H_2
water	H_2O
methane	CH_4
oxygen	O_2
ammonia	NH_3
carbon dioxide	CO_2

14 Consider the average bond enthalpies in Table 10.2.1 on page 276 to predict the enthalpy changes in the following reactions.

a $2HF(g) \rightarrow H_2(g) + F_2(g)$

b $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

c $C(g) + 2H_2(g) \rightarrow CH_4(g)$

10.3 Measuring energy changes

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain Hess's law
- use Hess's law to calculate change in enthalpy.
- understand the concept of specific heat capacity
- describe how calorimeters work
- use the equation $Q = mc\Delta T$ to calculate the enthalpy change measured using a calorimeter.

Have you ever wondered how the energy values on food nutrition labels, like those in Figure 10.3.1, are calculated? The chemical potential energy of food, like that of all chemicals, is the energy stored in the bonds between the atoms. Chemical reactions can result in this energy being released and measured. In the case of food, the energy content can be measured by burning the food in a bomb calorimeter.

It is the quantitative understanding of how energy is absorbed and released on a molecular scale, which is essential to understanding the heat flow that can be measured.

In this chapter, you will see how the energy changes of chemical reactions can be calculated and measured, both directly and indirectly.

FIGURE 10.3.1 (a) The energy values of a food are shown on its food label. (b) These values are measured by burning the food in a bomb calorimeter.

HESS'S LAW

According to Hess's law, if you add two or more thermochemical equations to give a final equation, then you can also add the enthalpy of reaction of these individual steps to give the final enthalpy of reaction.

Sometimes it is difficult to measure the enthalpy change of a reaction. The reaction may occur too quickly or too slowly to measure, or it may be a brief intermediate step in a series of reactions. Fortunately, in 1840, the Swiss–Russian chemist Germain Henri Hess (1802–50) discovered that it is possible to determine the enthalpy of reaction indirectly. His discovery is known as **Hess's law**. Hess's law states that the amount of heat energy released or absorbed in a chemical reaction is constant, irrespective of the number of steps or the kind of steps by which the reaction is carried out, provided that the same reactants and products are involved.

The energy cycle in Figure 10.3.2 shows the consequence of Hess's law in that no matter the pathway of a chemical reaction, the difference in energy between two substances is independent of the route taken for the conversion. That is, $\Delta H_1 = \Delta H_2 + \Delta H_3$.

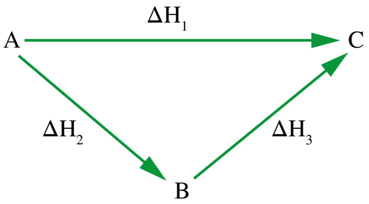


FIGURE 10.3.2 The difference in energy (Δ*H*) between two substances in a chemical reaction is independent of how they are converted. The diagram shows that the difference in energy between A and C (ΔH_1) is the same as the sum of the differences between A and B (ΔH_2), and B and C (ΔH_3).

Figure 10.3.3 shows a stepped energy diagram for the formation of carbon dioxide directly from carbon mixed with oxygen or the formation of carbon dioxide via carbon monoxide. It shows that whichever reaction pathway is taken to form carbon dioxide the amount of energy involved is the same.

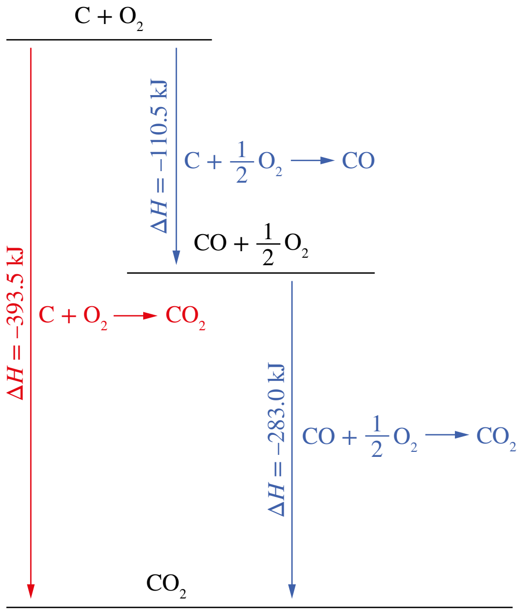


FIGURE 10.3.3 A stepped energy diagram for the formation of carbon dioxide from carbon and oxygen via carbon monoxide

$$\begin{array}{ll} C(s) + \tfrac{1}{2}O_2(g) \rightarrow CO(g) & \Delta H = -110.5\,kJ\,mol^{-1} \\ CO(g) + \tfrac{1}{2}O_2(g) \rightarrow CO_2(g) & \Delta H = -283.0\,kJ\,mol^{-1} \\ \hline C(s) + O_2(g) \rightarrow CO_2(g) & \Delta H = -110.5 + (-283.0) = -393.5\,kJ\,mol^{-1} \end{array}$$

This example demonstrates Hess's law as a restatement of the law of conservation of energy. The importance of Hess's law is that it allows the calculation of enthalpies that cannot be measured directly in the laboratory. Such calculations are based on three key properties of chemical reactions.

1 Chemical reactions are reversible if the appropriate conditions are available.

2 Energy is always involved in a chemical reaction. The products of an exothermic reaction contain less energy than the starting reactants: Δ*H* for an exothermic reaction is negative. Conversely, Δ*H* for the reverse endothermic reaction is positive.

For example, the combustion of carbon to form carbon dioxide can be reversed:

$$\begin{array}{ll} C(s) + O_2(g) \rightarrow CO_2(g) & \Delta H = -393.5\,kJ\,mol^{-1} \\ CO_2(g) \rightarrow C(s) + O_2(g) & \Delta H = +393.5\,kJ\,mol^{-1} \end{array}$$

3 Equations for chemical reactions can be added and subtracted in the manner of ordinary algebraic equations such that complex chemical equations can be reduced to a series of simpler equations of formation with known enthalpies of reaction.

Worked example 10.3.1

USING HESS'S LAW TO CALCULATE THE ENTHALPY CHANGE OF A REACTION

Calculate the enthalpy change for the reaction:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Use the following information.

$$S(s) + \tfrac{3}{2}O_2(g) \rightarrow SO_3(g) \qquad \Delta H = -395\,\text{kJ mol}^{-1} \qquad (1)$$

$$SO_2(g) + \tfrac{1}{2}O_2(g) \rightarrow SO_3(g) \qquad \Delta H = -98\,\text{kJ mol}^{-1} \qquad (2)$$

Thinking	**Working**
In this problem, you must find a path from S(s) to $SO_2(g)$. As the standard enthalpy change cannot be measured directly, an alternative route will enable you to determine the enthalpy change between S(s) and $SO_3(g)$. Reaction 1 gives you the correct starting point.	$S(s) + \tfrac{3}{2}O_2(g) \rightarrow SO_3(g) \quad \Delta H = -395\,\text{kJ mol}^{-1}$
Reverse reaction 2 to give the correct final product, $SO_2(g)$. Remember to change the enthalpy sign.	$SO_3(g) \rightarrow SO_2(g) + \tfrac{1}{2}O_2(g) \quad \Delta H = +98\,\text{kJ mol}^{-1}$
Combine the equations.	$S(s) + \tfrac{3}{2}O_2(g) + SO_3(g) \rightarrow SO_3(g) + SO_2(g) + \tfrac{1}{2}O_2(g)$ $\Delta H = -395 + 98\,\text{kJ mol}^{-1}$
Simplify.	$S(s) + \frac{\cancel{3}2}{2}\,O_2(g) + \cancel{SO_3(g)} \rightarrow \cancel{SO_3(g)} + SO_2(g) + \tfrac{1}{2}\cancel{O_2(g)}$ $\Delta H = -297\,\text{kJ mol}^{-1}$
State the answer.	$S(s) + O_2(g) \rightarrow SO_2(g) \quad \Delta H = -297\,\text{kJ mol}^{-1}$

➤ Try yourself 10.3.1

USING HESS'S LAW TO CALCULATE THE ENTHALPY CHANGE OF A REACTION

Calculate the enthalpy change for the reaction:

$$C(\text{graphite}) + \tfrac{1}{2}O_2(g) \rightarrow CO(g)$$

Use the following information.

$$C(\text{graphite}) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -394\,\text{kJ mol}^{-1} \qquad (1)$$

$$CO(g) + \tfrac{1}{2}O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -283\,\text{kJ mol}^{-1} \qquad (2)$$

CALORIMETRY

The enthalpy released in a chemical reaction can be measured experimentally using a technique called **calorimetry**. Calorimetry is the experimental method of measuring the heat energy released or absorbed by a chemical reaction or physical process, such as by the combustion of a fuel or the reaction between an acid and a base.

In the simplest form of calorimetry, the heat energy released by a chemical reaction can be transferred to water in a container. The quantity of energy transferred can be calculated from the increase in temperature of the water in the **calorimeter**.

A calorimeter is an instrument designed to measure the energy changes in a reaction. Calorimetry is the experimental method of measuring heat energy released or absorbed by a chemical process.

The heat energy transferred to the volume of water can be calculated by measuring the:

- initial temperature of the water
- highest temperature of the water
- volume of water.

The relationship used to calculate the energy that has been transferred to the water is:

$$Q = m \times c \times \Delta T$$

where Q is the energy that is transferred to the water (in J), m is the mass of the water (in g), c is the specific heat capacity of the water ($4.18\,\mathrm{J\,g^{-1}\,°C^{-1}}$), and ΔT is the change in temperature of the water (in °C or K):

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

The **specific heat capacity** of a substance is a measure of the amount of energy (usually in joules) needed to increase the temperature of a specific quantity of that substance (usually 1 gram) by 1°C. It reflects the types of bonds holding the molecules, ions or atoms together in the substance.

Specific heat capacity is given the symbol c and is usually expressed in joules per gram per degrees Celsius, i.e. $\mathrm{J\,g^{-1}\,°C^{-1}}$. It can also be expressed in joules per grams per kelvin, i.e. $\mathrm{J\,g^{-1}\,K^{-1}}$ (an increase of 1°C is the same as an increase of 1 K).

The specific heat capacities of some common substances are listed in Table 10.3.1. You can see that the value for water is relatively high, which is a consequence of the hydrogen bonding between its molecules. You will also note that the metals, copper and lead, have much lower specific heat because of their high conductivity.

TABLE 10.3.1 The specific heat capacity of some common substances

Substance	Specific heat capacity ($\mathrm{J\,g^{-1}\,°C^{-1}}$)
water	4.18
glycerine	2.43
ethanol	2.46
sand	0.48
copper	0.39
lead	0.16

Solution calorimetry

Energy changes that occur in reactions involving solutions can be measured with a **solution calorimeter**. A solution calorimeter may be as simple as a polystyrene foam coffee cup with a lid, as shown in Figure 10.3.4.

The insulation provided by the polystyrene foam prevents the transfer of heat to or from the surroundings of the calorimeter. The reaction is carried out in the calorimeter with an accurately known volume of water. The initial and final temperatures are measured and recorded, as are the amounts of reactants used.

If the temperature of the water in the calorimeter increases, the reaction occurring in the calorimeter is an exothermic reaction. The reaction has released heat energy, the water in the calorimeter has absorbed that energy and the temperature of the water has increased.

Similarly, if the temperature of the water in the calorimeter decreases, the reaction occurring in the calorimeter has absorbed energy from the water. In this case, the reaction is an endothermic reaction.

A coffee cup calorimeter has some limitations. The polystyrene container absorbs some heat, so the temperature change is lower than it would otherwise be and the calculated value for the heat released or absorbed by the reaction is lower than it should be.

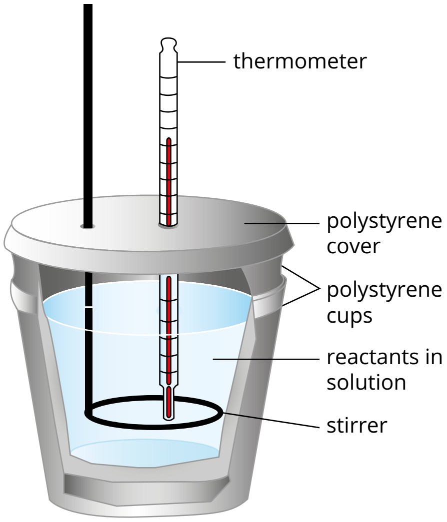


FIGURE 10.3.4 A simple 'coffee-cup' calorimeter

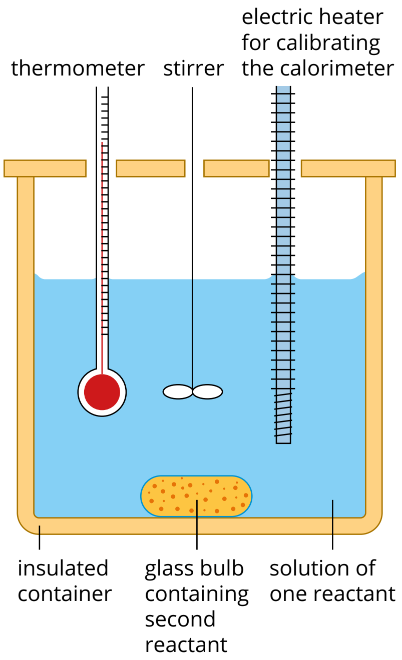


FIGURE 10.3.5 A solution calorimeter; breaking the glass bulb starts the reaction

Solution calorimetry is used in the laboratory to determine the enthalpy changes that occur when acids react with bases and solids dissolve in water. The construction of a laboratory solution calorimeter is shown in Figure 10.3.5. The stirrer is used to ensure the temperature of the water is uniform. The electrical heater is used for calibrating the calorimeter.

Worked example 10.3.2

CALCULATING ENTHALPY CHANGE USING A SOLUTION CALORIMETER

A 25.0 mL sample of water containing 0.025 mol HCl at 25.0°C is added to another 25.0 mL sample of water containing 0.025 mol NaOH at 25.0°C in a 'coffee cup' calorimeter. A reaction occurs. The highest temperature recorded is 32°C. Calculate the enthalpy change.

Assume the density of each solution is the same as water (1.00 g mL^{-1}) and the combined volume of the final solution is equal to the sum of the volumes of the reacting solutions.

Thinking	**Working**
The question is asking you to find the enthalpy change (ΔH). The first step is to list the information provided by the question.	$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O$ $\Delta T = T_f - T_i = 32.0 - 25.0°C = 7.0°C$ m = volume × density mass of the water (m) = $50\,mL \times \frac{1.00\,g}{1\,mL} = 50.0\,g$ c_{water} according to Table 10.3.1 = $4.18\,J\,g^{-1}°C^{-1}$
Use the formula $Q = m \times c \times \Delta T$ to calculate the energy transferred to the water. This is the heat gained by the water, but it is also the heat lost by the reacting HCl and NaOH, therefore change the sign to negative.	$\Delta T = T_f - T_i$ $= 32°C - 25°C$ $= 7°C$ $Q = mc\Delta T$ $= (50.0\,g)(4.18\,J\,g^{-1}\,°C^{-1})(7.0°C)$ $= (50.0\,\cancel{g})(4.18\,J\,\cancel{g^{-1}}\,\cancel{°C^{-1}})(7.0\cancel{°C})$ $= (50.0)(4.18\,J)(7.0)$ $= 1463\,J$ $= 1.46\,kJ$ $Q = -1.46 \times 10^3\,kJ$
Now you need to find the heat released *per mole*. To calculate the energy per mole of acid or base, divide the number of joules by the number of moles.	$n(HCl) = n(NaOH) = 0.025\,mol$ $\Delta H = \frac{Q}{n}$ $= \frac{-1463}{0.025}$ $= -58\,520\,J$ $= -58.5\,kJ$ (to 3 significant figures) The enthalpy change is $-58.5\,kJ\,mol^{-1}$.

WS 1.3.5

➤ Try yourself 10.3.2

CALCULATING ENTHALPY CHANGE USING A SOLUTION CALORIMETER

When 50.0 mL of water containing 0.5 mol HCl at 22.5°C is mixed with 50.0 mL of water containing 0.5 mol NaOH at 22.5°C in a calorimeter, the temperature of the solution increases to 26°C. Calculate the enthalpy change.

Assume the density of each solution is the same as water (1.00 g mL^{-1}) and the combined volume of the final solution is equal to the sum of the volumes of the reacting solutions.

HEAT OF COMBUSTION

Combustion reactions are rapid reactions between a fuel and oxygen. They are exothermic reactions that often release a large amount of energy. The **heat of combustion** is defined as the enthalpy change that occurs when a specified amount (e.g. 1 g, 1 L, 1 mol) of the fuel burns completely in oxygen. It is usually measured at conditions of 298 K (25°C) and 100 kPa (1 bar), which means that the water produced should be shown in the liquid state. The heat of combustion can be given the symbol ΔH_c.

Many fuels, including wood, coal and kerosene, are mixtures of chemicals and do not have a specific chemical formula or molar mass. This means their heat of combustion cannot be expressed in kJ mol^{-1}. Therefore, it is measured only as kJ g^{-1}, kJ L^{-1} or MJ/tonne.

The heats of combustion for some common elements and compounds present in fuels are listed in Table10.3.2. Heat energy is released during combustion, so ΔH_c always has a negative value.

Only fuels that exist as pure substances can have their heat of combustion measured in kJ mol^{-1}.

WS 1.3.3

TABLE 10.3.2 Heats of combustion for some common elements and compounds

Substance	Heat of combustion, ΔH_c (kJ mol^{-1})
methane	–890
ethane	–1560
propane	–2220
butane	–2886
octane	–5450
methanol	–725
ethanol	–1367
hydrogen	–286
carbon (graphite)	–394

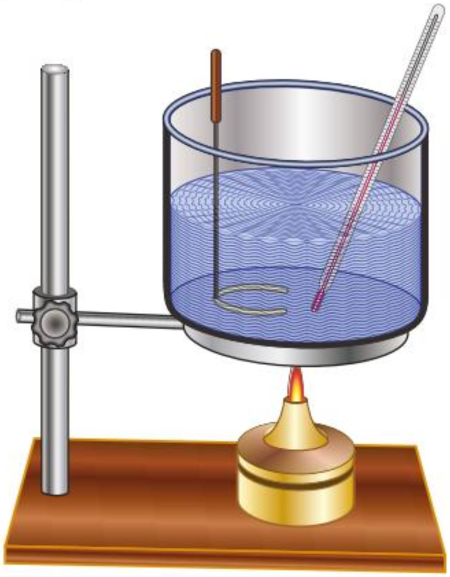


FIGURE 10.3.6 A copper calorimeter is used to measure the energy released by combustion reactions.

The heat of combustion of a substance can be measured using a copper calorimeter as shown in Figure 10.3.6. In a copper calorimeter, the heat released by the burning fuel is absorbed by the water above. The temperature of this water will increase, which allows the heat of combustion to be calculated using the formula:

$$\Delta H_c = \frac{Q}{n} = -\frac{mc\Delta T}{n}$$

A minus sign is placed before the *m* according to the convention that combustion reactions are always exothermic and will have negative enthalpy values.

Worked example 10.3.3

CALCULATING THE HEAT ENERGY TRANSFERRED TO WATER FROM FUEL

A sample of fuel is used to heat 100 mL of water. Calculate the heat energy, in kilojoules, that has been transferred if the temperature of the water increases from 18.5°C to 44.0°C.

Thinking	**Working**
Change the volume of water in millilitres to mass of water in grams. Remember that the density of water is 1.00 g mL^{-1}.	1 mL of water has a mass of 1 g, so 100 mL of water has a mass of 100 g.
Calculate the change in temperature, ΔT, by subtracting the initial temperature from the final temperature: $\Delta T = T_f - T_i$	$\Delta T = T_f - T_i$ $= 44.0 - 18.5$ $= 25.5$°C
Calculate the heat energy transferred to the water, in joules, using the formula: $Q = m \times c \times \Delta T$	$Q = m \times c \times \Delta T$ $= 100 \times 4.18 \times 25.5$ $= 10\,659$ J
Express the quantity of energy, in kJ, to the appropriate number of significant figures. Remember that to convert from joules to kilojoules, you divide by 10^3.	$Q = \frac{10\,659}{1000}$ $= 10.7$ kJ (3 significant figures)

➤ Try yourself 10.3.3

CALCULATING THE HEAT ENERGY TRANSFERRED TO WATER FROM FUEL

A sample of fuel is used to heat 200 mL of water. Calculate the heat energy, in kilojoules, that has been transferred if the temperature of the water increases from 15.5°C to 30.0°C.

Worked example 10.3.4

CALCULATING THE HEAT ENERGY PER GRAM TRANSFERRED TO WATER FROM A FUEL

A 0.850 g sample of fuel was burnt under a steel can containing 150 mL of water. After the flame went out, the mass of the fuel was 0.300 g and the temperature of the water had risen by 22.5°C. Calculate the energy content of the fuel in kJ g^{-1}.

Thinking	**Working**
Calculate the heat energy absorbed by the water in joules, using the formula: $Q = m \times c \times \Delta T$	$Q = m \times c \times \Delta T$ $= 150 \times 4.18 \times 22.5$ $= 14\,108$ J
Express the quantity of energy in kJ. Remember that to convert from joules to kilojoules, you divide by 10^3.	Q (in kJ) $= \frac{14108}{1000}$ $= 14.108$ kJ

Calculate the mass of the fuel that was burnt by subtracting the final mass from the initial mass: $\Delta m = m_i - m_f$	$\Delta m = m_i - m_f$ $= 0.850 - 0.300$ $= 0.550\,g$
Calculate the energy content of the fuel by dividing the energy transferred to the water by the change in mass during combustion: Energy content $= \frac{Q}{\Delta m}$	Energy content $= \frac{Q}{\Delta m}$ $= \frac{14.108}{0.550}$ $= 25.7\ kJ\,g^{-1}$ (3 significant figures)

➤ Try yourself 10.3.4

CALCULATING THE HEAT ENERGY PER GRAM TRANSFERRED TO WATER FROM A FUEL

A 2.500 g corn chip was burnt under a steel can containing 200 mL of water. After the flame went out, the mass of the corn chip was 1.160 g and the temperature of the water had risen by 35.0°C. Calculate the energy content of the corn chip in kJ g^{-1}.

Heat loss

When energy is transferred from a fuel, across an open space, heat is lost to the surroundings such as the air. Similarly, if there is no lid on a container of water, heat will be lost from the surface of the water. This heat loss is illustrated in Figure 10.3.7.

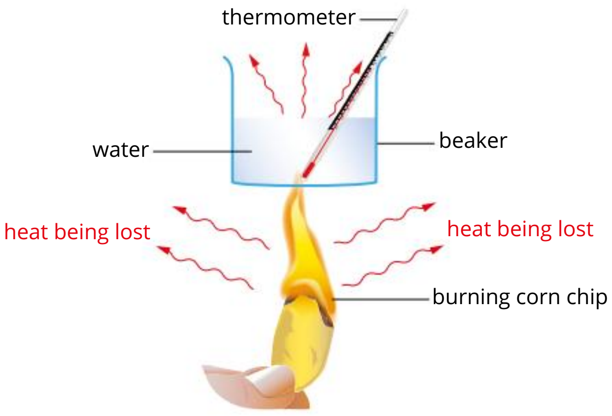


FIGURE 10.3.7 Heat is lost to the surroundings when a piece of food is burnt to heat some water.

When some heat energy from the burning food is transferred to the surrounding air, the temperature of the water does not increase as much as it would if all the energy was used to heat the water. A lower change in temperature, ΔT, of the water results in a lower energy value, Q.

There are several ways to reduce heat loss during the calorimetry experiment shown in Figure 10.3.7, including:

- putting a lid on the calorimeter
- insulating the beaker of water (with flameproof material)
- placing insulation around the burning food, although sufficient oxygen must reach the food for complete combustion to occur.

PA
1.3.2

MANDATORY PRACTICAL 2

Using a calorimeter: measuring the energy change during a chemical reaction

Research and planning

Aim

- To calibrate a calorimeter by measuring the increase in temperature that results from a measured input of electrical energy
- To determine the heat of solution when potassium nitrate dissolves in water

Rationale (scientific background to the experiment)

Calorimetry is a technique for measuring the energy change during a chemical reaction.

A calorimeter is often calibrated by passing an electric current through a heating coil in the instrument. The temperature of the calorimeter and its contents increases in proportion to the amount of energy supplied by the heating coil. The calibration factor (joules of energy required per degree Celsius of temperature increase) is determined by measuring the temperature increase that results from a measured energy input. The value of the calibration factor depends upon the characteristics of the particular calorimeter.

Once calibrated, the calorimeter can be used to measure enthalpy changes in chemical reactions. While bomb calorimeters are used to determine the energy content of food, solution calorimeters are used to determine energy changes for processes that occur in solution. Figure 10.3.8 summarises the steps involved in calorimetry calculations.

Potassium nitrate is often used as a preservative in food and to help keep the colour of cured meats. In this experiment, you will determine the heat of solution when potassium nitrate dissolves in water.

Timing

50 minutes—can be split into two individual practicals, as long as students use the same calorimeter

Materials

- 100 mL measuring cylinder
- calorimeter (with a code number for identification)
- dc power supply
- 5 × wires leads
- thermometer, –10 to 50°C
- stopwatch
- ammeter
- voltmeter
- safety glasses
- 3 g potassium nitrate, coarsely ground
- spatula
- weighing bottle
- electronic balance

Method

Risk assessment

Consideration of risks includes chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and Assessment book or download it from your eBook.

PART A

1. Record the code number of your calorimeter in the space below Tables 10.3.3 and 10.3.4. You will need to use the same calorimeter after you have calibrated it.
2. Pour 100 mL of water into the calorimeter. Stir the water and record its temperature every 30 seconds in Table 10.3.3 for several minutes. (You could enter data directly into a computer spreadsheet program for later processing.)

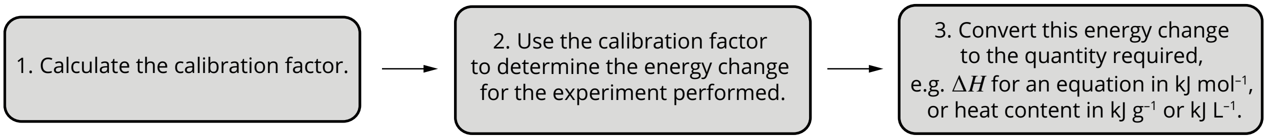


FIGURE 10.3.8 Steps involved in calorimetry calculations

PRE-LAB SAFETY INFORMATION		
Material used	**Hazard**	**Control**
use of electrical equipment	electric shock and damage to equipment	check all connections carefully
Please indicate that you have understood the information in the safety table. Name (print): ______________________________ I understand the safety information (signature): ______________________________		

3. Once the temperature is steady, apply a measured voltage of approximately 6 V for exactly three minutes using the circuit shown in Figure 10.3.9. Stir continuously and record the temperature every 30 seconds.

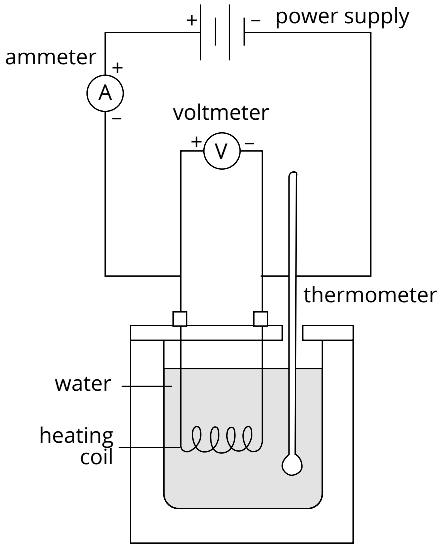


FIGURE 10.3.9 Experimental circuit set-up

4. Record the potential difference (voltage) and current while the water is heating. After three minutes, turn off the power supply. Continue to stir the water in the calorimeter and record its temperature every 30 seconds for a further three minutes.
5. Discard the water in the calorimeter and repeat steps 1–4 using a fresh supply of water. Record your results in Table 10.3.4 on page 292.

PART B

1. Pour 100 mL of water into a calorimeter. Stir and record the temperature of the water.
2. Add about 3 g of accurately weighed potassium nitrate to the calorimeter. Stir and record in your results in Table 10.3.5 on page 292.

Variables

- **i** Independent: the amount of potassium nitrate
- **ii** Dependent: the heat of solution when potassium nitrate dissolves in water
- **iii** Controlled: the calorimeter used, the balance used, the measuring cylinder used, the DC power supply, the ammeter and voltmeter, the thermometer used.

Reflect and check that your evaluation demonstrates these characteristics

☐ Effective and efficient investigation of phenomena is directed by a concise and relevant research question.
☐ Accurate application of scientific concepts, theories, models and systems is demonstrated by a considered rationale for the experiment including scientific justification for the chosen modifications.
☐ Effective and efficient investigation of phenomena is demonstrated by a considered methodology that enables the collection of sufficient, relevant data that addresses the research question.
☐ Safe investigation of phenomena is ensured through the consideration and management of risks, and ethical or environmental issues.

Analysing

Raw data

TABLE 10.3.3 Results for calibration 1

Time (min)	Temperature (°C)	Time (min)	Temperature (°C)	Time (min)	Temperature (°C)

Code number of your calorimeter: ______________________

TABLE 10.3.4 Results for calibration 2

Time (min)	Temperature (°C)	Time (min)	Temperature (°C)	Time (min)	Temperature (°C)

Code number of your calorimeter: ________________

Potential difference during heating for calibration 1: ________________ V

Current during heating for calibration 1: ________________ A

Potential difference during heating for calibration 2: ________________ V

Current during heating for calibration 2: ________________ A

TABLE 10.3.5 Dissolution of potassium nitrate results

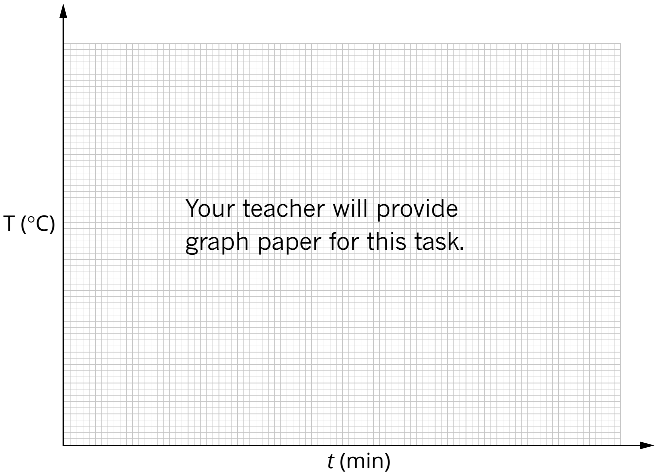
Mass (g) of KNO_3	
Initial temperature (°C)	
Final temperature (°C)	

Processed data

Reflect and check that your data analysis demonstrates these characteristics

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.
☐ Accurate application of algorithms, visual and graphical representations of data are demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

1 Graph the two sets of data you collected when calibrating the calorimeter, plotting temperature against time.



Your teacher will provide graph paper for this task.

FIGURE 10.3.10 Calorimeter calibration graphs

2 For each set of data, record the calculations in Table 10.3.6.

TABLE 10.3.6 Data calculations

	First calibration	Second calibration
Temperature increase (°C) over the three minute heating period		
Thermal energy supplied by the heating coil in the calorimeter E = voltage (V) × current (amps) × time (s)		
Calibration factor for each set of data Calibration factor = $\frac{E}{\Delta T}$, where ΔT = temperature change (°C)		

Mean value of the two calibration factors: ____________ J °C^{-1}

Analysis

Reflect and check that your analysis demonstrates these characteristics

☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.
☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.
☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Discussion

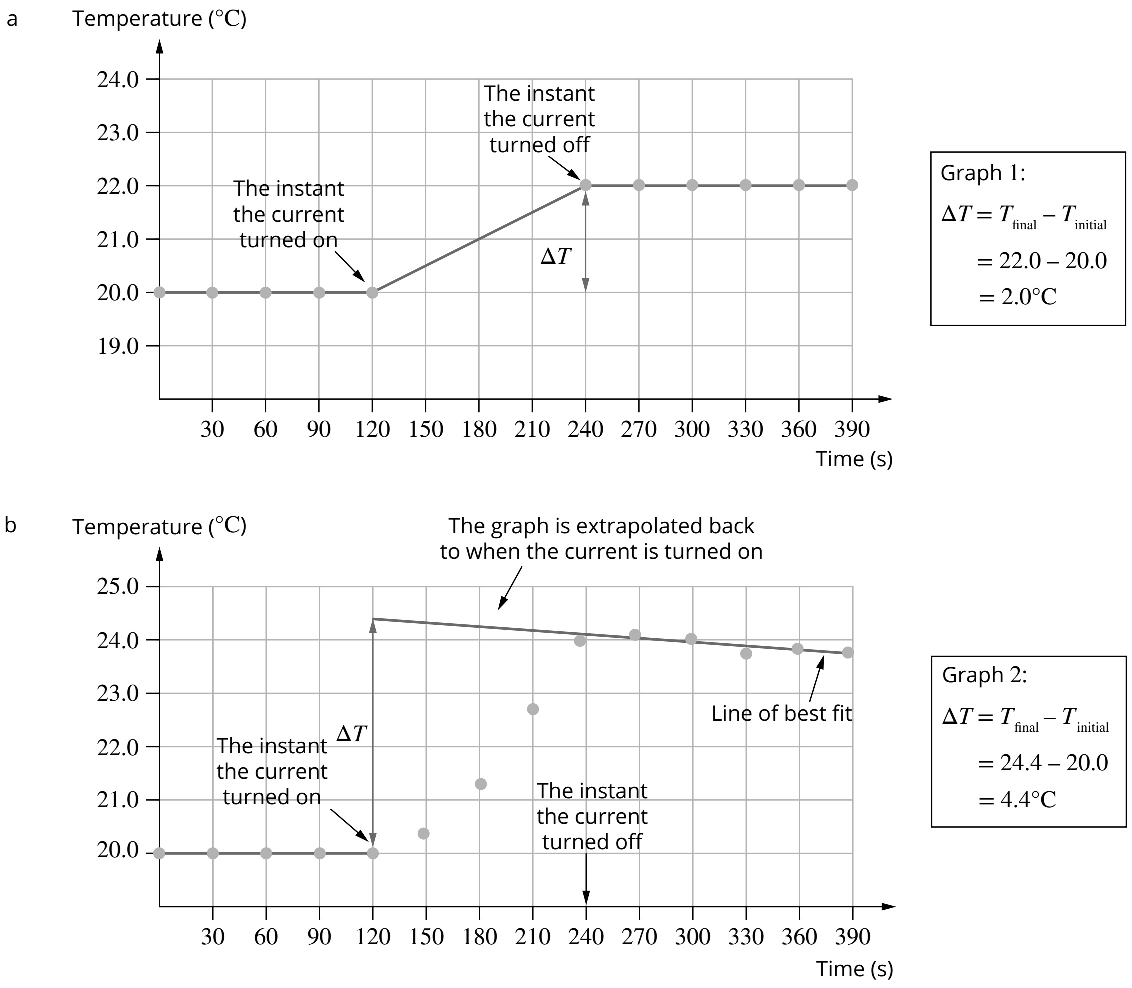
1 What energy transformation occurs when the calorimeter and its contents are being heated?

2 The graphs in Figure 10.3.11 show temperature against time for two calorimeters: (a) a graph for a well-insulated calorimeter and (b) a graph for a poorly insulated calorimeter with heat loss.
Compare these graphs to your graphs above. Comment on the insulation of your own calorimeter.

3 What would be the effect on the calibration factor obtained from this experiment of using 50 mL of water instead of 100 mL of water to calibrate the calorimeter?

4 Write an equation for the dissolution of solid potassium nitrate in water.

5 Using your measurements of the changes in temperature and the calibration factor of the calorimeter, calculate the energy change, in joules, that occurred. (If the calibration factor is unknown, an estimate of the energy changed can still be calculated because 1 mL of water requires 4.18 J to raise the temperature by 1°C.)



Graph 1:

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$
$$= 22.0 - 20.0$$
$$= 2.0°\text{C}$$

Graph 2:

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$
$$= 24.4 - 20.0$$
$$= 4.4°\text{C}$$

FIGURE 10.3.11 Temperature–time graphs for the calibration of two calorimeters

6 Calculate the heat of reaction (ΔH) and give the sign for the ΔH.

7 For this reaction, which was the greater: the energy required to break the bonds in the reactants or the energy released when new bonds were formed to make the products? To explain your answer, draw an energy profile diagram and label it clearly.

8 Energy is neither created nor destroyed in a chemical reaction. Explain where the energy released (or absorbed) by the reaction comes from (or goes to).

Interpreting and communicating

Conclusion

1 The experiment had the following aims.

a To calibrate a calorimeter.
Based on your results were you able to determine a calibration factor for the calorimeter? What do your results suggest about the adequacy of your calorimeter?

b To determine the heat of solution when potassium nitrate dissolves in water.
Based on your results, write a thermochemical equation for the reaction and state whether this was an exothermic or endothermic reaction.

Evaluation

2 Considering your analysis and conclusion, did the experiment provide an effective and efficient method of determining the heat of solution when potassium nitrate was dissolved in water?

3 How accurate do you think the results are based on the equipment used; that is, what percentage uncertainty do you expect? (Round to the nearest 10 per cent.)

Improvements

4 If you were to repeat the experiment, identify the steps that you would do differently. When considering this question you need to include in your answer issues with:

- the methodology and how you would change it to improve the results
- how you performed your tasks and the skills and whether you need to improve your technique
- the collection of data and how this could be improved or uncertainty reduced.

Extension

5 Suggest some practical improvements that could be made to the calorimeter you have calibrated and used to make your results more accurate.

6 Identify any limitations and possible problems that might restrict the use of this methodology in determining the heat of solution when a substance is dissolved in water.

Reflect and check that your evaluation demonstrates these characteristics

☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis).

☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respects to the error analysis and limitations or sufficiency of the data.

☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment that are logically derived from the analysis of the evidence.

10.3 Review

SUMMARY

- According to Hess's law, if you add two or more thermochemical equations to give a final equation, then you can also add the enthalpy of reaction for each equation to give the overall enthalpy of reaction.
- Calculations involving Hess's law rely on three key properties of chemical reactions.
  1. Chemical reactions are reversible if the appropriate conditions are available.
  2. ΔH for an exothermic reaction is negative. ΔH for the reverse endothermic reaction is positive.
  3. Equations for chemical reactions can be added and subtracted in the manner of ordinary algebraic equations such that complex chemical equations can be reduced to a series of simpler equations of formation with known enthalpies of reaction.
- A calorimeter is an instrument designed to measure the energy changes in a reaction. Calorimetry is the experimental method of measuring heat energy released or absorbed by a chemical process.
- The enthalpy of reaction can be determined using a calorimeter: $\Delta H_c = \frac{Q}{n} = -\frac{mC\Delta T}{n}$.
- The heat of combustion is defined as the enthalpy change that occurs when a specified amount of the fuel burns completely in oxygen.

KEY QUESTIONS

Retrieval

1 Indicate if the following statement is true or false. 'According to Hess's law, if you add two or more thermochemical equations to give an equation, you cannot always add the enthalpy of each reaction to give the overall enthalpy of reaction.'

2 State what a calorimeter is.

3 Describe a simple calorimeter that could be used to measure the enthalpy change of reactions between two solutions.

4 Define the term 'heat of combustion'.

Comprehension

5 Hess's law calculations rely on three fundamental properties of chemical reactions. Explain what they are.

6 Explain why different substances have different specific heat capacities.

Analysis

7 Determine the enthalpy change ΔH (in kJ mol^{-1}), for the following reaction.

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

Use the information below to help you.

$N_2(g) + O_2(g) \rightarrow 2NO(g) \quad \Delta H = +180.5\ \text{kJ mol}^{-1} \quad (1)$

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g) \quad \Delta H = +66.4\ \text{kJ mol}^{-1} \quad (2)$

8 Determine the enthalpy change, ΔH (in kJ mol^{-1}), for the following reaction.

$MnO_2(s) + C(s) \rightarrow Mn(s) + CO_2(g)$

Use the equations below to help you.

$C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -390\ \text{kJ mol}^{-1} \quad (1)$

$Mn(s) + O_2(g) \rightarrow MnO_2(s) \quad \Delta H = -520\ \text{kJ mol}^{-1} \quad (2)$

9 Determine the energy required to increase the temperature of the following.

a 100 mL of water by 15.2°C

b 500 mL of water from 16.0°C to 28.0°C

c 1.50 kg of water from 20.0°C to 30.0°C

10 Determine the final temperature of 250 g of water that is initially at 18.0°C if the amount of energy transferred to the water is 4.50 kJ.

11 A 20 mL sample of NaOH solution is added to 100 mL of H_2SO_4 solution. The temperature is measured to rise from 24.0°C to 49.5°C. Determine the enthalpy change.

12 Determine the energy content, in kJ g^{-1}, of a wafer biscuit if combustion of 5.00 g of biscuit heats 500 mL of water by 38.3°C.

Chapter review

10

KEY TERMS

activation energy
average bond enthalpy
boiling point
bond enthalpy
calorimeter
calorimetry
Celsius scale
change of state
chemical energy
combustion
condensation
deposition
endothermic
energy
enthalpy
enthalpy change
enthalpy level diagram
evaporation
exothermic
freezing
gaseous
heat
heat content
heat of combustion
heat of reaction
Hess's law
intermolecular force
joule
Kelvin scale
kinetic theory
law of conservation of energy
liquid
melting
phase of matter
product
reactant
solid
solution calorimeter
specific heat capacity
state of matter
state symbols
sublimation
surroundings
system
temperature
thermochemical equation
thermochemistry
vaporisation
vapour

KEY QUESTIONS

Retrieval

1 Describe what is meant by the term 'system'.

2 Define the concept of temperature.

3 Define the concept of heat.

4 Identify whether each of the following statements related to activation energy is true or false.

- **a** Activation energy is the energy required to break bonds in the reactants.
- **b** Reactions that start immediately do not have an activation energy.
- **c** Reactions that release energy overall do not need to absorb activation energy.
- **d** The match used to light a fire is providing activation energy.

5 Identify which one of the following is correct about the energy profile diagrams of both endothermic and exothermic reactions.

- **A** There is always less energy absorbed than released.
- **B** The enthalpy of the products is always less than the energy of the reactants.
- **C** Some energy is always absorbed to break bonds in the reactants.
- **D** The ΔH value is the difference between the enthalpy of the reactants and the highest energy point reached on the energy profile.

6 Define specific heat capacity.

7 The equation $Q = m \times c \times \Delta T$ describes the relationship between heat energy and the temperature change of a substance being heated. It is used to determine how much energy has been absorbed by a material, such as a mass of water, when its temperature increases by a measured amount.

Identify which of the following statements about this equation is correct.

- **A** The specific heat capacity, c, depends on the material that is generating the heat energy.
- **B** The mass, m, represents the mass of the material that is being heated.
- **C** Q is measured in $\text{kJ}\,^{\circ}\text{C}^{-1}$.
- **D** ΔT refers to the change in temperature of the fuel generating the heat.

Comprehension

8 Explain the three states of matter in terms of kinetic theory.

9 Your fingers feel cold when you touch an ice cube. Explain the important thermochemical principle that this change illustrates.

10 Determine whether the following processes are exothermic or endothermic. Explain your answers.

- **a** burning of wood
- **b** melting of ice
- **c** recharging of a car battery
- **d** decomposition of plants in a compost heap

Analysis

11 Assess why steam at 100°C has greater energy content than water at 100°C.

12 Determine the energy levels in the units specified (to 3 significant figures).

- **a** 2205 J to kJ
- **b** 0.152 kJ to J
- **c** 1 890 000 J to MJ
- **d** 0.0125 MJ to kJ

13 Compare evaporation and boiling.

14 The combustion reaction of ethyne gas that occurs in a welding machine can be represented by the thermochemical equation:

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$
$$\Delta H = -2619\,\text{kJ}$$

- **a** Judge if this reaction is endothermic or exothermic.
- **b** Determine the new value of ΔH if the equation was now written as follows:

$$4C_2H_2(g) + 10O_2(g) \rightarrow 8CO_2(g) + 4H_2O(l)$$

15 The combustion of butane gas in portable stoves can be represented by the thermochemical equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$
$$\Delta H = -5772\,\text{kJ}$$

- **a** Compare the overall energy of the bonds in the reactants to the overall energy of the bonds in the products.
- **b** Draw an enthalpy level diagram for the reaction, labelling ΔH and activation energy.

16 Calculate the enthalpy change for the following reaction using the average bond energies from Table 10.2.1 on page 276:

$$C(g) + O_2(g) \rightarrow CO_2(g)$$

17 Consider why reversing a chemical reaction reverses the sign of ΔH.

18 Calculate the enthalpy change (ΔH) in kJ for the following reaction:

$$2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

Use the enthalpy changes for the combustion of aluminium and iron to help you.

$$2Al(s) + \tfrac{3}{2}O_2(g) \rightarrow Al_2O_3(s)$$
$$\Delta H = -1676.0\,\text{kJ} \qquad (1)$$
$$2Fe(s) + \tfrac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$$
$$\Delta H = -822.1\,\text{kJ} \qquad (2)$$

19 Calculate the energy needed to heat the following.

- **a** 100 mL of water from 20.0°C to 80.0°C
- **b** 250 mL of water from 25.0°C to 100.0°C
- **c** 1.5 kg of water from 20.0°C to 30.0°C
- **d** 2300 g of water from 18.0°C to 100.0°C
- **e** 300 g of cooking oil from 18.0°C to 100.0°C ($C_{\text{(cooking oil)}} = 2.2\,\text{J}\,\text{g}^{-1}\,°\text{C}^{-1}$).

20 A 0.254 g piece of black coal was burnt and used to heat 300 mL of water. The temperature of the water rose from 18.25°C to 24.92°C. Calculate the heat of combustion, in $\text{kJ}\,\text{g}^{-1}$, of the coal. Assume all of the heat released during combustion was used to heat the water.

21 A peanut with a mass of 1.200 g was burnt under a steel can containing 200.0 mL of water. After the flame went out the mass of the peanut was 0.750 g and the temperature of the water had risen by 13.20°C. Calculate the energy content of the peanut in $\text{kJ}\,\text{g}^{-1}$, assuming all the energy from the burning peanut is transferred to the water.

22 An aluminium mass is heated and placed into a cup calorimeter containing 40.0 mL of water at 17.0°C. The water reaches a temperature of 20.0°C. Calculate the amount of energy (in J) released.

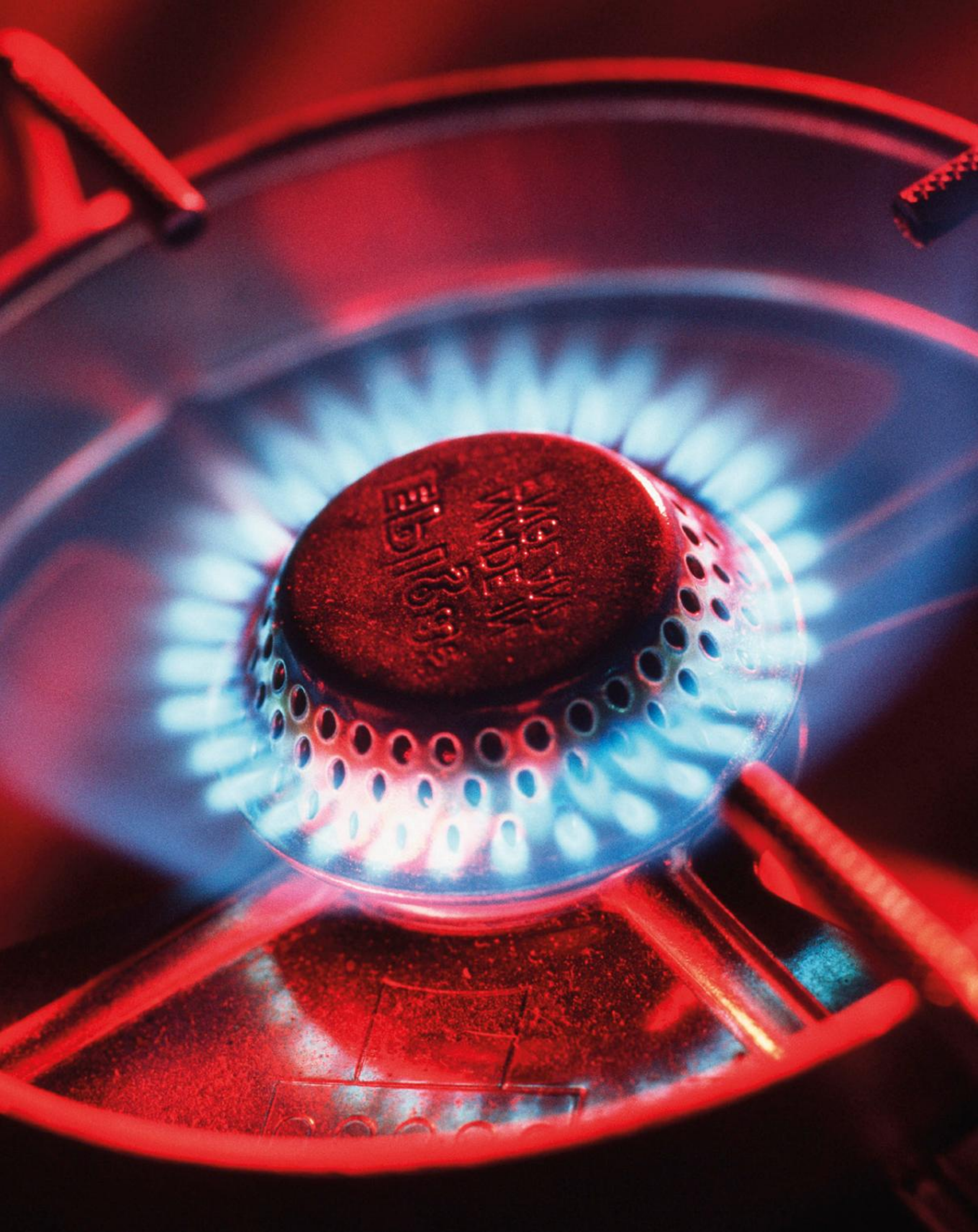
Knowledge utilisation

23 A 500 mL volume of water in a beaker was heated using the energy released by the combustion of a 3.00 g piece of wood. The observed temperature rise was 22.9°C.

- **a** Determine the heat of combustion of wood in $\text{kJ}\,\text{g}^{-1}$ and in $\text{MJ}\,\text{t}^{-1}$.
- **b** Determine if the heat of combustion you calculated in part **a** would be higher or lower than the actual heat of combustion of the wood. Give reasons for your answer.

24 A 2.5 g sample of a hydrocarbon mixture was completely burnt in a calorimeter containing 750 mL of water. The initial temperature of the water was 18°C. The temperature of the water on all the fuel had been used was 46°C.

- **a** Deduce if this was an endothermic or exothermic reaction.
- **b** Determine whether ΔH will be positive or negative for this reaction.
- **c** Predict the amount of energy (in kJ) that was used to heat the water, assuming all the energy from the reaction was transferred.



CHAPTER 11

Fuels

While the ability to feed and clothe ourselves is an ancient hallmark of human endeavour, the harnessing of energy for our use is a more recent and possibly more far-reaching achievement. You use energy everywhere in large quantities to power our machines, light our homes and fuel our technologies. In forming carbohydrates during photosynthesis, plants have made molecular packages of energy—chemical potential energy. The energy stored in these bonds is released as you go about your daily life. The focus of our energy industry has become realigned during the last 50 years—scientists, engineers and politicians are concerned to maintain energy supply and protect the long-term health of our planet by using cleaner, more efficient energy to fuel our communities.

In this chapter, you will learn how fossil fuels and biofuels are used to meet global energy needs. You will gain an appreciation of the chemistry that underpins decisions about the use of different fuels for a particular purpose, and will consider the environmental impact of using different types of fuels, including the emission of carbon and other pollutants into the atmosphere. New technologies related to the production of renewable fuels and the potential for reducing the harmful impact of fossil fuels will also be explored.

Syllabus subject matter

Topic 3 • Chemical reactions—reactants, products and energy change

FUELS

- compare fuels, including fossil fuels and biofuels, in terms of their energy output, and evaluate their suitability for purpose, and the nature of products of combustion.

SCIENCE AS A HUMAN ENDEAVOUR

- **Minimising use of energy in industry:** Green chemistry principles can be applied to industrial processes to reduce energy requirements.
- **Use of fuels in society:** Biofuels are more efficient and have less environmental impact than fossil fuels.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

11.1 Fuels in society

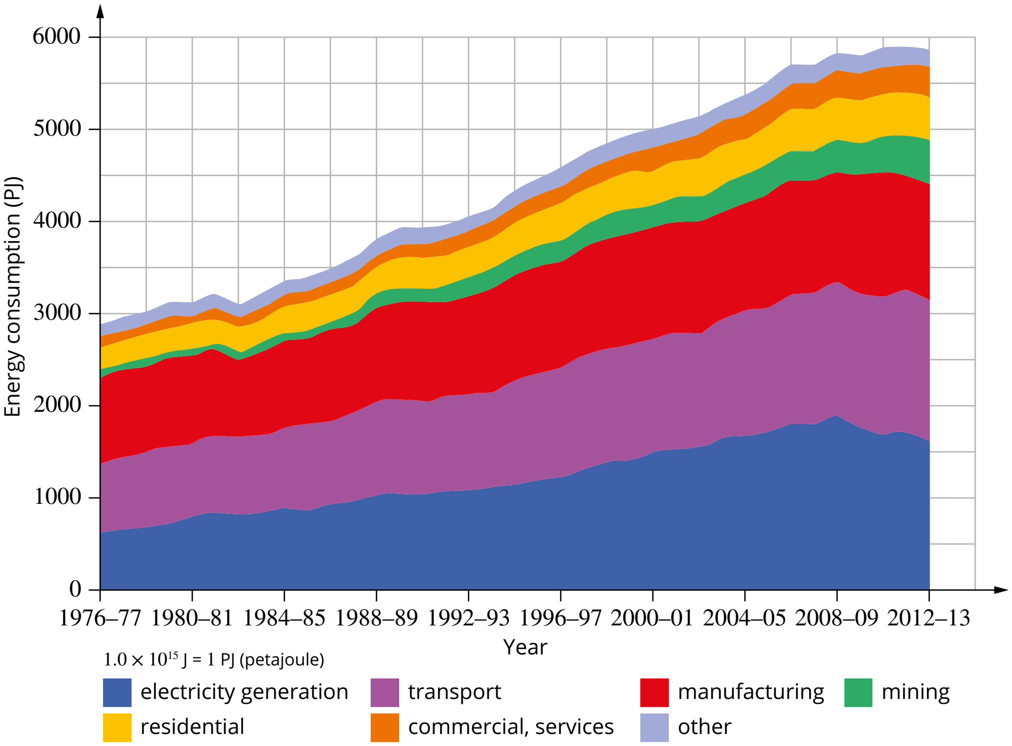
BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that fuels are vital for the needs of our society, and that their use is increasing
- describe the combustion reaction, which is common to all fuels
- understand that most fuels used are fossil fuels, which are not renewable.

In Australia, and around the world, most of the energy used to power a range of transport and manufacturing engines comes from **fossil fuels**. About 86% of Australian electricity is generated from these fuels (73% from coal and 13% from natural gas). The remaining 14% is produced from **renewable energy** sources including hydroelectricity (7%) and wind, biofuels and solar energy (another 7%). Initiatives are in place to decrease the use of fossil fuels and increase the use of renewable energies.

THE NEED FOR FUELS

World energy consumption is around 4×10^{20} joules per year. The United States consumes a quarter of the world's energy. Australia consumes about one-hundredth of the world's energy. But energy consumption per person in Australia is only just below that of the United States. Figure 11.1.1 shows the ways in which Australians use energy. The graph shows that manufacturing, electricity generation and transportation account for more than three-quarters of Australia's total energy consumption, and that energy consumption has doubled in the last 30 years.



Source: Bureau of Resources and Energy Economics

FIGURE 11.1.1 National primary energy consumption by sector

The world first became aware that fossil fuels are a finite energy reserve during the 'oil crisis' of the early 1970s. Several Middle Eastern oil exporters restricted production for political reasons. This dramatically increased the cost of crude oil and caused huge increases in the price of petrol around the world.

Given the limited reserves and concerns about the link between fossil fuels and climate change, there is considerable interest in identifying and developing new energy sources. The development of alternative sources for large-scale energy production is not a simple task. Replacement energy sources need to meet a range of requirements, such as being reliable, **sustainable** and cost-effective. Figure 11.1.2 shows the increase in world energy production from different sources.

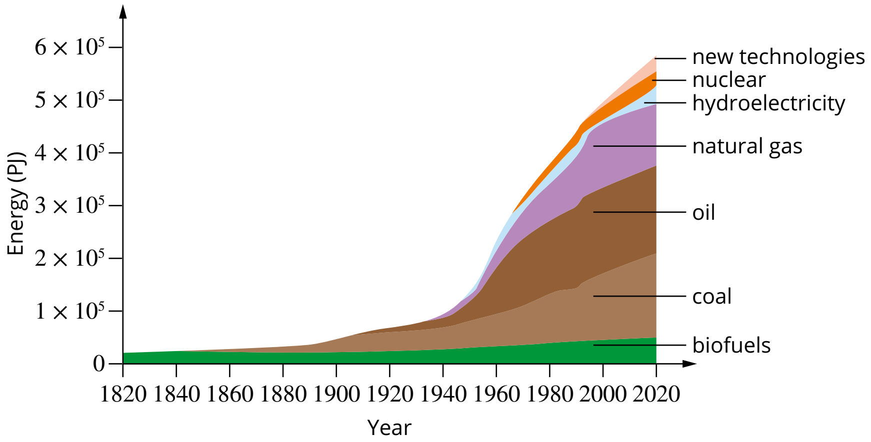


FIGURE 11.1.2 This graph shows the increase in world energy production in the last 200 years.

Oil and coal reserves

Australia's relatively small oil reserves are likely to be exhausted later this century, and the longevity of our much larger coal supplies relies upon continual expansion and discovery of new coalfields. The availability of this resource is uncertain beyond the year 2100 unless further deposits are found.

The past few decades have seen a growth in development of strategies to reduce our reliance on fossil fuels and a corresponding increase in the number of chemical engineers working in the energy industry. Government bodies such as environmental protection agencies have developed and ratified a set of principles, known as **green chemistry principles**, to be pursued in a wide range of contexts.

The overarching goals of green chemistry, explored further in Unit 4, are to be more resource efficient and to design safer molecules, materials, products and processes. As you engage with ideas explained in this chapter, you will see that new ways of providing and using energy use these principles.

COMBUSTION AND ENERGY

The **combustion** of fuels is an exothermic reaction that provides us with energy for daily use, including heat for warmth and cooking, and a source of power for transport. In Chapter 10, you learnt that **thermochemical equations** show the energy released or absorbed during a chemical reaction by including a sign and numerical value for the **enthalpy** change that occurs in the reaction represented by the equation. Thermochemical equations are used to calculate the energy released by the combustion of specified quantities of fuel.

The amount of energy released by the combustion of fuels depends on:

- the type of fuel burnt
- the amount of fuel burnt
- whether **complete combustion** or **incomplete combustion** is involved.

Consider the thermochemical equations for the complete combustion of methane and pentane:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -890\,\text{kJ mol}^{-1}$$
$$C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l) \qquad \Delta H = -3509\,\text{kJ mol}^{-1}$$

These equations show us that the combustion of 1 mole of pentane (C_5H_{12}) releases much more energy than the combustion of 1 mole of methane (CH_4).

When there is a plentiful supply of oxygen, complete combustion occurs and carbon dioxide is formed; however, if the oxygen supply is limited carbon monoxide is formed instead through incomplete combustion. Complete combustion of a fuel releases more energy than incomplete combustion of the same amount. For example, for ethane, complete combustion releases 1131 kJ mol^{-1} more energy than incomplete combustion:

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \quad \Delta H = -3120\,kJ$$

$$2C_2H_6(g) + 5O_2(g) \rightarrow 4CO(g) + 6H_2O(l) \quad \Delta H = -1989\,kJ$$

Heat of combustion and units of energy

The heat of combustion of a fuel is defined as the enthalpy change (ΔH), in joules, that occurs when a specified amount (e.g. 1 g, 1 L, 1 mole) of the fuel burns completely in oxygen. The SI unit for energy is the joule, symbol J. This means that the energy released (enthalpy change) can have units such as $J\,mol^{-1}$, $kJ\,mol^{-1}$, $kJ\,g^{-1}$ or $kJ\,L^{-1}$. It is usually measured at conditions of 298 K and 100 kPa, which means that the water produced should be shown in the liquid state. As 1 J of energy is a relatively small amount, the following units are commonly used.

- kilojoules, 1 kJ = 10^3 J
- megajoules, 1 MJ = 10^6 J
- gigajoules, 1 GJ = 10^9 J
- terajoules, 1 TJ = 10^{12} J
- petajoules, 1 PJ = 10^{18} J

11.1 Review

SUMMARY

- Solid, liquid and gas fuels are used primarily for manufacturing, transport and electricity generation.
- The majority of fuels used are fossil fuels, with renewable fuels a small but increasing proportion.
- World energy production and use has doubled since 1960.
- Australian oil, gas and coalfields are significant, but will be depleted by the end of this century at the current rate of use.
- Combustion of fuels uses oxygen and produces carbon dioxide water and energy (measured in joules, J).

KEY QUESTIONS

Retrieval

1 Describe a combustion reaction.

2 Convert the following energy values to kJ.
  - **a** 0.180 MJ
  - **b** 1.5×10^6 J
  - **c** 10.0 J
  - **d** 2.0×10^{-3} J

Comprehension

3 Refer to Figure 11.1.2 on page 301 to answer the following questions.
  - **a** Determine which type of fuel produced most of the world's energy in 1980.
  - **b** List the order of energy production for each fuel type (from most to least) to be produced in 2020.

4 Represent the incomplete combustion of ethanol (C_2H_5OH) when carbon monoxide is formed with a balanced equation.

Analysis

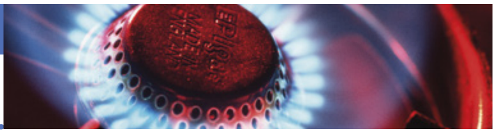
5 Wood from forests is a renewable resource that supplied global energy needs for thousands of years. Determine why wood is no longer able to meet the needs of today's society.

6 Refer to Figure 11.1.1 on page 300 and Figure 11.1.2 on page 301 to answer the following questions.
  - **a** Identify two trends in national energy consumption between 1976 and 2012.
  - **b** Explain each of these trends.

11.2 Fossil fuels

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- analyse the key uses of petrol, diesel and liquid petroleum gas.



Using engines has created the freedom to move materials around and to travel as desired. You connect to a power source numerous times a day, mostly without a thought for the prehistoric molecules, whose chemical degradation explodes within turbines and sends countless electrons surging to our appliances and devices. Fossil fuels are a reservoir of plant (and animal) hydrocarbons from long ago, having taken many years to change from living tissue to a blackened sludge-like material. This crude oil contains molecules with short and long carbon chains. In this module, you will explore the range of fossil fuels available, how they are sourced and the purpose for which they are used.

The bonds of hydrocarbons store energy from the sun.

FORMATION OF FOSSIL FUELS

Coal, oil and natural gas are non-renewable resources—they are used faster than they can be replaced. This is because they were formed from ancient plants, animals and microorganisms. Buried under tonnes of mud, sand and rock, this once biological material has undergone complex changes to become the fossil fuels used by societies today.

Coal

As wood and other plant material turn into coal, gradually the proportion of carbon content increases and the proportion of hydrogen and oxygen decreases. Over millions of years, the biomass progressively becomes peat, brown coal and then black coal (Figure 11.2.1). Some deposits are estimated to be 350 million years old. Coal is a mixture of large molecules made from carbon, hydrogen, nitrogen, sulfur and other elements.

Coal-fired power stations are the dominant generators of the world's electricity because they are often the cheapest form of generation. Electricity from coal-fired power stations is reliable and coal is currently very abundant. The combustion of coal generates over three-quarters of Australia's electricity. The **chemical energy** in coal is converted to electrical energy at a power station. Electricity is transmitted easily from the power station by metal cables and wires to other regions. The reaction occurring when coal burns can be written as:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

The energy released from the combustion of coal is about $32\,kJ\,g^{-1}$. This energy converts water to steam, which moves turbines resulting in electricity generation. The overall efficiency of a coal-fired power station is 30–40%. The combustion of brown coal is usually at the lower end of this efficiency range. Energy is lost during each step of the process, mainly as heat.

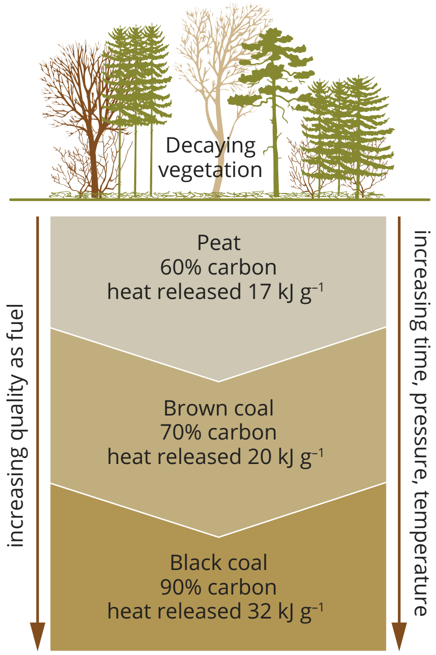


FIGURE 11.2.1 Steps in the formation of coal. Values of the carbon content and heat released upon combustion are for dried coal.

Oil

The main deposits of **crude oil** were formed from small marine animals (zooplankton) and plants (phytoplankton) that lived 1–4 billion years ago. The largest crude oil deposits are in Russia, Iran, Iraq and Saudi Arabia.

FUELS FROM CRUDE OIL

Crude oil is the source of most of the fuel used for transport (Figure 11.2.2). Crude oil is a mix of molecules mostly from the alkane homologous series. These are separated into a series of fractions (parts) by **fractional distillation**. (Figure 11.2.3). Some of these fractions are important fuels, such as liquid petroleum gas (LPG), petrol, kerosene and petrodiesel. These fractions can be used as fuels, or

FIGURE 11.2.2 Petrol is just one type of fuel from crude oil that is used each day to meet our energy needs.

as feedstock for further chemical processes. The plastics industry relies on ethene gas in the smallest fraction to make a range of plastics.

Fuels, such as petrodiesel and petrol, are mixtures of hydrocarbons of different sizes.

Fractional distillation uses heat to separate a mixture into different parts or fractions. The temperature of the tower decreases gradually with increasing height. The fractions collected from trays higher in the tower will be more **volatile** (have lower boiling points). These more volatile compounds are smaller molecules. They have lower boiling points as a result of the weaker intermolecular forces that exist between non-polar molecules of low mass. The relationship between molecular structure, intermolecular forces and boiling point is explained further in Chapter 12.

The composition and boiling range of each fraction are summarised in Figure 11.2.3.

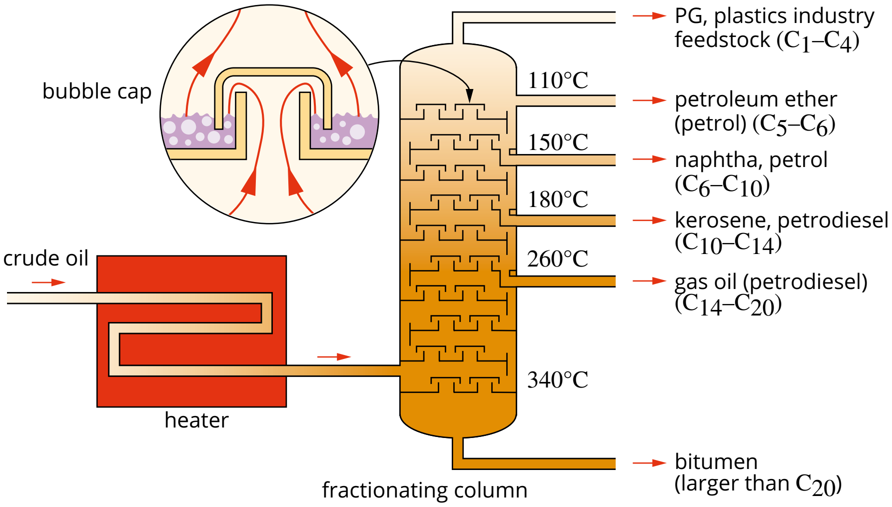


FIGURE 11.2.3 Fractional distillation of oil, and the bubble caps that slow the upward movement of gases

Petrol

The hydrocarbon combustion reactions that you probably use daily are those that occur when petroleum (known as petrol) is burnt in the engines of cars. Petrol is a mixture of **hydrocarbons**—molecules composed of hydrogen and carbon. Octane makes up a large proportion of the hydrocarbons in petrol. Other hydrocarbons include branched, cyclic and aromatic hydrocarbons of similar size to octane. The combustion reactions of these chemicals power most of Australia's 17.6 million motor vehicles. The equation for the combustion of octane is:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$$

The combustion of 1 mole of octane gas releases 5450 kJ of energy, equivalent to 47.8 $kJ\,g^{-1}$. Combustion occurs in the cylinder of a car engine. The hot gases formed push the piston in the engine, enabling the car to move. The combustion efficiency of a typical internal combustion engine ranges from about 20 to 40%.

Diesel

Diesel fuel, often called **petrodiesel**, is obtained from lower in the crude oil fractional distillation column. It is a mixture of around 75% **alkanes** containing 10 to 15 carbon atoms ($C_{10}H_{22}$ to $C_{15}H_{32}$) and 25% aromatic hydrocarbons (ring-shaped molecules). Its composition will vary depending on the source. Some of the uses of diesel in transport are in ocean liners, farming machinery and in large hauling mining trucks. Diesel has a higher density than petrol, so although the energy content of diesel and petrol, measured in $kJ\,g^{-1}$, is similar, diesel fuel yields more energy per litre.

Liquefied petroleum gas

Propane and butane gases are some of the lightest products of crude oil distillation. They become liquids under pressure and are sold as **liquefied petroleum gas (LPG)**. LPG is used as a fuel in cars and in home gas bottles. It can be sourced from fractionation of crude oil or from natural gas. Most of the vehicles that use LPG as a fuel have a standard petrol engine with a fuel tank and fuel injection system modified to suit a gaseous fuel. The equation for the combustion of propane, a major component of LPG, is:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

The combustion of 1 mole of propane releases 2220 kJ of energy, equivalent to $50.5\,kJ\,g^{-1}$.

In Australia, LPG is a significantly cheaper fuel than petrol, yet its popularity is still limited. There are many reasons for this.

- Most new vehicles are designed to run on petrol; therefore, the owner has to pay for a conversion.
- The LPG fuel tank takes up boot space.
- There are fears that LPG cylinders might explode if the vehicle crashes.
- The prices of fuels fluctuate, so often it is difficult to make meaningful price comparisons.

NATURAL GAS

Liquefied petroleum gas is a product of distillation of crude oil but similar compounds are also found as **natural gas** in deposits in Earth's crust. Natural gas is mainly composed of methane (CH_4) together with small amounts of other hydrocarbons such as ethane (C_2H_6) and propane (C_3H_8).

Natural gas is used to generate electricity for the power grid. In a gas-fired power plant, **methane** and other small alkanes are burnt to release energy. A gas-fired plant is more efficient than a coal-fired power station, reaching efficiencies just over 40%. Gas-fired plants also emit less carbon dioxide and **particulate matter** (small solid particles of solid combustion products) per unit of energy released.

11.2 Review

SUMMARY

- Fossil fuels have stored energy from biological material that is millions of years old. They are found underground and are not renewable.
- Black coal is older than brown coal and, as a result of the maturing process, contains less water and has a higher energy content.
- Oil deposits are found in rock underground or under the ocean.
- Petrol, diesel, liquid petroleum gas and other fuels are separated by fractional distillation of crude oil. The separation is based on shorter hydrocarbons having a lower boiling point.
- The uses of petrol, diesel, liquid petroleum gas and other fuels are related to their chemical properties.
- Natural gas is found near oil fields and is composed of very small molecules such as methane.

KEY QUESTIONS

Retrieval

1 Describe the basis for separation during the fractional distillation of crude oil.

2 Recall the molecule size (number of carbon atoms) of the components of diesel. Refer to Figure 11.2.3 on page 304.

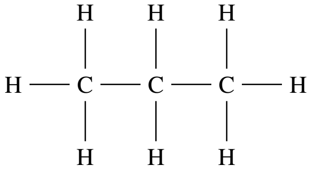
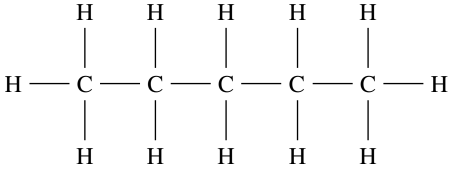
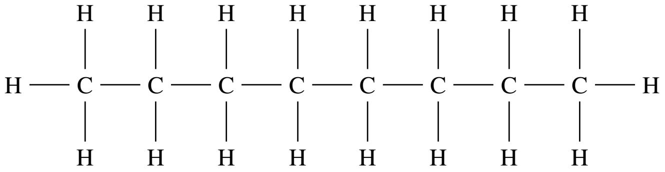
3 State why natural gas is a fossil fuel.

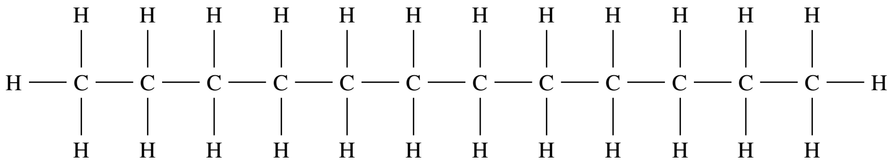
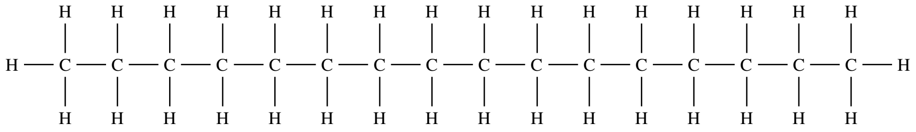
4 When liquid petroleum gas is burnt in a car engine, the major products of the combustion reaction are carbon dioxide and water.

   a Taking propane (C_3H_8) as one of the constituents of liquid petroleum gas, show a balanced equation to represent its combustion to carbon dioxide and water. Under the conditions in a car engine, propane is a gas and the water produced is also a gas.

   b Sometimes fuel does not burn completely; for example, in an engine that has not been properly tuned. This is because the oxygen supply is limited and is not sufficient to completely convert the propane to carbon dioxide and water. Under these conditions, one of the products of combustion will be carbon monoxide. Show a balanced equation to show the production of carbon monoxide when propane burns in a limited supply of oxygen. Assume that carbon monoxide is the only carbon-containing compound produced.

5 Using Figure 11.2.3 on page 304, complete the following table of information about fractional distillation fractions.

Fuels in fraction	**Number of carbons**	**Example**	**Structural formula**
		propane	
		pentane	
		octane	

		dodecane	 $\mathrm{CH_3(CH_2)_{10}CH_3}$
		hexadecane	 $\mathrm{CH_3(CH_2)_{14}CH_3}$

Comprehension

6 Explain which molecules have the stronger intermolecular forces—those collected from the base of the crude oil fractional distillation column or those collected from the top of the crude oil fractional distillation column.

7 Even though the fuel came from the ground, cars using petrol to power their engines can be described as 'solar powered'. Explain this reasoning.

8 Use Figure 11.2.1 on page 303 to answer the following questions.

- **a** Identify the type of coal that takes the longest time to form.
- **b** Describe which type of coal is a better quality fuel.

9 For each of these equations, determine if complete or incomplete combustion is shown.

- **a** $C_2H_5OH + 2O_2 \rightarrow 2CO + 3H_2O$
- **b** $C_2H_5OH + O_2 \rightarrow 2C + 3H_2O$

Analysis

10 Consider the relative volumes of the reactants and products in the combustion of octane and explain the origin of the force that moves the pistons in the engine of a car.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g) \qquad \Delta H = -10\,900 \text{ kJ}$$

11 The water content of brown coal is 60–70%.

- **a** Determine why less heat is released when the coal containing water is burnt.
  Predict the implications this high water content has for the energy released from burning the coal.
- **b** Describe a pre-treatment that could raise the energy output per gram of brown coal consumed in a power station.
- **c** Describe two ways that the fuel could be treated to reduce the impact of impurities on the environment.

12 Coal seam gas has become a significant energy source in Queensland since the year 2000.

- **a** Identify the main component of coal seam gas.
- **b** Identify two environmental concerns associated with the production of coal seam gas.

13 Carbon dioxide gas and sulfur dioxide gas are products of coal-fired power stations. Typical properties of various coals are shown in the table below.

- **a** Calculate the volume of each of these gases produced by 1 kg each of anthracite coal and bituminous coal using the data in the table below. Use the average value of each data range.
- **b** Research the cause and effects of sulfur dioxide acid rain.

	Anthracite coal	**Bituminous coal**	**Lignite coal**
fixed carbon, weight %	80.5–85.7	44.9–78.2	31.4
moisture, weight %	2.8–16.3	2.2–15.9	39.0
ash, weight %	9.7–20.2	3.3–11.7	3.3–11.7
sulfur, weight %	0.6–0.77	0.7–4.0	0.3

11.3 Biofuels

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- distinguish between renewable and non-renewable fuels
- understand the importance of biofuels in managing carbon dioxide in the atmosphere
- understand the chemical nature of bioethanol, biodiesel and biogas.

Governments and industry are exploring alternatives to fossil fuels to meet our future energy needs and limit the impact of burning fossil fuels on the environment. Ideally, new sources of energy will be renewable. Renewable energy is energy that can be obtained from natural resources that can be constantly replenished, at least at the rate that it is used.

Biochemical fuels (or **biofuels**) are fuels derived from animals, such as animal fat, or plant materials, such as grains (corn, maize, wheat, barley or sorghum), sugar cane, vegetable waste and plant oils. The three main biofuels are bioethanol, biodiesel and biogas. They can be used alone or blended with fossil fuels such as petrol and diesel. As well as being renewable, biofuels are predicted to have fewer detrimental effects on the environment than do fossil fuels.

RENEWABILITY

Fossil fuels are a non-renewable source of energy. It took millions of years for their formation, so no new deposits will be usable in the near future. Biofuels are renewable (able to be replaced by natural processes within a relatively short period of time). This means that their availability can be ongoing into the future. Renewability is explored in detail in Unit 4.

The plant materials used in the generation of biofuels are produced by **photosynthesis**, which removes carbon dioxide from the atmosphere and produces glucose ($C_6H_{12}O_6$) in the reaction:

$$6CO_2(g) + 6H_2O(l) \xrightarrow[\text{chlorophyll}]{\text{sunlight}} C_6H_{12}O_6(aq) + 6O_2(g)$$

Plants convert the glucose into cellulose and starch. Although carbon dioxide is released back into the atmosphere when the biofuel is burnt, the net impact should be less than for fossil fuels.

From this point of view, you see that biofuels can be more **carbon neutral** than fossil fuels and contribute less to **climate change**. However, other environmental concerns exist, such as the use of agricultural land required for food, and the further clearing of native forests and bushland to produce biofuels. In addition, the energy used during the manufacture of these fuels is a cost of production that must be considered to fully understand the benefits of biofuels. In this module, you will explore the range of biologically derived fuels, how they are sourced and the purpose for which they are used.

Currently, biofuels produce only a small percentage of Australia's fuel needs but other countries have made significant advances in their use. In the European Union, for example, 10% of European vehicles are expected to run exclusively on biofuels by 2020. If biofuel production in Australia were to increase significantly, crops would need to be grown specifically or processes to harvest ethanol from waste organic matter would need to be developed.

Some of the non-renewable and renewable fuels in use in Australia are listed in Table 11.3.1.

TABLE 11.3.1 Types of renewable and non-renewable fuels in use in Australia

Non-renewable fuels	Renewable fuels
coal	bioethanol
oil	biodiesel
liquefied petroleum gas (LPG)	biogas
natural gas	
coal seam gas (CSG)	

BIOETHANOL

For thousands of years, humans have employed biological catalysts (enzymes) from yeasts to convert starches and sugars to ethanol. Enzymes catalyse the breakdown of the starch in grain crops (such as corn, barley and wheat, Figure 11.3.1) to glucose. Other enzymes in yeasts then convert glucose and other small sugar molecules to ethanol (CH_3CH_2OH) and carbon dioxide in the **fermentation** reaction:

$$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast enzymes}} 2CH_3CH_2OH(aq) + 2CO_2(g)$$

Bioethanol can be produced in this way from crops such as sugar cane and corn (maize). Sugar cane produces a slightly greater amount of ethanol than corn. The Pentland Bioenergy plant under construction in Queensland will begin ethanol production in 2018, and production is planned to be 344 million litres per year by June 2020.

FIGURE 11.3.1 Harvesting sugar cane in Queensland. Sugar cane is a source of the raw materials for the production of bioethanol.

Bioethanol is used extensively in Australia. This includes **E10 petrol blend**, which contains 10% ethanol. This mix can be used by most modern car engines and its use reduces the consumption of petrol derived from crude oil. The presence of ethanol also reduces the emissions of particulates and gases such as oxides of nitrogen, but higher levels of ethanol can damage engines, especially in older vehicles.

The equation for the complete combustion of ethanol is:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

The combustion of 1 mole of ethanol releases 1367 kJ of energy, equivalent to 29.7 $kJ\,g^{-1}$. As you will see in Module 11.4, the energy content of ethanol is about 62% that of the main component of petrol—octane. This means that a larger mass (and volume) of ethanol is required to provide the same amount of energy. At a simple level, the lower energy content of ethanol can be regarded as the result of the carbon atoms in an ethanol molecule being partly oxidised ('partly burnt'). This is due to the presence of oxygen in the ethanol molecule.

BIODIESEL

Biodiesel is a mixture of organic compounds called **esters**. These esters are produced in a process, known as **transesterification**, using vegetable oils or animal fats and an alcohol (most commonly methanol, CH_3OH). The composition of biodiesel varies due to differences in types of raw materials, which can be oils from sources such as soya bean, canola or palm oil. The structure of a typical biodiesel molecule is shown in Figure 11.3.2.

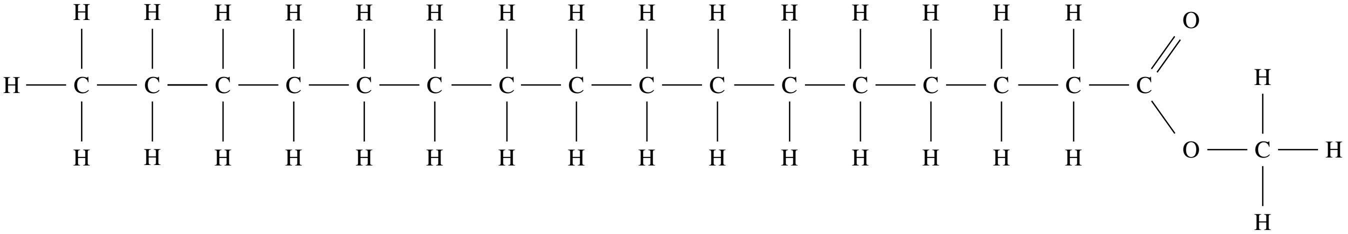


FIGURE 11.3.2 Structural formula of a typical biodiesel molecule

The equation for the combustion of biodiesel is shown below. The reaction releases about 10% less energy than diesel, but has a safety advantage of a higher **flashpoint** (temperature at which a vapour forms) than diesel.

$$2C_{17}H_{34}O_2(l) + 49O_2(g) \rightarrow 34CO_2(g) + 34H_2O(g)$$

BIOGAS

Biogas is gas that is released in the breakdown of organic waste by anaerobic bacteria. These bacteria decompose the complex molecules contained in substances such as carbohydrates and proteins into the simple molecular compounds carbon dioxide and methane.

A range of materials including rotting rubbish, sewage and decomposing plant material can be used to produce biogas.

The composition of biogas depends on the original material from which it is obtained and the method of decomposition. The typical composition of a sample of biogas is shown in Table 11.3.2.

TABLE 11.3.2 Typical percentage composition of different molecules found in biogas

Gas	Formula	Percentage composition (by volume)
methane	CH_4	60
carbon dioxide	CO_2	32
nitrogen	N_2	4.5
hydrogen sulfide	H_2S	2
oxygen	O_2	1
hydrogen	H_2	0.5

As you can see from Table 11.3.2, biogas consists mainly of methane and carbon dioxide. Biogas can be used for heating and to power homes and farms.

The main reaction occurring in the combustion of biogas is the same reaction of methane burning in a gas-fired power station.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

WS 1.3.6

The energy released per gram of biogas is less than that of natural gas because the methane content in biogas is lower.

SCIENCE AS A HUMAN ENDEAVOUR

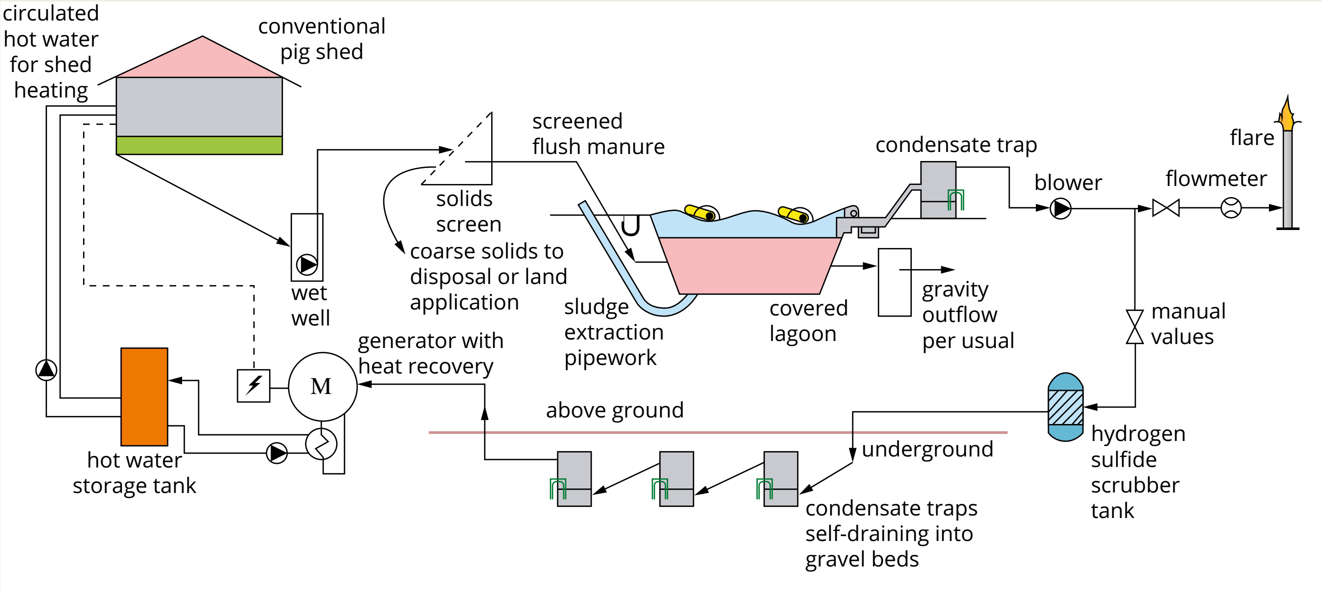
Biogas generation and use at a piggery

The Queensland Natural Pork Holdings Palahra piggery at Grantham has been producing and using biogas since 2010. A moderate-sized piggery, it has reduced the cost of its LPG gas use by $23 000 annually by setting up a 'covered pond' in which waste from the piggery produces biogas, which is about 70% methane gas. The process involves digestion of the waste (mostly manure) by anaerobic bacteria. This procedure enables the production of fuel to warm the animal areas and digested effluent suitable as a fertiliser. Importantly, this technique of collecting methane is able to significantly reduce the greenhouse emissions of the farm because methane is a potent greenhouse gas—its strength is 21 times that of carbon dioxide.

Covering areas where waste is collected is a practice adopted by farmers to manage emissions of potent gases. Often, when systems to collect the biogas are not yet in place, the excess biogas under the pond cover is 'flared'. This means that it is burnt so that the large amounts of methane are converted to carbon dioxide. While not using the biogas efficiently, this process will, at least, mean that methane does not enter the atmosphere. Biogas from animal refuse contains sulfur, in the form of hydrogen sulfide. Sulfur dioxide produced by its combustion causes acid rain. So, the biogas is passed through a 'scrubber tank' to remove this component before the combustion stage at the water heater. A schematic diagram and image of the pond set-up is shown in Figure 11.3.3.

The biogas from the extraction system (covered pond, Figure 11.3.4) is supplied to a heating system used to heat water circulated through underfloor heating pads in the piggery farrowing sheds. This gas-fired water-heating system reduces on-farm energy use and costs by replacing a significant proportion of the LPG previously used for farrowing shed heating. Larger farms can cost-effectively install generators to produce electricity from the gas (Figure 11.3.5 on page 312). In this way, the biofuel has even greater use including supply of electricity back to the grid.

FIGURE 11.3.4 A covered pond for biogas generation



*A coarse solids screening step may not be in use

FIGURE 11.3.3 A schematic representation of biogas generation and use at a piggery

Review

1 State the chemical process that digests waste to methane.

2 Explain how the use of biogas is of benefit to the environment.

3 Converting methane produced by piggery waste to carbon dioxide decreases the effect of gases on the environment. Explain this statement by referring to the chemical equation for combustion of methane.

4 Research the properties of hydrogen sulfide (H_2S) gas and propose a reason why piggeries with covered ponds are more acceptable to local communities.

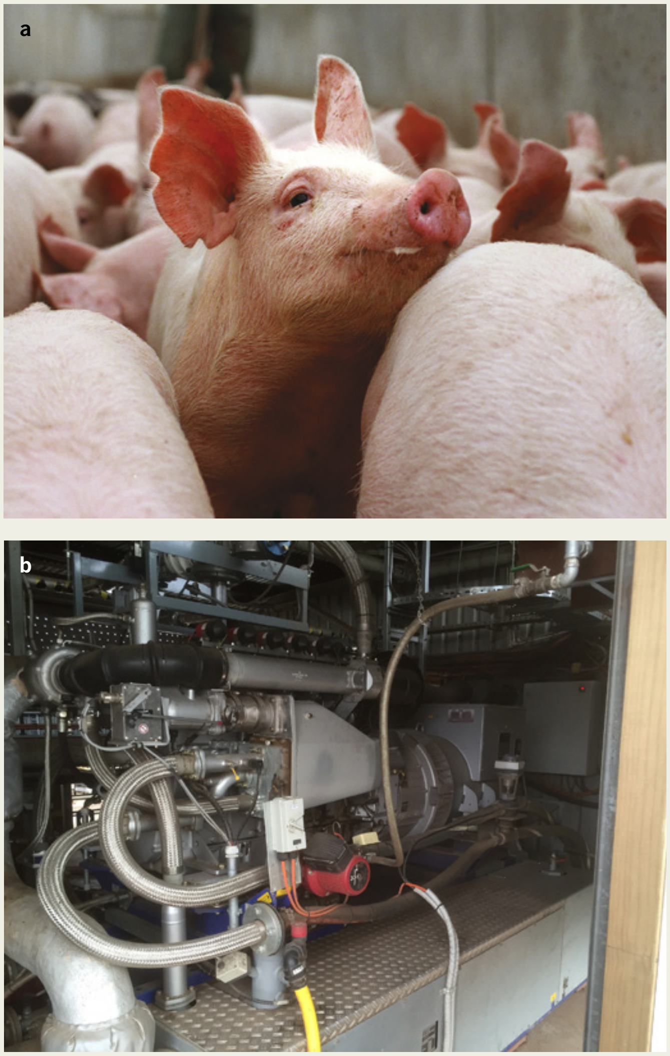


FIGURE 11.3.5 (a) Pig waste is a source of biogas. (b) Biogas can be used to fire generators that produce electricity from methane in the gas.

11.3 Review

SUMMARY

- Biofuels are renewable because their hydrocarbon molecules can be replenished.
- Unlike fossil fuels, biofuels are mostly carbon neutral because the carbon dioxide released during combustion was previously sourced from the atmosphere as the plant photosynthesises.
- Bioethanol can be used as a proportion of car fuel, and as a 100% fuel after conversion of the engine.
- Enzymes in plants catalyse fermentation, which produces bioethanol from starch and sugar (and more recently cellulose).
- Biodiesel is produced by reacting biologically derived oil with alcohol and is readily exchanged for petrodiesel in engines.
- Biogas is a result of anaerobic digestion of waste and is produced at rubbish tips.

KEY QUESTIONS

Retrieval

1 Describe the difference between a renewable fuel and a non-renewable fuel.

2 State the name of the reaction in which a triglyceride molecule is broken down to form biodiesel molecules.

3 State the following statements about biodiesel are true or false.

- **a** Biodiesel can contain several different compounds.
- **b** Biodiesel molecules contain ester groups.
- **c** Biodiesel is an alkane.
- **d** Three molecules of biodiesel can be obtained from each molecule of a triglyceride.
- **e** All hydrocarbon chains in biodiesel molecules are saturated.
- **f** Biodiesel is formed in a transesterification reaction.

4 Identify the two gases that form the largest percentage of a typical biogas sample.

5 State three types of vegetable oil that can produce biodiesel.

Comprehension

6 Use the definition of renewability to explain why alternatives to fossil fuels must be found.

7 Describe examples of renewable fuel sources and non-renewable fuel sources used in Australia.

8 Describe a renewable and a non-renewable source of methane.

9 Explain why there are variations in the percentage of gases making up different samples of biogases.

Analysis

10 Determine which of Australia's resources is likely to last longer before it is depleted—coal, oil or natural gas.

11 Determine why CO_2 emissions from the use of biodiesel are not considered as problematic as those produced from the use of petrodiesel.

12 The following energy sources are used across the world: coal, bioethanol, biodiesel and natural gas. Identify which of these are likely to be available for the longest time.

13 Determine why bioethanol is sometimes described as a 'carbon neutral' fuel. Use chemical equations for photosynthesis, fermentation and combustion to support your answer.

14 Trials are being conducted to source biodiesel from algae grown in the warm water of a power station cooling pond. The water is warm and carbon dioxide emitted from the power station can be trapped and bubbled through the water to enhance the growth of the algae. Classify the following as advantages or disadvantages of large-scale production of biodiesel from algae.

- **a** fewer particulate emissions than petrodiesel
- **b** renewable
- **c** less reliance on fossil fuels
- **d** distribution outlets required
- **e** fewer net CO_2 emissions than petrodiesel
- **f** oil can be 'harvested' many times per year
- **g** reduction of CO_2 emissions from coal-fired power stations
- **h** more additives need to be added to the fuel in cold climates

15 Analyse this statement by explaining the two rates that are mentioned.
'A non-renewable energy source cannot be produced at the same rate as the rate at which it is used by society.'

11.4 Comparing fuels

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- use information from stoichiometric calculations learnt in Chapter 10 to compare heat of combustion of fuels using alcohols as examples
- interpret energy density data to compare the energy content of fuels
- distinguish key differences between liquid petroleum gas and petrol
- distinguish key differences between petrodiesel and biodiesel
- distinguish key differences between octane and ethanol
- analyse factors regarding fuels to evaluate which fuel is the most suitable for a purpose.

If you were to start a small business you would make a multitude of decisions about the structures and processes needed to make it successful. The use of energy would be central to most of these. What type of building, equipment and machinery would you use or build? Where would it be located? How would you transport materials and/or people? On a larger scale, governments need to make similar decisions on behalf of the nation. For the past 100 years, the choice of type of energy was limited. Fuels used were found quite easily in the ground, and power derived from these fossil fuels was reasonably cheap. Both situations have changed, and the undeniable warming of the Earth provides us with a challenge to manage our use of fuels more effectively. Earlier in the chapter, you learnt about the nature of fossil fuels and the ever-increasing types of biofuels. In this module, you will use chemistry to compare these fuels, focusing on suitability for purpose, energy produced and impact on the environment.

COMPARING FUELS USING STOICHIOMETRIC CALCULATIONS

During combustion, the conversion of a fuel and oxygen to carbon dioxide and water releases heat known as the **heat of combustion** (ΔH_c). The reaction is usually rapid and the heat of combustion is usually measured at conditions of 298 K and 100 kPa. If conditions do not provide sufficient oxygen, the reaction is called incomplete combustion and the key product is carbon monoxide and sometimes soot (C). The enthalpy change of the fuel, overall, is negative, which indicates that energy is released to the environment. These ideas have been described in Chapter 10. For pure substances, the unit is kJ mol^{-1}. However, fuel mixtures (such as diesel or kerosene) do not have a specific chemical formula and so units such as kJ g^{-1}, kJ L^{-1} or MJ t^{-1} must be used.

Only fuels that exist as pure substances can have their heat of combustion measured in kJ mol^{-1}.

Heat of combustion

In Chapter 10, you learnt how to use a thermochemical equation to represent the energy change (heat of combustion) for fuels involved in combustion reactions. A thermochemical equation includes a ΔH_c value, which shows the amount of heat released when a specific amount of a fuel (shown by the stoichiometric ratio in the equation) undergoes combustion.

Calculating energy change from thermochemical equations

Worked examples 11.4.1 and 11.4.2 show you how to use a thermochemical equation to calculate the energy released by the combustion of different amounts of fuels or the amount of fuel required to produce a specified amount of energy.

Worked example 11.4.1

CALCULATING ENERGY RELEASED BY THE COMBUSTION OF A PARTICULAR AMOUNT OF A FUEL

Calculate the heat energy released, in MJ, when 10.0 kg of octane undergoes complete combustion.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l) \; \Delta H = -10\,900\text{ kJ}$$

Thinking	**Working**
Calculate the number of moles of the fuel using: $n = \frac{m}{M}$	$n(C_8H_{18}) = \frac{10000}{114.26}$ $= 87.5$ mol
Using the thermochemical equation, write a relationship between the number of moles of fuel burned and energy released.	2 moles of C_8H_{18} releases 10 900 kJ. Let 87.7 moles of C_8H_{18} release x kJ.
Calculate the energy released by the fuel in kilojoules.	By proportion: $\frac{87.5}{2} = \frac{x}{10\,900}$ $x = \frac{87.5}{2} \times 10\,900 = 476\,875$ kJ
Convert the energy released to MJ. 1 MJ = 10^3 kJ	476 875 kJ = 477 MJ

Try yourself 11.4.1

CALCULATING ENERGY RELEASED BY THE COMBUSTION OF A PARTICULAR AMOUNT OF A FUEL

Calculate the heat energy released, in MJ, when 10.0 kg of ethane undergoes complete combustion.

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \; \Delta H = -3120\text{ kJ}$$

Worked example 11.4.2

CALCULATING THE AMOUNT OF FUEL THAT MUST BE BURNED TO PRODUCE A PARTICULAR AMOUNT OF ENERGY

Calculate the volume of methane, measured at standard temperature and pressure (STP), that burns completely to provide 4.00×10^4 kJ.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \; \Delta H = -890\text{ kJ mol}^{-1}$$

Thinking	**Working**
Using the thermochemical equation, write a relationship between the number of moles of fuel burned and energy released.	1 mole of CH_4 releases 890 kJ. Let x moles release 4.00×10^4 kJ.
Calculate the amount of the fuel that was burnt to produce the energy.	By proportion: $\frac{1}{x} = \frac{890}{4.00 \times 10^4}$ $x = \frac{4.00 \times 10^4}{890}$ $= 44.9$ mol
Calculate the volume of the fuel at STP.	$n = \frac{V}{V_m}$ $V(CH_4) = n \times V_m$ $= 44.9 \times 22.7$ $= 1019$ L $= 1.09 \times 10^3$ L

➤ Try yourself 11.4.2

CALCULATING THE AMOUNT OF FUEL THAT MUST BE BURNT TO PRODUCE A PARTICULAR AMOUNT OF ENERGY

Calculate the volume of methane, measured at STP, that burns completely to provide 5.00×10^3 kJ.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890\ kJ\ mol^{-1}$$

The heats of combustion for some common elements and compounds present in fuels are listed in Table 11.4.1. Heat energy is released during combustion, so ΔH_c always has a negative value.

TABLE 11.4.1 Heats of combustion for some common fuels

Substance	Heat of combustion, ΔH_c (kJ mol^{-1})
methane	–890
ethane	–1560
propane	–2220
butane	–2886
octane	–5450
methanol	–725
ethanol	–1367
propanol	–2021
pentanol	–3331
dodecane	–7902
methyl linoleate	–11 690

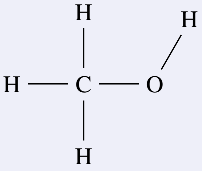
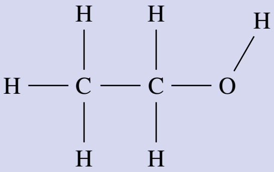
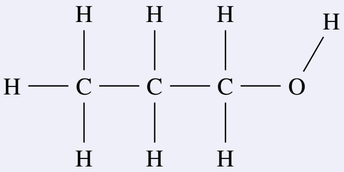
The conversion between mole, mass and volume amounts is an important skill in analysing fuels using thermochemical equations. This allows us to compare the amount of energy per mole, per gram or per litre. Some of these conversions are described in Chapter 9 and Chapter 14.

ΔH_c trends in three alcohols—effect of size of fuel molecule

An experiment using a calorimeter and three spirit burners containing fuels of the alkanol (Figure 11.4.1) series allows confirmation of the heats of combustion, and illustrates the link between fuel structure and energy produced. Calorimetry experiments, described in Chapter 10, are investigations where the heat from a burning fuel is used to raise the temperature of a known mass of water. Information related to the experiment is shown in Table 11.4.2.

FIGURE 11.4.1 Combustion of 1 g of ethanol in these spirit burners releases almost 30 kJ of energy.

TABLE 11.4.2 Structural formulas and thermochemical equations for some alcohol fuels

Fuel	Structure	Thermochemical equation
methanol		$CH_3OH(g) + 1.5O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = -725\,kJ\,mol^{-1}$
ethanol		$C_2H_5OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ $\Delta H = -1367\,kJ$
propanol		$C_3H_7OH(g) + 4.5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$ $\Delta H = -2021\,kJ$

The trend in heats of complete combustion for these three alcohols is clear. As the number of carbons in the fuel increases, the heat produced (ΔH_c) also increases. You can see that, for each single increase in number of carbons in the chain, there is an increase of about 600 kJ in the heat of combustion. This trend can be explained, using bond energies for each bond broken and formed, through the understanding that bond formation is exothermic and bond breaking is endothermic (refer to Worked example 10.2.2 on page 277).

The key to this trend is the formation of the carbon dioxide molecule. The bond energy of the C=O bond is high (799 kJ). As larger fuel molecules will form one extra CO_2 molecule for each extra carbon in the fuel molecule, larger molecules will have higher ΔH_c values due to the relatively high energy of the C=O bond (799 kJ) being formed.

Energy density

Liquid fuels, such as petrol, are normally sold by volume rather than by mass. They are also mixtures, so for these fuels it is often more convenient to refer to the heat of combustion per litre or per gram, rather than per mole. Figure 11.4.2 compares volumetric heats of combustion (energy per litre) and gravimetric heats of combustion (per kilogram) for common liquid fuels. The energy released per litre of fuel is often called the fuel's **energy density**.

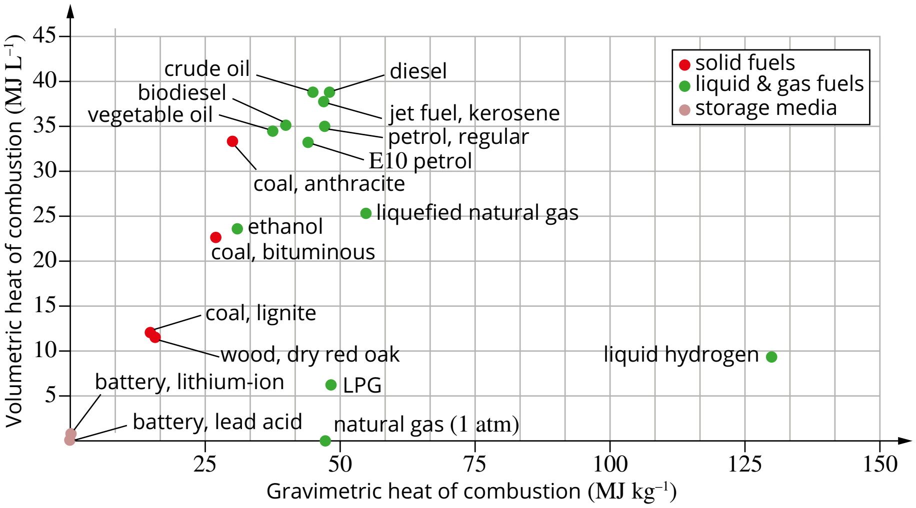


FIGURE 11.4.2 Volumetric and gravimetric heats of combustion of some common fuels

COMPARING PETRODIESEL AND BIODIESEL

Diesel engines are favoured in situations requiring high fuel efficiencies and for heavy haulage vehicles. They have a 20 to 40% better fuel economy than petrol engines. Although its inventor described that the diesel engine (produced in 1895) could operate using vegetable oil, it is only in the past 20 years that biodiesel has emerged as an alternative to petrodiesel. At present, biodiesel consumption is about a tenth that of petrodiesel, both locally and internationally. This gap is likely to close in the future as the biodiesel industry is the subject of extensive scientific research.

The typical structures for biodiesel and petrodiesel (which are typical fractions of the mixture) are shown in Figure 11.4.3. The main combustion products of both forms of diesel are carbon dioxide and water as represented by the equations:

$$\text{petrodiesel: } 2C_{12}H_{26}(l) + 37O_2(g) \rightarrow 24CO_2(g) + 26H_2O(g)$$

$$\text{biodiesel: } 2C_{17}H_{34}O_2(l) + 49O_2(g) \rightarrow 34CO_2(g) + 34H_2O(g)$$

The carbon dioxide emissions from both petrodiesel and biodiesel are about $73\,g\,MJ^{-1}$, which is almost the same as for petrol. In the case of biodiesel, carbon dioxide is absorbed in the growth of the plant or animal the biodiesel is derived from, reducing the net impact on **greenhouse gas** levels. Biodiesel production and use upholds green chemistry principles but petrodiesel use follows fewer of these principles.

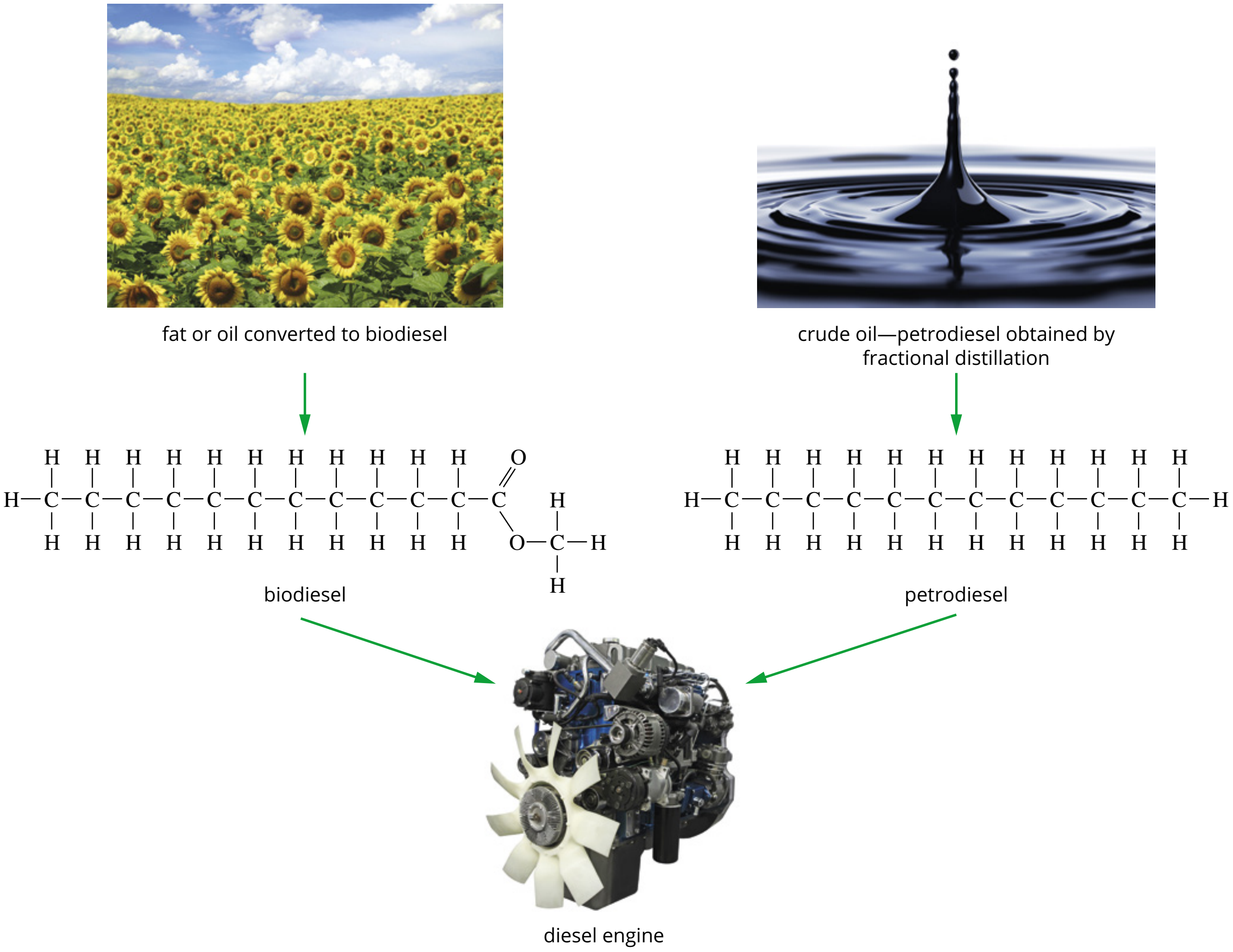


FIGURE 11.4.3 Two methods of producing diesel fuel, and the typical structures of each fuel

Some of the properties of petrodiesel and biodiesel are compared in Table 11.4.3

TABLE 11.4.3 Properties of petrodiesel and biodiesel compared

Property	Explanation of property	Petrodiesel	Biodiesel
density (kg L^{-1})	mass of fuel in a given volume	0.84	0.87
viscosity (cSt)	measure of a fluid's resistance to flow at 40°C	2.5	4.9
energy content (kJ g^{-1})	energy released per gram of fuel	48	41
energy density (kJ L^{-1})	energy released per litre of fuel	40	36

Table 11.4.3 shows that petrodiesel produces up to 10% more energy than biodiesel (this is offset by the fact that biodiesel burns more smoothly and efficiently). Biodiesel has a higher **viscosity** than petrodiesel, and does not flow as easily along fuel lines and through filters. This is related to stronger intermolecular forces in biodiesel, and renders biodiesel less useful than petrodiesel.

An overall comparison of the two fuels is provided in Table 11.4.4. For the motorist, the differences in the properties of the two fuels imply that:

- diesel blends containing up to 20% biodiesel have a similar performance to petrodiesel and the engines do not need to be modified
- in cold climates, high-biodiesel fuels require additives and service stations need to ensure the pump is kept warmer than the surroundings.

Biodiesel also has the potential to keep fuel systems cleaner, but higher quality seals and fuel hoses may be required. Biodiesel cannot be stored for as long as petrodiesel because it is more likely to react with atmospheric water and oxygen, and biodegrade.

TABLE 11.4.4 An overall comparison of petrodiesel and biodiesel

Property	Petrodiesel	Biodiesel
renewability	non-renewable	renewable
energy content	marginally higher	marginally lower
viscosity	lower	higher
emissions	CO_2, particulates and other pollutants are produced.	CO_2 is produced but CO_2 is consumed in production of the triglycerides used to make the fuel, making it more carbon neutral. Biodiesel also produces fewer particulates and other pollutants.
environmental impact of production	Oil fields are prone to spills and damage to local habitats. Oil refineries produce emissions that have harmful effects on local regions.	Spilt biodiesel will biodegrade. Using farm crops can add to land degradation and might lead to higher food prices. Biodiesel made from waste products, such as used cooking oil, has beneficial effects for the environment.

COMPARING OCTANE AND ETHANOL

Octane (C_8H_{18}) and ethanol (C_2H_5OH) are used as fuels for transport. Octane is a key component of petrol and is derived from a fossil fuel, whereas ethanol is a biofuel. Comparing these two fuels provides an insight into the use of E10 petrol and informs consideration of the further use of bioethanol in cars. In Brazil, the flourishing sugar cane industry supports cars running on 85% and 100% ethanol.

You have seen how energy produced per gram or kilogram is calculated and fuels are compared (Worked example 11.4.1). Of equal importance is to identify the effect on the environment by calculating the amount of carbon dioxide, a greenhouse gas, produced as energy is released. The thermochemical equations for each reaction (shown below) are used to calculate the amounts of these gases.

octane: $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g) \quad \Delta H = -10\,108\,kJ$

ethanol: $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \quad \Delta H = -1235\,kJ$

Calculations for octane are shown below. A comparison of each of these quantities for octane and ethanol is given in Table 11.4.5.

Volume of CO_2 (a greenhouse gas) released at STP per MJ of energy obtained

From the equation:

16 mol of CO_2 is released when 10 108 kJ of energy is released.

Calculate the volume occupied by 16 mol of CO_2 at STP:

$$n = \frac{V}{V_m}$$

$$\text{so } V = n \times V_m$$

$$\text{at STP } V(CO_2) = 16 \times 22.7$$

$$= 363.2\,L$$

So, production of 363 L of CO_2 releases 10 108 kJ (10.108 MJ).

Let the production of x L of CO_2 release 1 MJ of energy.

By proportion:

$$\frac{x}{363} = \frac{1}{10.108}$$

$$x = \frac{363}{10.108}$$

$$= 35.9\,L$$

Complete combustion of octane to produce 1 MJ of energy produces 35.9 L of CO_2, measured at STP. A similar calculation can be performed for the other product, H_2O, which is also a greenhouse gas.

TABLE 11.4.5 Quantities calculated for the complete combustion of octane and ethanol

Fuel	Complete combustion of octane	Complete combustion of ethanol
energy content (MJ kg^{-1})	44.3	26.8
volume of CO_2 produced per MJ of energy (L MJ^{-1})	35.9	36.8
total volume of greenhouse gases (CO_2 and H_2O) produced per MJ of energy at 500 K and 1 atm (L MJ^{-1})	138	166

The data in Table 11.4.5 shows that octane has a greater energy content (MJ kg^{-1}) than ethanol. This means that octane is more useful than ethanol as a portable transport fuel. Their carbon dioxide emissions per megajoule of energy produced are similar. However, octane produces less greenhouse gases overall. Also, in engines ethanol burns smoothly and produces fewer particulates than petrol. Today, very few cars are not readily able to use petrol containing ethanol.

Comparing petrol and LPG

Liquid petroleum gas (LPG) and petrol are both mixtures consisting almost entirely of alkane molecules, but having a key difference that petrol is a liquid and LPG is a gas stored as a liquid under pressure in canisters. LPG is a mixture of propane and butane (C_3H_8 and C_4H_{10}), and petrol contains molecules that are slightly larger (C_8H_{18} to $C_{12}H_{26}$). LPG is versatile as a fuel being used for heating, cooking, refrigeration and in transport.

In cars, LPG is less efficient than petrol, due to a lower heat of combustion per mole. This is due to the difference in size of the molecules in the fuel mixture. Molecules in LPG are smaller than petrol, and so, using the reasoning described above for alcohols, petrol has a greater energy density (35 $MJ\,L^{-1}$) than LPG (25 $MJ\,L^{-1}$). Petrol's combustion, however, produces more particulate emissions (polluting particles) than burning LPG. Another concern is that, as LPG is heavier than air, it will not disperse if leakage occurs in a walled area even if windows are open, creating a safety hazard. For this reason, LPG cars are not allowed in some indoor parking areas.

As crude oil distillation is occurring in massive amounts there is always a ready supply of LPG gas. However, a product of LPG processing, ethene, is feedstock for the plastics industry and therefore large quantities are used elsewhere. Transportation of LPG gas is often in tankers or in canisters as liquid under pressure. However, it will flow readily along a pipeline from the oil refinery to, for example, a marine terminal. This removes the need for road transport during this stage.

CHOOSING A FUEL

If you were a business owner, the chemistry-based comparisons of liquid and gas fuels explained above would be useful in deciding which of these fuels to use. Perhaps the effect on the environment of both fuel use and fuel sourcing would also be a consideration. Further practical imperatives might inform your choice, such as how readily the fuel is transported and whether it can be used in existing machines and infrastructure. While you are able to make a decision about which fuel to put in your car or use in a business, similar decisions need to be made on a national and global scale. Sustainable alternatives that do not contribute to the enhanced greenhouse effect are required.

There are many factors that need to be considered when selecting a fuel for a particular purpose.

- What is the energy released per unit mass or unit volume?
- What mass of greenhouse gases is produced per energy released?
- Is the source of the hydrocarbons renewable or non-renewable?
- What is the availability and cost of the fuel?
- Is it easy to transport?
- Are there hazards to people and the environment related to sourcing and production of the fuel?
- Are there hazards to people and the environment associated with its use and its waste products?
- What are the social, economic, cultural and political values that can affect the choice of a fuel?

Table 11.4.6 is an overview of some key chemistry-related factors relating to evaluating the use of fuels for a specific purpose.

TABLE 11.4.6 Choosing a fuel—consideration of advantages and disadvantages of some fuels

Fuel	Advantages	Disadvantages
coal	• large reserves • relatively high energy content	• non-renewable • high level of emissions • less easily transported than liquid or gaseous fuels
natural gas and coal seam gas	• more efficient than coal for electricity production • easy to transport through pipes • relatively high energy content	• non-renewable • limited reserves • polluting (but less than coal and petrol production) • fracking may pollute water reserves
biogas	• renewable • made from waste • reduces waste disposal • low running costs • releases the CO_2 that was absorbed previously in photosynthesis	• low energy content • supply of waste as raw materials is limited
petrol	• high energy content • ease of transport	• non-renewable • polluting (but less than coal) • limited reserves
LPG	• low cost • easily separated from natural gas • relatively high energy content • fewer particulates produced than petrol	• non-renewable • polluting (but less than petrol)
bioethanol and biodiesel	• renewable • can be made from waste • releases the CO_2 that was absorbed previously in photosynthesis • burns smoothly • fewer particulates produced than petrol • can be used in existing liquid fuel-based systems	• limited supply of raw materials from which it can be produced • lower energy content than petrol • may require use of farmland otherwise used for food production

A discussion of the environmental impact of fuels needs to consider both the impact on the environment when obtaining the fuels and the effect of the greenhouse gas emissions from the combustion of the fuel. All fuels have to be mined or produced. The decrease in photosynthesising biomass as a result of deforestation continues to increase the amount of carbon dioxide in the atmosphere. If plants or algae are to be used to produce biofuels, further land may be required. Table 11.4.7 is a summary of practical considerations relating to the sourcing and producing of fossil fuels compared to biofuels.

TABLE 11.4.7 A comparison of the impacts of sourcing and producing fossil fuels and biofuels

	Fossil fuels	Biofuels
effects of sourcing	 • Fossil fuels can be mined on land or at sea. • Fossil fuels are non-renewable. • Oil and coal deposits are often located in environmentally sensitive locations and mining approval is often contested. • Spills, leaks or run-off could significantly damage local ecology. • Fossil fuels are flammable with associated fire risks.	 • Increased biofuel production requires intensive farming. • Crops and algae are renewable because they can be harvested each year. • Growing crops requires resources such as water and fertiliser. The quality of the soil deteriorates if the same crops are grown each year. • Farm machinery produces CO_2 and other pollutants that reduce the net benefits of renewable fuels. • High-quality farmland may be used for growing fuels instead of food.
effects of production	 • Oil refineries can produce significant emissions due to the burning of excess hydrocarbons. • Energy is required to operate the refinery. • Energy is required to transport coal, gas and oil.	 • Biofuel plants should produce fewer emissions than oil refineries. • Energy is required to operate the plant and transport the fuel.

In the immediate future, the growth of biofuel production may depend upon economics. If biofuel manufacturers cannot produce a product, ready for use in existing infrastructure, for a price lower than the market price of fossil fuels, then the industry is unlikely to expand. A complicating factor is the fluctuations that occur in fuel prices. Global and national initiatives will be required to support the industry in the short to medium term.

11.4 Review

SUMMARY

- Thermochemical equations indicate the stoichiometry of the combustion reaction and energy produced; this heat of combustion is different for different fuels.
- Calculations using bond energies and thermochemical equations can predict the energy released from combustion, and reveal a link between fuel molecule size and heat of combustion (e.g. in fuels of the alkanol series).
- The energy density ($J\,L^{-1}$ or $J\,kg^{-1}$) varies for different fuels.
- Diesel (a hydrocarbon) has greater energy density than biodiesel due to differences in molecular structure.
- Diesel and biodiesel have other differences in physical properties and performance related to differences in their chemical structure and related intermolecular forces.
- Thermochemical equations of fuel combustion can also be used to compare the volume of exhaust gases produced per amount of fuel or per amount of energy produced; for example, for octane and ethanol.
- Petrol has a higher energy density than liquid petroleum gas because its main component (octane) is a larger molecule.
- Choice of a fuel involves many considerations including energy produced per kg or per L, gases produced, whether the fuel is sustainable and the effect on the environment (including effect of production/mining and effect of fuel use).

KEY QUESTIONS

Retrieval

1 Diesel, petrol and ethanol are used to power vehicles.
List the fuels in order of their energy content per gram (from highest to lowest). You may need to refer to Figure 11.4.2 on page 317.

2 Define the term 'energy efficiency'.

Comprehension

3 Energy density and ΔH per litre are alternative ways to describe the energy content of a fuel. Explain why energy density is used more commonly to compare fuels.

4 Calculate the energy released, in MJ, by 20 kg propane using this thermochemical equation.

$$C_3H_8(l) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) \quad \Delta H = -2220\,kJ\,mol^{-1}$$

5 Calculate how much octane is required to produce 200 MJ of energy.

6 Calculate the volume of ethane that burns to produce 0.2 MJ at STP.

7 Determine if each of the following is an advantage or a disadvantage of the use of bioethanol compared to petrol as an energy source.

- **a** less CO_2 impact overall
- **b** lower energy content ($kJ\,g^{-1}$)
- **c** can be produced from waste products
- **d** renewable resource
- **e** greater amount of CO_2 emitted to travel a set distance

8 Information about two hydrocarbon fuels, propane and octane, is given below.

Characteristic	**Propane (C_3H_8)**	**Octane (C_8H_{18})**
heat of combustion ($kJ\,mol^{-1}$)	–2220	–5450
molar mass ($g\,mol^{-1}$)	44.0	114.0

Calculate the heat of combustion of each fuel in $kJ\,g^{-1}$ and use your answer to state which fuel produces more energy per kilogram.

9 Using the information in Figure 11.4.2 on page 317, determine the amount of energy released when the following amounts of each fuel undergo complete combustion.

- **a** 25 L vegetable oil
- **b** 403 kg of ethanol
- **c** 5 tonnes (5×10^6 g) of brown coal (lignite)

10 Large quantities of black coal are burned in Australia to generate electricity. In the process, significant amounts of the greenhouse gas carbon dioxide are generated. The equation for this combustion reaction is:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Determine the volume of carbon dioxide produced by the combustion of 1 tonne of coal, assuming that the coal is 85 % carbon.

11 Use Table 11.4.5 on page 320 to calculate the extra fuel weight needed to be carried when a car converts from 100% petrol to 100% ethanol. Provide your answer as a percentage.

Analysis

12 Methane and propane are used as alternative transport fuels. The heats of combustion of methane and propane are 890 kJ mol^{-1} and –2220 kJ mol^{-1} respectively.

- **a** Calculate the mass, in kg, of each gas required to release 100 MJ of energy.
- **b** Calculate the mass of carbon dioxide produced by each fuel during the production of 100 MJ of energy.
- **c** Calculate the energy produced per tonne of carbon dioxide released for each fuel.
- **d** Compare methane and propane as transport fuels, based on the values you calculated in parts **a–c**.

13 Evaluate characteristics that make a fuel suitable for:

- **a** heating the home
- **b** vehicle propulsion

14 The ethanol produced by the fermentation of glucose is used as a biochemical fuel or biofuel. Fermentation of glucose produces ethanol and carbon dioxide according to the following equation:

$$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast}} 2C_2H_5OH(aq) + 2CO_2(g)$$

Calculate the mass of ethanol produced by the fermentation of 80.0 g of glucose.

11.5 Environmental impact of the use of fuels

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the greenhouse effect and how excess carbon dioxide from combustion is enhancing it
- analyse temperature changes to the Earth over recent years
- compare carbon dioxide produced per joule of energy for various fuels
- recognise pollutants that are released during combustion of fuels.

The fact that humans are changing their environment seems obvious when watching smoke billowing from a campfire, exhaust from a car or smoke from a factory. However, it was not until the second half of the twentieth century that scientists obtained definitive trends in data relating to changes to the environment. Increasing carbon dioxide levels, increasing average temperatures, loss of Arctic ice, sea level rises, change to ocean pH and increases in extreme weather events have been documented. Initially slight at the start of the **industrial era**, the rate of environmental change has increased dramatically in the past 40 years. Some scientists believe that warming trends are part of the long-term sequence of ice ages and warm periods in Earth's history.

In the past 20–30 years, the Intergovernmental Panel on Climate Change, a United Nations committee, has produced extensive scientific analyses, provided recommendations and sought agreements from all governments to control CO_2 emissions. The most recent was negotiated in Paris in October 2016. Signatories to the agreements have set goals to lower and control greenhouse gas emissions. The majority of worldwide governments have agreed to manage their emissions and set clean energy targets. This gives hope to the idea that, although a 1°C increase in global temperature has been documented since pre-industrial times, it might be possible to limit the increase to 2°C.

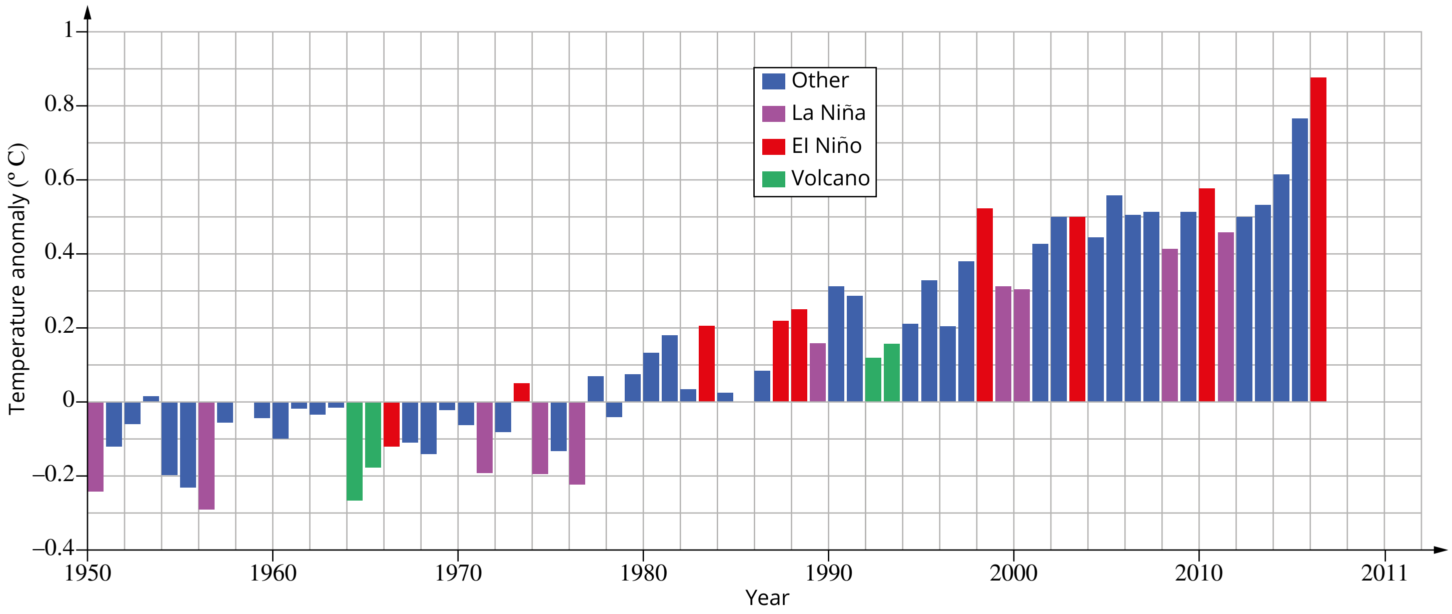
In this module, you will build on ideas from Module 11.4 by continuing to evaluate the effect of the use of fuels on the environment. A particular focus is gases emitted from combustion of fuels and the consequences for our climate.

EMISSIONS FROM FUEL COMBUSTION AND OUR ATMOSPHERE

The atmosphere stores the gaseous elements of the biosphere, protects us from solar radiation and maintains Earth at a habitable temperature. Energy (**infrared radiation**) from the Sun heats the surface of the Earth. The Earth in turn radiates energy back towards space but greenhouse gases in the atmosphere absorb and re-radiate the energy in a process known as the **greenhouse effect**. The higher the concentration of greenhouse gas, the more energy is trapped in the Earth's atmosphere. The greenhouse effect occurs naturally due to the gases present in the atmosphere. However, scientists are concerned that increasing levels of greenhouse gases produced by our use of fossil fuels, predominantly carbon dioxide, are causing global warming and triggering consequential shifts in weather patterns and climate. This is referred to as the **enhanced greenhouse effect**. Methane, water vapour, nitrogen oxides and ozone are also greenhouse gases. Methane is 21 times more effective at trapping heat than carbon dioxide.

The greenhouse effect is caused by heat being trapped in Earth's atmosphere by greenhouse gases, which causes an increase in temperatures at Earth's surface. As the amount of greenhouse gases in Earth's atmosphere increases due to human activities, more heat is trapped, which is predicted to cause global changes in climate.

The graph shown in Figure 11.5.1 supports concerns that the Earth is warming. Many countries are choosing alternatives to fossil fuels to address this issue.



Source: World Meteorological Organisation

FIGURE 11.5.1 Global surface temperature anomalies of the Earth (land and ocean) for 1950–2015. Anomalies are with respect to the 1961–1990 base period. Major tropical volcanic eruptions are associated with cooler global temperatures.

Carbon dioxide

Each fuel discussed in this chapter produces carbon dioxide when burnt. Because large quantities of fuel are burnt every day to meet society's energy needs, the level of carbon dioxide production is high.

Table 11.5.1 compares the theoretical mass of carbon dioxide produced from the complete combustion of 1 gram of each fuel and per unit of energy produced. Refer to Module 11.4 for examples of these calculations.

TABLE 11.5.1 Mass of CO_2 produced from the combustion of 1 gram of fuel

Fuel	Mass of $CO_2(g)$ emitted per gram of fuel	Mass of $CO_2(g)$ per megajoule of energy produced (approx.)
coal	3.7	93
natural gas	2.8	56
LPG	3.0	65
petrol	3.1	73
ethanol	1.9	72

Bioethanol is a renewable fuel derived from plants. Although carbon dioxide is produced when bioethanol burns, carbon dioxide is also absorbed by the plants as they grow. For this reason, the widespread use of bioethanol should lead to a net reduction in the levels of carbon dioxide emitted.

However, bioethanol is not carbon neutral. This is because energy is required, and emissions are produced, in the growing, transport and refining of the fuel. Note also that although a relatively low mass of carbon dioxide is emitted per gram of fuel burnt, because bioethanol produces much less energy than the same mass or volume of petrol, a similar mass of carbon dioxide is emitted to produce the same quantity of energy.

Research is being conducted into ways of reducing carbon dioxide emissions from coal-fired power stations such as:

- storing carbon dioxide deep underground (**geosequestration**)
- removing water from brown coal by heating; the volatile hydrocarbons that are driven off are also burnt to produce electricity
- absorbing carbon dioxide into solvents (chemical capture).

High-energy low-emissions (HELE) power stations can produce up to 30% lower emissions than conventional power stations.

Other emissions

Carbon dioxide and water vapour are not the only products formed when fuels are burnt. Fuels may contain elements such as sulfur, which burns to form sulfur dioxide, or the high temperatures of combustion can lead to reactions with nitrogen in the air. Table 11.5.2 lists the more common pollutants formed.

TABLE 11.5.2 Pollutants formed from fuel combustion

Emission	Formation	Effects
sulfur dioxide (SO_2)	sulfur in fuel reacts with oxygen: $S(s) + O_2(g) \rightarrow SO_2(g)$	• causes respiratory problems • leads to the formation of **acid rain**
nitrogen oxides (NO and NO_2)	nitrogen in fuel, or nitrogen in air, reacts with oxygen due to high temperatures: $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$	• causes respiratory problems • leads to the formation of other pollutants and acid rain
ozone (O_3)	nitrogen oxides react with oxygen at high temperatures	• causes respiratory problems
particulates, such as ash	combustion of impurities in fuel	• irritant • adheres to houses and plants
carbon monoxide	incomplete combustion of fuel	• poisonous gas
organic chemicals such as methanal and ethanal	reactions of other organic chemicals in fuel	• toxic and carcinogenic compounds

The levels of each pollutant mentioned in Table 11.5.2 varies with the composition of the fuel and the efficiency of the combustion reaction. However, it is possible to make general statements.

- Ash is usually more of a problem with coal than other fuels. Ash is produced when coal is burnt without any purification or removal of impurities.
- Sulfur levels are lower in natural gas and LPG than in liquid or solid fossil fuels.
- The molecules in petrol are larger than the molecules in natural gas, LPG and ethanol. As a consequence, the combustion of petrol tends to be less complete. This produces more carbon monoxide and particulates. All new cars now have a catalytic converter attached to the exhaust pipe. Inside, a temperature resistant porous ceramic structure coated with platinum, palladium or rhodium catalyses the oxidation or reduction of some of the harmful pollutants in exhaust gases. Carbon monoxide is oxidised to carbon dioxide. Also, nitrogen oxides are reduced to nitrogen and water.
- Although biofuels are renewable, they can still produce the same pollutants as fossil fuels when burnt.

11.5 Review

SUMMARY

- Scientists and politicians generally agree that our current fuel-using strategies need to change so that fuel reserves will last and that the environment is not further damaged.
- Emissions generated by combustion since our society became industrialised are changing our climate and ecosystems.
- Carbon dioxide absorbs and releases infrared radiation and is responsible for the greenhouse effect—enhancement of this effect has resulted in a 1°C increase in average global temperature (and associated changes) in the past 100 years.
- The fuels you use produce different amounts of carbon dioxide per joule of energy.
- Combustion releases other pollutants such as nitrogen and sulfur compounds into the atmosphere.

KEY QUESTIONS

Retrieval

1 State the type of radiation that is absorbed by greenhouse gases.

2 Identify three greenhouse gases and two other atmospheric gases that are not greenhouse gases.

Comprehension

3 The emissions of carbon dioxide per gram from the combustion of bioethanol are less than that from octane (petrol). However, a car using bioethanol produces more carbon dioxide when driving the same distance as a car using octane. Explain the reasons for this.

4 Explain the various factors involved in the greenhouse effect shown in the diagram below.

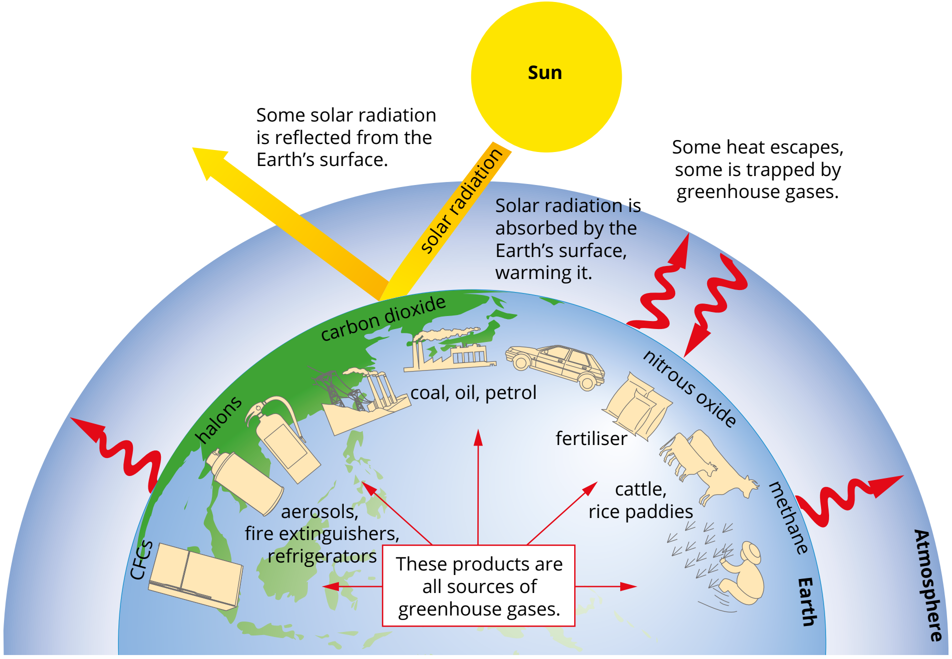
5 The 2016 Paris Agreement was an international response to global warming in which all nations were asked to commit to keeping the global average temperature rise to below 2°C, through reductions in greenhouse gas emissions.

a Discuss impacts in Australia of committing to reductions in greenhouse gas emissions.

b Discuss the role that biofuels could play in helping Australia meet its target for reducing greenhouse gas emissions.

6 Water is a greenhouse gas, and as the Earth becomes warmer, water from oceans will vaporise and enter the atmosphere. This will then increase the warming effect in a cycle known as positive feedback.

Draw a flow diagram to illustrate these ideas.



Chapter review

WS 1.3.10 WS 1.3.11

11

KEY TERMS

acid rain
alkane
biodiesel
bioethanol
biofuel
biogas
carbon neutral
chemical energy
climate change
combustion
complete combustion
crude oil
E10 petrol blend
energy density
enhanced greenhouse effect
enthalpy
ester
fermentation
flashpoint
fossil fuels
fractional distillation
geosequestration
green chemistry principles
greenhouse effect
greenhouse gases
heat of combustion
hydrocarbons
incomplete combustion
industrial era
infrared radiation
liquefied petroleum gas (LPG)
methane
natural gas
particulate matter
petrodiesel
photosynthesis
renewable energy
sustainable
thermochemical equation
transesterification
viscosity
volatile

KEY QUESTIONS

Retrieval

1 Identify which of the following statements about combustion reactions is correct.
  A Combustion reactions are usually exothermic, meaning energy is absorbed during the reaction.
  B Combustion reactions are always exothermic, meaning energy is absorbed during the reaction.
  C Combustion reactions are usually endothermic, meaning energy is absorbed by the system.
  D Combustion reactions are always exothermic, meaning energy is released during the reaction.

2 Select which of the following statements describes a renewable energy source.
  A A renewable energy source can be produced at a slower rate than the rate at which it is used by society.
  B A renewable energy source can be produced at a faster rate than the rate at which it is used by society.
  C A renewable energy source can be produced at the same rate as the rate at which it is used by society.
  D The rate at which a renewable energy source can be produced is unrelated to the rate at which it is used by society.

3 Select the correct statement about petrodiesel.
  A Petrodiesel is a renewable fuel as it is formed from plants.
  B Petrodiesel is purified to ensure all molecules include the same number of carbon atoms.
  C Petrodiesel is a mixture of substances including LPG and petrol.
  D Petrodiesel contains 75% alkanes with 10–15 carbon atoms.

4 Octane is the major component of petrol. Write a balanced equation for the complete combustion of octane (C_8H_{18}).

5 Convert the following units of energy to the unit given.
  a 2200 J to kJ
  b 0.774 kJ to J
  c 4 420 000 J to MJ
  d 0.056 MJ to kJ

6 List the following fuels in order of increasing energy output, as measured in kJ g^{-1}?
  brown coal, petrol, ethanol and diesel

Comprehension

7 A power station converts 85 MJ of input energy into 35 MJ of useable output energy. Determine its efficiency percentage.

8 **a** Explain the term ‘E10 petrol’.
  b Describe how the introduction of E10 petrol helps with the potential shortage of crude oil.

9 Ethanol is produced industrially by reacting ethene with water using a phosphoric acid catalyst at 300°C in the following reaction:

$$C_2H_4(g) + H_2O(g) \rightarrow CH_3CH_2OH(g)$$

  a Explain whether ethanol produced by this method is a biofuel or fossil fuel.
  b Describe how bioethanol is produced.

10 Butane is used as the fuel in some portable camping stoves. It is a liquid when stored under pressure in a butane cylinder, but vaporises when the valve is opened. Combustion of butane is represented by the equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l) \quad \Delta H = -5772 \text{ kJ}$$

- **a** Determine the energy produced when 10.0 g of butane burns completely.
- **b** Determine the energy produced when 0.1 L of butane, measured at STP, burns completely.
- **c** Calculate the volume of butane, measured at STP, that must be burnt to produce 1.0 kJ of energy.

11 Compare the chemical equations for complete and incomplete combustion of ethane. Explain the nature of the products of each in terms of the amount of oxygen present as a reactant in each case.

Analysis

12 Refer to Figure 11.1.1 on page 300 to answer the following questions.

- **a** Identify the time period during which national energy consumption increased from 2900 PJ to 5800 PJ.
- **b** Consider the change to our population and evaluate the two sets of data.

13 Analyse the importance of fossil fuels in your everyday life. Make a list of activities and items used during a single day that use energy and products of fossil fuels. Include transport, use of power and use of plastics.

14 The table below shows data for composition by weight of some hydrocarbons in petrol. Determine why the composition of petrol varies. Use values from the data as examples.

Hydrocarbon	Average (%)	Range (%)
paraffins (alkanes)	30	15–60
naphthanes (cycloalkanes)	49	30–60
aromatics	15	3–30
asphaltics	6	remainder

15 Methane and methanol both burn in air according to the thermochemical equations:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890 \text{ kJ}$$

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l) \quad \Delta H = -1450 \text{ kJ}$$

- **a** If 1 mole of methane and 1 mole of methanol are completely burnt in separate experiments, determine which experiment will release the most energy.
- **b** If each of the above reactions is used to produce 2000 kJ of energy, determine which fuel will release the most carbon dioxide.

16 Octane (C_8H_{18}) is a major component of petrol. It burns in oxygen to produce carbon dioxide and water. Energy is released during this reaction. The equation for this reaction is:

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

- **a** Calculate the mass of oxygen required to react with 200 g of octane.
- **b** Calculate the mass of carbon dioxide produced in part **a**.

Knowledge utilisation

17 The world has become very dependent on the products of the petrochemical industry, but the raw materials of coal, crude oil and natural gas are likely to be virtually exhausted by 2100. Assuming that current production remains unchanged and no alternative sources are available, evaluate the impact of the lack of these raw materials on our lifestyle.

18 Evaluate the two biodiesel sources using the information in Module 11.3 and the table of data below as a starting point. Appraise the:

- need for prime agricultural land
- removal of native ecosystems
- oil yield
- current understanding of the technology

Oil source	Yield ($L\,ha^{-1}$)
algae	>47 000
palm oil	5900
canola	1200
sunflower	950
soy	450

19 Propose three ways in which your family could decrease its electricity use to contribute to the minimisation of carbon dioxide release into the atmosphere.

UNIT 1 • REVIEW

REVIEW QUESTIONS

TR 1.1 TR 1.2 TR 1.3

Chemical fundamentals—structure, properties and reactions

Topic 1: Properties and structure of atoms

Multiple-choice questions

1 Compare the electronegativity and the first ionisation energy of caesium, Cs, and barium, Ba.

A Cs is less electronegative and has a lower first ionisation energy.

B Cs is more electronegative and has a higher first ionisation energy.

C Cs is less electronegative and has a higher first ionisation energy.

D Cs is more electronegative and has a lower first ionisation energy.

2 Select the case in which the radius of the first-named species is smaller than the radius of the second-named species.

A chloride ion (Cl^-): chlorine atom (Cl)

B iron(II) ion (Fe^{2+}): iron(III) ion (Fe^{3+})

C magnesium atom (Mg): magnesium ion (Mg^{2+})

D potassium ion (K^+): potassium atom (K)

3 Determine the effective nuclear charge of an atom that has the electronic structure $1s^22s^2p^63s^23p^4$.

A 2

B 4

C 6

D 16

4 Consider the following oxides:

Al_2O_3, Cl_2O, Na_2O

Identify the sequence that correctly lists these oxides in order, from left to right, from basic to amphoteric to acidic.

A Al_2O_3, Cl_2O, Na_2O

B Cl_2O, Na_2O, Al_2O_3

C Na_2O, Al_2O_3, Cl_2O

D Cl_2O, Al_2O_3, Na_2O

5 Consider the following electrostatic forces:

I attraction between oppositely charged particles

II repulsion between positively charged particles

III repulsion between negatively charged particles

The electrostatic forces operating in an atom are:

A I only

B I and II only

C I and III only

D I, II and III

6 The atomic number, mass number and electron configuration of four particles, W, X, Y and Z, are given below.

	Atomic number	Mass number	Electron configuration
W	17	37	$1s^22s^22p^63s^23p^6$
X	19	39	$1s^22s^22p^63s^23p^6$
Y	17	37	$1s^22s^22p^63s^23p^63p^5$
Z	19	40	$1s^22s^22p^63s^23p^64s^1$

Identify the correct statement.

A W is a noble gas.

B X is a positively charged ion.

C Y is in group 5 of the periodic table.

D Z is a negatively charged ion.

7 Element X is in group 4 of the periodic table. Element Y is in group 16. Deduce the most likely formula for a compound between X and Y.

A X_4Y_{16}

B $X_{16}Y_4$

C XY_4

D XY_2

8 The formula for hydrogen cyanide is HCN. Deduce the number of electrons shared between the carbon and nitrogen atom.

A 2

B 4

C 6

D 8

9 Select the pair of elements that will form an ionic compound.

A sulfur with oxygen

B copper with tin

C tin with chlorine

D carbon and chlorine

10 Bromine has two isotopes, ^{79}Br and ^{81}Br. It has a relative atomic mass of 79.9. The percentage of ^{81}Br in a sample of naturally occurring bromine is closest to:

A 45

B 55

C 50

D 35

11 The atomic number, mass number and electron configuration of four particles, W, X, Y and Z, are given below.

	Atomic number	Mass number	Electron configuration
W	17	37	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}$
X	19	39	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}$
Y	20	40	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}4s^{2}$
Z	19	40	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}4s^{1}$

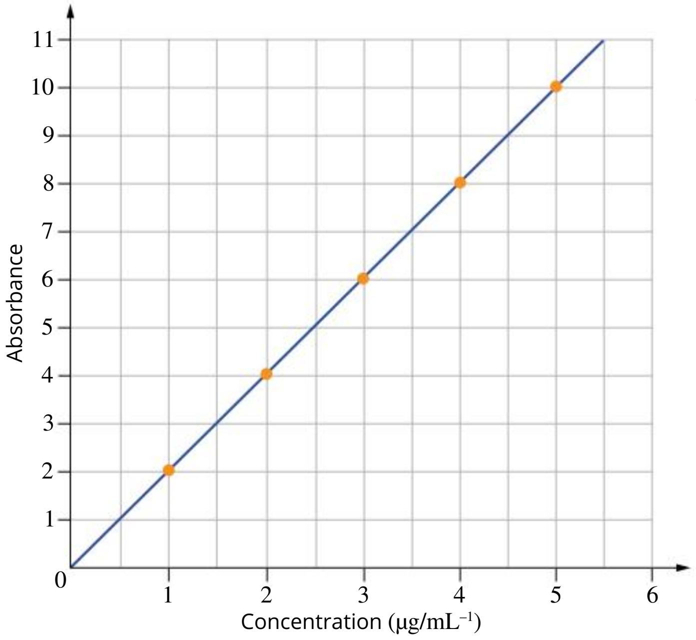
Identify the particles that are isotopes of the same element.

A W and X only
B X and Z only
C Y and Z only
D W, X and Y only

12 Copper has two isotopes, ^{63}Cu and ^{65}Cu. It has a relative atomic mass of 63.55. Any one particular copper atom would:

A contain either 63 or 65 protons
B contain either 34 or 36 neutrons
C have a mass number of 63.55
D have an atomic number of 29 or 31

13 Cadmium is a toxic heavy metal which can cause kidney disease. The concentration of cadmium ions in river water was determined using atomic absorption spectrometry (AAS). The absorbencies of several solutions of known concentration of cadmium ions were measured and the results plotted as a calibration graph.



A sample of river water had an absorbance of 5.0. The cadmium ion concentration, in μg mL^{-1}, in the sample of water is:

A 2.0
B 2.5
C 5.0
D 10.0

14 When a salt was heated in the flame of a Bunsen burner, the flame colour changed to an intense yellow. Identify the salt.

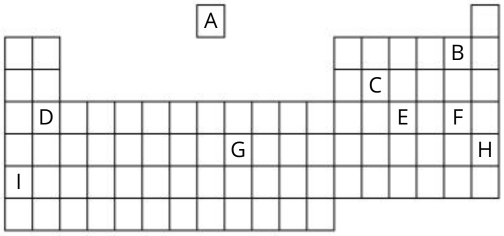
A barium chloride
B calcium chloride
C potassium chloride
D sodium chloride

Short-answer questions

1 **a** With the aid of a periodic table, identify the correct chemical symbol for each of the following.
  i the element that is in group 2 and period 4
  ii a noble gas with exactly three occupied electron shells
  iii an element from group 14 that is a non-metal
  iv an element that has exactly three occupied electron shells and is in the *s*-block
  v the element in period 2 that has the largest atomic radius
  vi the element in group 15 that has the highest first ionisation energy
  vii the element in period 2 with the highest electronegativity

b Describe the trend in chemical reactivity of elements going down group 1 of the periodic table and give a brief explanation for this trend.

2 Below is an outline of the periodic table, with a number of elements identified using letters that are not their normal symbols. Use these letters when answering the following questions.



a Select the element that has the highest electronegativity.

b Identify the elements that have an electron configuration ending in $s^{2}p^{5}$.

c Identify the element that does not usually react with other elements.

d Select the element that has an electron shell configuration of 2,8,18,5.

3 Complete the table below by placing a tick (√) in the box that correctly describes the trend in the property listed as you move down a group and across a period from left to right.

Property	Trend in property			
	down a group		across a period (from left to right)	
	Increase	Decrease	Increase	Decrease
atomic radius				
ionisation energy				
electronegativity				
metallic character				
non-metallic character				
ease of formation of positive ions				
ease of formation of negative ions				
acidic nature of oxides				

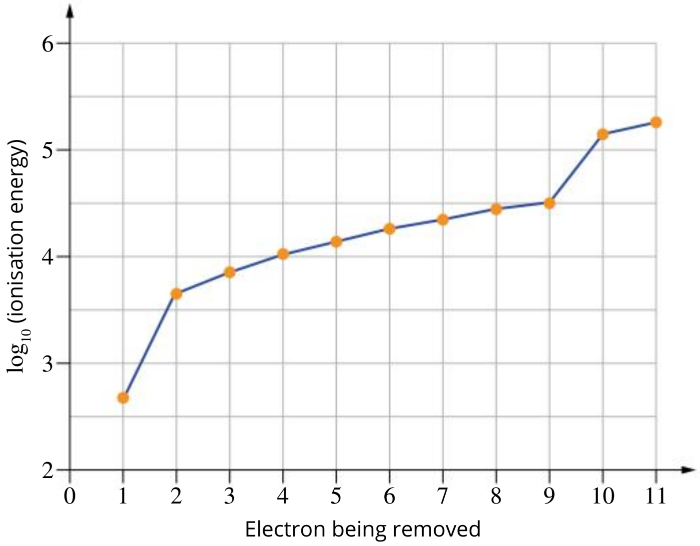
4 The ionic radii of some ions are provided in the following table.

Ion	Ionic radius (pm)	Number of protons	Electron configuration
N^{3-}	146		
O^{2-}	140		
F^-	133		
Na^+	102		
Mg^{2+}	72		
Al^{3+}	53		

a Complete the table by indicating the number of protons in the nucleus of each ion and the electron configuration of each ion.

b Suggest a reason why the ionic radii decrease down this set of ions.

5 The graph below shows the successive ionisation energies for sodium.



Use the information provided in this graph to explain how the energies for the successive ionisation of sodium are related to its electronic structure.

6 **a** Using an atom of sodium as an example, clearly explain the difference between an electron shell, a subshell and an orbital.

b State the Pauli exclusion principle.

c Consider the following electron configuration: $1s^22s^22p^63s^23p^63d^1$.

i Deduce if this electron configuration represents a ground state or excited atom.

ii Give a reason for your answer.

7 This question concerns two elements, identified only as X and Y. Information concerning ions of these two elements is shown in the table below.

Element	Mass number	Atomic number	Neutron number	Electron number
X	24	12	i	10
Y	ii	15	16	iii

a Supply the following information to complete the table.

i Determine how many neutrons are in an atom of X.

ii Determine the mass number of an atom of Y.

iii Determine how many electrons the Y^{3-} ion contains.

iv Give the symbol for the ion formed by X.

b **i** Deduce the group in the periodic table to which element X belongs.

ii Deduce the period in the periodic table to which element Y belongs.

c Explain why an atom of X has a larger radius than an atom of Y.

8 **a** Write the formula of each of these compounds.

- **i** potassium sulfate
- **ii** aluminium oxide
- **iii** sodium nitrate
- **iv** iron(III) sulfide

b Write the name of each of the compounds listed below.

- **i** $Ca(NO_3)_2$
- **ii** $CuCO_3$
- **iii** $Mg(OH)_2$
- **iv** Cr_2O_3

c The formula of the ionic compound calcium phosphide is Ca_3P_2.

- **i** Write the formula for the phosphide ion.
- **ii** Provide the ground state electron configuration of the phosphide ion.

9 Discuss this statement in terms of ionic and covalent bonding, using potassium chloride and chlorine as examples: 'All chemical bonding is electrostatic in nature'.

10 Draw the Lewis dot structure in the space below and state the number of bonding and non-bonding pairs of electrons for the following species.

a CO_2 number of bonding pairs number of non-bonding pairs

b NH_4^+ number of bonding pairs number of non-bonding pairs

c HF number of bonding pairs number of non-bonding pairs

11 Iridium has two naturally occurring isotopes. Their relative abundances and masses are shown in the table below.

Isotope	Relative isotopic mass	Percentage abundance
^{191}Ir	190.97	37.30%
^{193}Ir	192.97	62.70%

a Recall the name of the instrument that is commonly used to experimentally obtain the abundances and relative isotopic masses data.

b Using the information in the table, determine the relative atomic mass of iridium. Show your working and give your answer to the appropriate number of significant figures.

12 Lithium has two isotopes. Their relative isotopic masses are 6.02 and 7.02. The relative atomic mass of lithium is 6.94. Calculate the relative abundance of the lighter isotope.

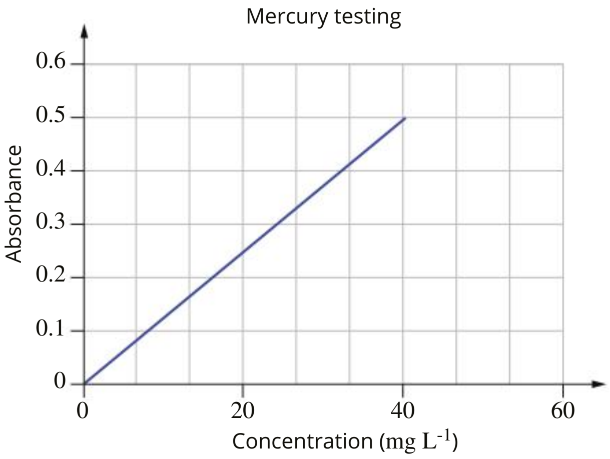
13 **a** Describe how Bohr used evidence from spectral analysis to propose his model of the atom.

b State the name given to the set of lines found in the visible region of the hydrogen emission spectrum. Describe the electron transitions that generate these lines.

c Describe one way in which the Schrödinger model of the atom differs from that of Bohr.

14 Deep sea fish can build up high levels of mercury from contaminated water. A 3.0 g sample of fish is ground up and mixed with water. The mixture is made up to 100 mL with water. A sample of the solution is analysed by atomic absorption spectroscopy, giving an absorbance reading of 0.23.

Use the following calibration curve to determine the quantities below.



a percentage mass of mercury in the fish

b concentration in ppm of mercury in the fish

15 Atomic absorption spectroscopy was used to determine the concentration of sodium ions in ground water.

The absorbance readings of a series of standard solutions are shown in the table below.

Concentration of sodium (mg L^{-1})	Absorbance
2.0	0.14
4.0	0.28
6.0	0.42
8.0	0.56
10.0	0.70

a Identify what evidence there is in the data to indicate that the calibration graph is linear.

b Plot the calibration graph.

c Determine the concentration of sodium in a sample that had an absorbance of 0.38.

16 The atomic number of four elements I, II, III and IV and their atomic radii are provided in the following table. Refer to the information provided in this table when answering parts **a** and **b**.

element	I	II	III	IV
atomic number	11	19	20	35
atomic radius	186	227	147	114

a Complete the following table by writing the electronic structure of these elements in the long and short form.

Element	Long form	Short form
I		
II		
III		
IV		

b Describe, using elements I, II, III and IV as examples, how the structure of the periodic table is determined by the atomic number and the properties of the elements. Identify the group and period to which each of the four elements belong.

c Explain why the periodic table consists of four blocks and why the number of columns within each block differs.

17 Describe and explain the trends in atomic radii in the periodic table. Your answer should include the following points.

- Explain the meaning of the term 'effective nuclear charge' and outline how this may be calculated.
- Explain the role of effective nuclear charge in determining the radius of an atom.
- Identify and explain other factors that determine the radius of atoms.
- Illustrate your answer with appropriate examples using the elements included in the table in Question **16**.

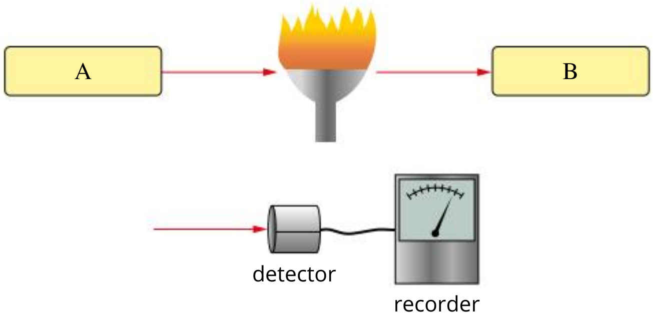
18 **a** Use orbital diagrams to illustrate the application of the Aufbau principle, Hund's rule and the Pauli exclusion principle in the determination of the electron configuration of sodium, carbon and iron.

b Identify two elements that have an electron configuration different from that predicted by the Aufbau principle. Provide the actual and predicted electron configurations for each element.

19 Mass spectrometry and atomic absorption spectroscopy are two analytical techniques used by chemists. Compare these two techniques by:

- describing the information provided by each technique
- outlining the basic principles upon which each technique is based.

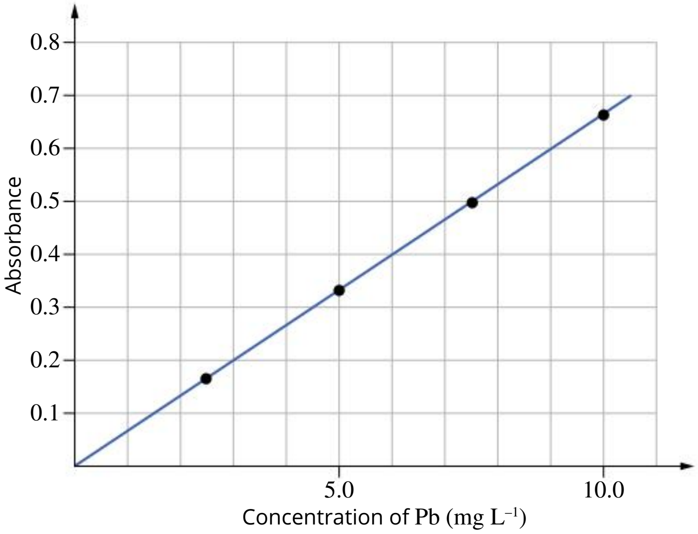
20 The diagram below shows a simplified flow chart for the operation of an atomic absorption spectrometer.



a **i** How is the part labelled A chosen for a particular analysis?

ii Identify the function of part B.

b The lead level in a sample of contaminated soil was investigated using atomic absorption spectrometry. A 3.0 g sample was dissolved in acid and then diluted to a total volume of 25.0 mL. The absorbance of this solution, determined in a spectrometer set at an appropriate wavelength, was found to be 0.40. Several standard solutions of lead were tested under the same conditions and the calibration curve shown below was generated.



- **i** Determine the mass of lead present in the 3.0 g sample.
- **ii** Determine the concentration of lead in the soil as a % m/m.
- **iii** Explain why it was necessary to construct a calibration curve in this investigation.

c Briefly discuss the difference between the emission and absorption spectrum of hydrogen. Describe, in terms of electron transitions, how each spectrum is generated.

Topic 2: Properties and structure of materials

Multiple-choice questions

1 Graphenes and fullerenes are classified as carbon nanomaterials. Select the measurements that are equivalent to one nanometre.

- **I** 10^{-6} mm
- **II** 10^{-9} mm
- **III** 10^{-12} mm
- **IV** 10^{-9} m

A I only
B I and IV only
C II and IV only
D III and IV only

2 Zinc is an element. Determine which statement is true.

A Zinc has no isotopes.
B All zinc atoms are identical.
C Zinc atoms always contain the same number of protons.
D Zinc atoms contain equal numbers of protons and neutrons.

3 Select the statement that correctly lists the properties of oxygen gas.

A less dense than air and highly soluble in water, where it forms a neutral solution
B less dense than air and very slightly soluble in water, where it forms an acidic solution
C denser than air and very slightly soluble in water, where it forms a neutral solution
D denser than air and highly soluble in water, where it forms a neutral solution

4 Select the statement about solids liquids and gases that is incorrect.

A Liquids and gases are both fluids and they can be poured.
B Solids and liquids have a definite volume but gases have a variable volume.
C Solids are incompressible but liquids and gases are highly compressible.
D Solids have a definite shape, but liquids and gases have variable shapes.

5 Select the correct statement about the boiling points of pure water and salt water.

A The boiling points of pure water and salt water are the same.
B The boiling point of pure water is less than the boiling point of salt water.
C The boiling points of pure water and salt water depend upon the volume of the liquids.
D The boiling point of pure water remains constant, while the boiling point of salt water decreases as more salt is added.

6 You have been provided with two solutions. Solution A contains 1 g of NaCl dissolved in 50 mL of water. Solution B contains 5 g of NaCl dissolved in 100 mL of water. Select the statement that correctly describes the difference between the properties of solutions A and B.

A Solution A has a higher boiling point than solution B.
B Solution A has a lower electrical conductivity than solution B.
C Solution A has a higher density than solution B.
D Solution A has a lower freezing point than solution B.

7 Identify the statement that does NOT apply to ionic bonding.

- **A** When molten, ionic compounds are conductors of electricity.
- **B** An ionic lattice contains both cations and anions in fixed positions.
- **C** Ionic bonding involves the sharing of electrons between two different atoms.
- **D** Compounds held together by ionic bonding generally have high melting temperatures.

8 It would be possible to tell whether the particles in a compound were held together by ionic or covalent bonds by measuring which characteristic of the compound?

- **A** melting temperature
- **B** hardness and brittleness
- **C** electrical conductivity in the solid state
- **D** electrical conductivity in the liquid (molten) state

9 Identify the member of the alkane family.

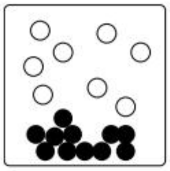
- **A** C_2H_4
- **B** C_3H_8
- **C** C_6H_6
- **D** C_8H_{16}

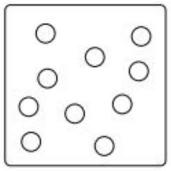
10 Predict which one of the following ionic compounds will have the highest melting point.

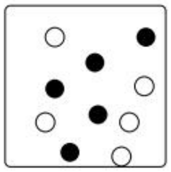
- **A** MgO
- **B** $MgCl_2$
- **C** Na_2O
- **D** NaCl

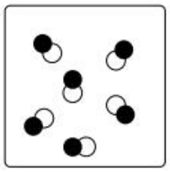
Short-answer questions

1 Two different elements are represented in the following diagram as black and white circles.


A


B


C


D

Match each option below with the appropriate diagram.

- **a** a pure substance
- **b** an element
- **c** a compound
- **d** a heterogeneous mixture
- **e** a homogeneous mixture

2 **a** Classify each of the following substances by placing a tick in the appropriate place in the table below.

Substance	Element	Compound	Homogeneous mixture	Heterogeneous mixture
salt water				
silver				
concrete				
sugar				
oil and water				
air				

b Deduce which of the substances listed in the above table are pure substances.

3 Ionic compounds have a variety of uses in everyday life.

- **a** Give the formula and a use of one ionic compound.
- **b** Use the ionic bonding model to explain the following properties of ionic compounds.
  - **i** Solid ionic compounds cannot conduct electricity.
  - **ii** Ionic compounds are generally brittle.
  - **iii** Ionic compounds have high melting temperatures.

4 Magnesium is a commonly used structural metal.

- **a** The radius of a magnesium atom is 160 pm (1 pm = 10^{-12} m). Convert the radius of the magnesium atom to nanometres.
- **b** **i** Describe the model commonly used to describe the structure of metals such as magnesium and the nature of the bonding between its particles. You may include a labelled diagram in your answer.
  - **ii** Use the metallic bonding model to explain why magnesium is a good conductor of electricity.
  - **iii** The metallic bonding model is useful to explain many of the properties of metals. However, like most models, it has some limitations. Give one example of the limitations of the metallic bonding model.
- **c** Magnesium is most commonly used as an alloy for structural purposes. Define an alloy. Give one example of how alloying affects the properties of a metal.

5 Provide concise explanations for each of the following observations.

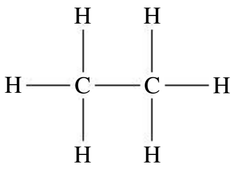
- **a** The melting temperature of ice (solid H_2O) is 0°C but a temperature of over 1000°C is needed to decompose water molecules to hydrogen and oxygen gases.
- **b** Most substances are denser in their solid state than in their liquid state, yet solid water (ice) floats on liquid water.

6 Explain the difference between the structure of diamond and that of a fullerene.

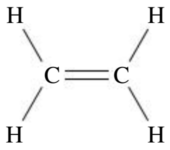
7 Explain the following observations in terms of structure and bonding.

- **a** Nickel chloride only conducts electricity when molten or in aqueous solution, not when solid.
- **b** Diamond is widely used as an abrasive on the ends of drill bits.

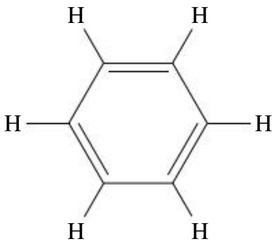
8 Complete the table below by matching the names of the three hydrocarbons with their bond energies, bond length and reactivity with bromine.



ethane



ethene



benzene

Hydrocarbon	C–C bond energy (kJ mol^{-1})	Bond length (pm)	Reaction with Br_2
	356	154	no reaction
	636	134	reaction
	518	140	no reaction

9 The physical properties of the components in a mixture can be used to separate them. Identify a separation technique that will separate the components of the following mixtures. Describe the properties of the mixture's components that are used to separate them.

- **a** iron filings from a mixture of iron filings and sand
- **b** salt from saltwater
- **c** water from a mixture of sand and water
- **d** alcohol from a mixture of alcohol and water
- **e** oil from a mixture of oil and water

10 The properties of some substances are described below.

- **a** Substance W is hard, brittle and does not conduct electricity and sublimes at approximately 3500°C.
- **b** Substance X melts at 645°C and does not conduct electricity in the solid state, but conducts electricity when molten.
- **c** Substance Y melts at 842°C, and conducts in both the solid and liquid states.
- **d** Substance Z is a diatomic gas at room temperature and does not conduct electricity in either the solid or liquid state.

Identify and describe the type of bonding in each substance and explain the properties listed in terms of the type of bonding.

11 Duralumin is a light but strong metal used in the manufacture of aircraft parts and racing bikes. It is an alloy of aluminium, containing around 4% copper and 1% magnesium and manganese.

- **a** Describe the structure and bonding in Duralumin.
- **b** Explain why Duralumin is:
  - **i** malleable
  - **ii** able to conduct electricity
- **c** Duralumin is harder than aluminium. Account for the difference in terms of structure and bonding.

12 Diamond and graphite are allotropes of carbon. Although they are similar in some respects, they are very different in their structure and uses.

- **a** Give a brief definition of the term 'allotrope'.
- **b** Describe the similarity in the bonding of these two materials.
- **c** Compare the structures of diamond and graphite.
- **d** Graphite is an excellent conductor of electricity, while diamond is unable to conduct electricity. In terms of the structure of their lattices, explain this difference.
- **e** Graphite is used as lead in pencils. With reference to the bonding in the lattice, explain why graphite can be used in this way.
- **f** Describe the structure of graphene.

Topic 3: Chemical reactions—reactants, products and energy change

Multiple-choice questions

1 Identify the set of coefficients that would balance the following equation.

$$wC_2H_5OH(l) + xO_2(g) \rightarrow yCO_2(g) + yH_2O(l)$$

w, x, y, z

- **A** 1, 3, 2, 3
- **B** 2, 7, 4, 6
- **C** 2, 6, 4, 5
- **D** 1, 6, 2, 3

2 The reaction of sodium sulfite with hydrochloric acid may be represented by the equation:

$Na_2SO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g)$

On the basis of the information in the equation, identify the correct statement.

A When 1 mol of Na_2SO_3 is treated with acid, 1 mol of SO_2 will always be produced.
B Na_2SO_3 and H^+ must always be present in the mole ratio of 1:2 for a reaction to take place.
C The ratio by mass of Na_2SO_3 to H^+ in a reaction mixture containing the correct stoichiometric proportions is 1:2.
D 0.4 mol of H_2O will be produced when 0.8 mol of H^+ reacts completely with excess Na_2SO_3.

3 Identify the statement that best defines the empirical formula of a molecular compound.

A actual number and type of each element present
B simplest whole number mole ratio of each ion present
C simplest whole number mole ratio of each element present
D simplest whole number mass ratio of each element present

4 Determine which one of the following quantities contains the same number of ions as there are in 1 mol of NaOH.

A 1 mol of Na_2S
B 0.5 mol of K_3PO_4
C 1 mol of $MgCl_2$
D 1.5 mol of MgO

5 A particular brand of plant fertiliser contains 58.5% urea ($CO(NH_2)_2$) as the only nitrogen-containing compound. Determine the mass of nitrogen, in g, in 125 g of this fertiliser.

A 29.2
B 40.0
C 50.0
D 73.1

6 Determine the mass, in g, of calcium chloride that contains 6.02×10^{23} chloride ions.

A 35.5
B 55.6
C 75.6
D 111

7 A 54 g sample of hydrogen cyanide, HCN, combined with 80 g of oxygen to produce 18 g of water, 28 g of nitrogen and carbon dioxide, but nothing else. Calculate the mass of carbon dioxide produced.

A 44 g
B 88 g
C 134 g
D 286 g

8 In an experiment to determine the empirical formula of a chloride of a metal *X*, a student reduced 0.277 g of the chloride of *X* to the metal *X* itself. She obtained 0.207 g of *X*.

(relative atomic masses: $X = 207$; $Cl = 35$)

From these results, the formula of the compound is shown to be:

A XCl
B XCl_2
C XCl_3
D XCl_4

9 The combustion of 1 mol of ethanol, CH_3CH_2OH, in excess oxygen generates 1364 kJ of heat energy. Select the correct thermochemical equation for the reaction.

A $CH_3CH_2OH + 2O_2 \rightarrow 2CO_2 + 3H_2O \quad \Delta H = +1364\ kJ\ mol^{-1}$
B $CH_3CH_2OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l) \quad \Delta H = -1364\ kJ\ mol^{-1}$
C $CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1364\ kJ\ mol^{-1}$
D $CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = +1364\ kJ\ mol^{-1}$

10 The thermochemical equation for the combustion of methanol in excess oxygen is:

$$CH_3OH(l) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -726\ kJ\ mol^{-1}.$$

Calculate the mass of methanol that must be consumed to produce 1.00 MJ of heat energy.

A 0.726 g
B 22.9 g
C 23.2 g
D 44.1 g

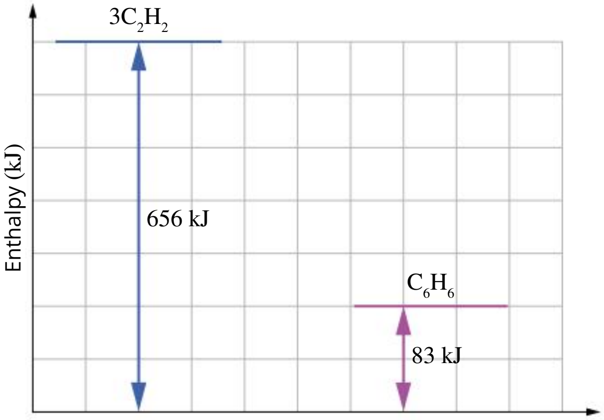
11 Select the statement that is true for an exothermic reaction.

I The system absorbs heat.
II The enthalpy change is negative.
III The reactants are more stable than the products.

A II only
B I and II only
C II and III only
D I, II and III

12 Use the information in the diagram to determine Δ*H* for the reaction below from the options.

$$3C_2H_2(g) \rightarrow C_6H_6(g)$$



A +739 kJ
B +573 kJ
C –573 kJ
D –656 kJ

13 For the reaction $2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$; Δ*H* = +92 kJ, select the correct statement.

A When 4 mol of ammonia decomposes, 92 kJ of heat will be evolved.
B When 1 mol of nitrogen reacts with 3 moles of hydrogen, 92 kJ of heat will be absorbed by the system.
C When 1 mol of ammonia reacts, 46 kJ of heat will be absorbed from the surroundings.
D When 2 mol of ammonia decomposes, the temperature will decrease by 92°C.

14 Identify the list that only contains compounds likely to be found in crude oil.

A C_2H_4, C_3H_7OH, C_8H_{16}, C_8H_{18}
B C_2H_6, C_4H_{10}, C_8H_{18}, $C_{16}H_{34}$
C C_2H_4, C_2H_6, C_2H_5OH, CH_3COOH
D C_2H_4, C_2H_5COOH, C_6H_{12}, C_8H_{16}

15 Select the statement about petrodiesel and biodiesel that is false.

A The energy content of 1 kg of biodiesel is less than the energy content of 1 kg of petrodiesel.
B Biodiesel is obtained from renewable sources. Petrodiesel is derived from non-renewable sources.
C The combustion of biodiesel does not produce greenhouse gases, whereas the combustion of petrodiesel does.
D A typical formula of a molecule found in biodiesel is $C_{16}H_{34}O_2$. A typical formula of a molecule found in petrodiesel is $C_{12}H_{23}$.

Short-answer questions

1 Write a balanced molecular equation for each of the following reactions.

a aluminium is heated with steam to give hydrogen gas and solid aluminium oxide
b chlorine gas is bubbled through a solution of sodium iodide to form solid iodine and a solution of sodium chloride
c solid manganese(IV) oxide (MnO_2) is added to dilute hydrochloric acid to produce water, chlorine gas and a solution of manganese(II) chloride
d solutions of calcium hydroxide and phosphoric acid are mixed to give a precipitate of calcium phosphate and water
e copper(II) oxide reacts with ammonia gas to produce copper, nitrogen gas and water

2 A sample of aluminium nitrate ($Al(NO_3)_3$) has a mass of 30.5 g.

a Calculate the amount, in mol, of aluminium nitrate in the sample.
b Calculate the amount, in mol, of nitrogen atoms present.
c Calculate the total number of atoms present in the sample.
d Determine the percentage by mass of oxygen in aluminium nitrate.

3 Copper forms two oxides, copper(I) oxide (Cu_2O) and copper(II) oxide (CuO). The following experiment was carried out to determine whether a particular oxide of copper was Cu_2O or CuO.

A sample of the oxide was dissolved in sulfuric acid. An excess of zinc powder was added to precipitate the copper as copper metal. The copper metal was collected, washed, thoroughly dried and then weighed. The results are given below:

$$m(\text{oxide}) = 2.127\text{ g}$$
$$m(\text{dried copper}) = 1.704\text{ g}$$

a Determine the empirical formula of this oxide of copper.
b If the sample analysed was a mixture of the two oxides instead of one of the oxides only, determine if the mass of copper obtained would be greater than or less than 1.704 g. Explain your answer.

4 Marble chips (calcium carbonate) dissolve readily in hydrochloric acid to produce calcium chloride, carbon dioxide and water. The equation for this reaction is:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

a Calculate the expected mass of carbon dioxide produced by the complete reaction of 2.75 g of marble chips.
b Calculate the percentage yield if the mass of carbon dioxide produced was 0.89 g.
c Suggest one reason why the mass of carbon dioxide obtained was less than the value calculated in part **a**.

5 The specific heat capacity of some common substances are listed in the table below.

Substance	Specific heat capacity (J g^{-1}°C^{-1})
water	4.18
ethanol	2.4
copper	0.39

- **a** Calculate the amount of heat energy, in J, required to increase the temperature of 250.0 g of water from 18°C to 100°C.
- **b** Calculate the amount of energy required to raise the temperature of an equal mass of copper by the same amount.
- **c** If the same amount of energy as in part **b** is used to heat 250.0 g ethanol, determine the temperature rise of the ethanol.
- **d** Explain why water has a high heat capacity.
- **e** Give one use for water based on its high heat capacity.

6 The specific heat capacity of water is 4.18 J g^{-1}°C^{-1} and the specific heat capacity of liquid hexane (C_6H_{14}) is 2.29 J g^{-1}°C^{-1}.

- **a** State what is measured by the specific heat capacity of a substance.
- **b** A 100 g sample of water and 100 g of hexane are both at a temperature of 20.0°C. The same amount of heat is applied to each liquid. The temperature of the water rises to 32.0°C. Calculate the temperature the hexane rises to after heating.

7 Use the average bond energies provided to calculate the enthalpy change for the following reaction. Indicate if the reaction is exothermic or endothermic. Explain your reasoning.

$$N_2(g) + 3Cl_2(g) \rightarrow 2NCl_3(g)$$

Bond	Average bond energy (kJ mol^{-1})
N ≡ N	941
Cl–Cl	242
N–Cl	200

8 Butane undergoes combustion with oxygen according to the equation

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$
$$\Delta H = -5774\ \text{kJ}$$

Calculate the mass of butane gas needed to boil 500.0 mL of water if the initial temperature of the water was 23.0°C.

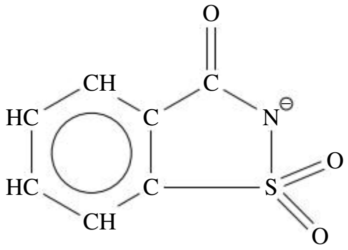
9 Ammonium sulfate and calcium hydroxide react in accordance with the following equation:

$$(NH_4)2SO_4(s) + Ca(OH)_2(s) \rightarrow CaSO_4(s) + 2NH_3(g) + 2H_2O(l)$$

A mixture of 109 g of ammonium sulfate and 53.0 g of calcium hydroxide is gently heated, causing the reaction to occur.

Calculate the mass of ammonia produced.

10 A liquid sweetener contains the artificial sweetener calcium saccharin. The negative ion of saccharin is shown below:



To determine the percentage, by mass, of calcium saccharin in the liquid sweetener, 52.4 g of the liquid sweetener was completely converted to saccharin ions and sulfate ions.

A solution of barium chloride was then added to form an insoluble solid called barium sulfate, from all the sulfate ions. The insoluble barium sulfate solid was filtered from the mixture, dried and weighed. Its mass was 1.27 g.

- **a** Determine the formula of calcium saccharin.
- **b** Calculate the amount, in mole, of barium sulfate formed from the liquid sweetener.
- **c** Calculate the amount, in mole, of sulfur atoms contained in this amount of barium sulfate.
- **d** Assuming all the sulfur atoms in the barium sulfate were originally present in the calcium saccharin, calculate the mass of calcium saccharin in the 52.4 g sample of the liquid sweetener.
- **e** Calculate the percentage, by mass, of calcium saccharin in the liquid sweetener.

11 A white crystalline solid of molecular formula $C_xH_yO_z$ was isolated from certain fruits and found to have a relative molecular mass of 180.

A 0.270 g sample of this solid was burnt to completion in air and produced 0.400 g of carbon dioxide and 0.160 g of water. Assume that all the carbon in the compound was converted to CO_2 and that all the hydrogen was converted to H_2O.

- **a** Calculate the masses of carbon, hydrogen and oxygen present in the compound.
- **b** Determine the empirical formula of the compound.
- **c** Determine the molecular formula of the compound.

12 Hydrated salts have set number water molecules per unit formula incorporated into their crystal structure. One example of a hydrated salt is copper sulfate pentahydrate, $CuSO_4{\cdot}5H_2O$. The water molecules, known as water of crystallisation, can be driven off by heating the hydrated salt.

$$CuSO_4{\cdot}5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$

The product does not contain any water of crystallisation and is known as anhydrous copper sulfate.

Outline a procedure you would use to experimentally determine the number of water molecules per unit formula of hydrated barium chloride, which can be represented as $BaCl_2 \cdot xH_2O$.

The following information should be included in your description:

- a list of equipment and chemical to be used
- a diagram showing the arrangement of the equipment
- steps in the experimental procedure
- a result table to record any measurements that are to be made
- an outline of steps you would take to calculate the extent of hydration represented by the number x in the above formula.

13 A compound of antimony, antimony sulfide, is used to make safety matches. A small amount of antimony sulfide is mixed with potassium chlorate and glued to the tip of the matchstick. When the match is struck on the side of the matchbox, the red phosphorus on the side of the box ignites, which in turn ignites the antimony sulfide/potassium chlorate mixture. Heat from this reaction then makes the match wood burn.

a Predict the formula of antimony sulfide (which contains the Sb^{3+} ion) and potassium chlorate.

b When antimony sulfide and potassium chlorate react, they form antimony oxide, potassium chloride and sulfur dioxide. Write a balanced equation for this reaction.

c Deduce if the reaction is endothermic or exothermic. Explain how you decided your answer.

d Sketch an energy profile diagram for the reaction of antimony sulfide with potassium chlorate, indicating the relative potential energies of the reactants and products and the activation energy.

14 The density of a liquid was determined by finding the mass of a measured volume of the liquid in a flask.

The following results were obtained.

mass of flask = 30.00 ± 0.01 g

mass of flask + liquid = 46.00 ± 0.01 g

volume of liquid = 20.0 ± 0. 2 mL

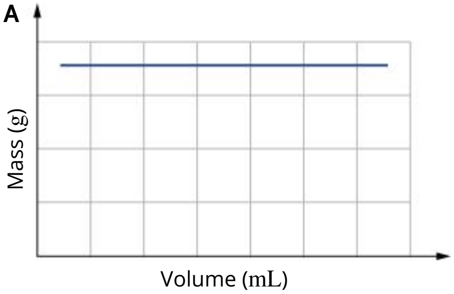
a Use the data provided to calculate the density of the liquid and its uncertainty.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

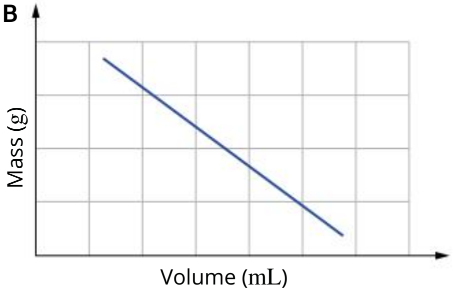
b The mass of different volumes of the same liquid were measured. The results were plotted on a graph using a line of best fit.

Select the graph that correctly represents the relationship between volume and mass of the liquid. Justify your selection.

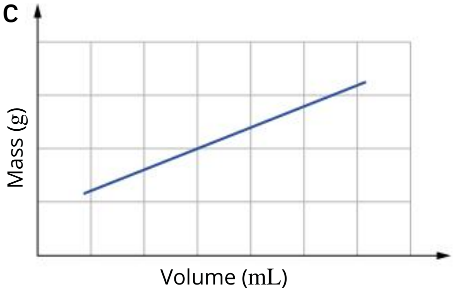
A



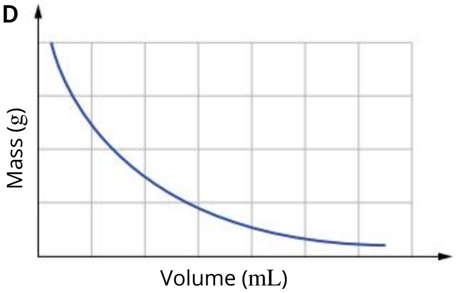
B



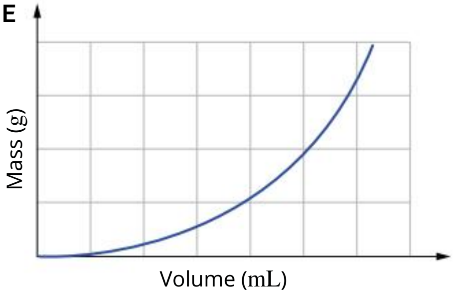
C



D



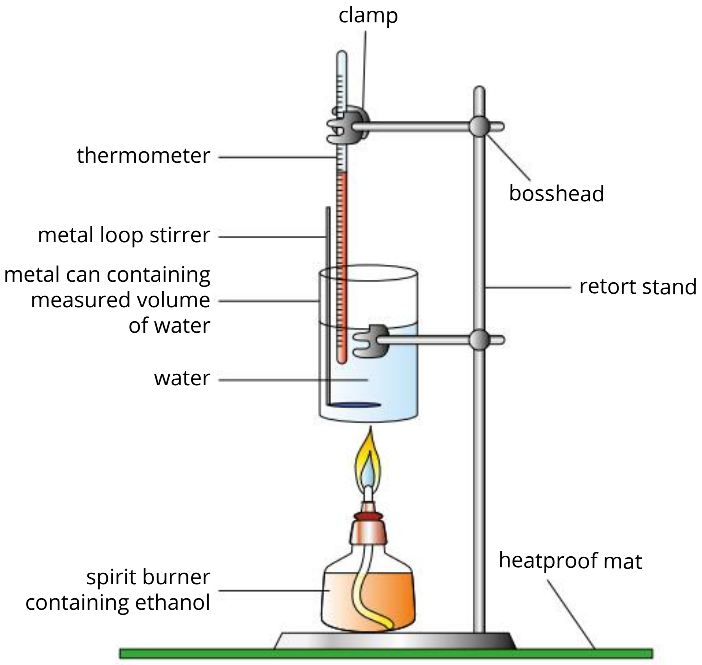
E



c Describe the data gathered in this experiment as either qualitative or quantitative. Explain your selection.

d Explain why analytical chemists aim to produce a set of measurements that are both accurate and precise.

15 The heat of combustion of ethanol, CH_3CH_2OH, was experimentally determined using the apparatus shown below.



The measurements taken in this experiment are provided below.

initial mass of burner and ethanol = 18.615 g
final mass of burner and ethanol = 18.298 g
volume of water = 200.0 mL
initial temperature of water = 21.34°C
final temperature of water = 31.97°C

The following information is also provided.

specific heat of water = 4.18 J g^{-1} °C^{-1}
density of water = 1.00 g mL^{-1}

a Use the experimental data provided to calculate the enthalpy of combustion of ethanol. Express your answer to an appropriate number of significant figures.

b Deduce if the combustion reaction is endothermic or exothermic. Explain your answer in terms of the experimental evidence.

c Discuss the endothermic or exothermic nature of the reaction in terms of bond energies.

d Use the enthalpy value calculated in part **a** above to write a thermochemical equation for the combustion reaction of ethanol.

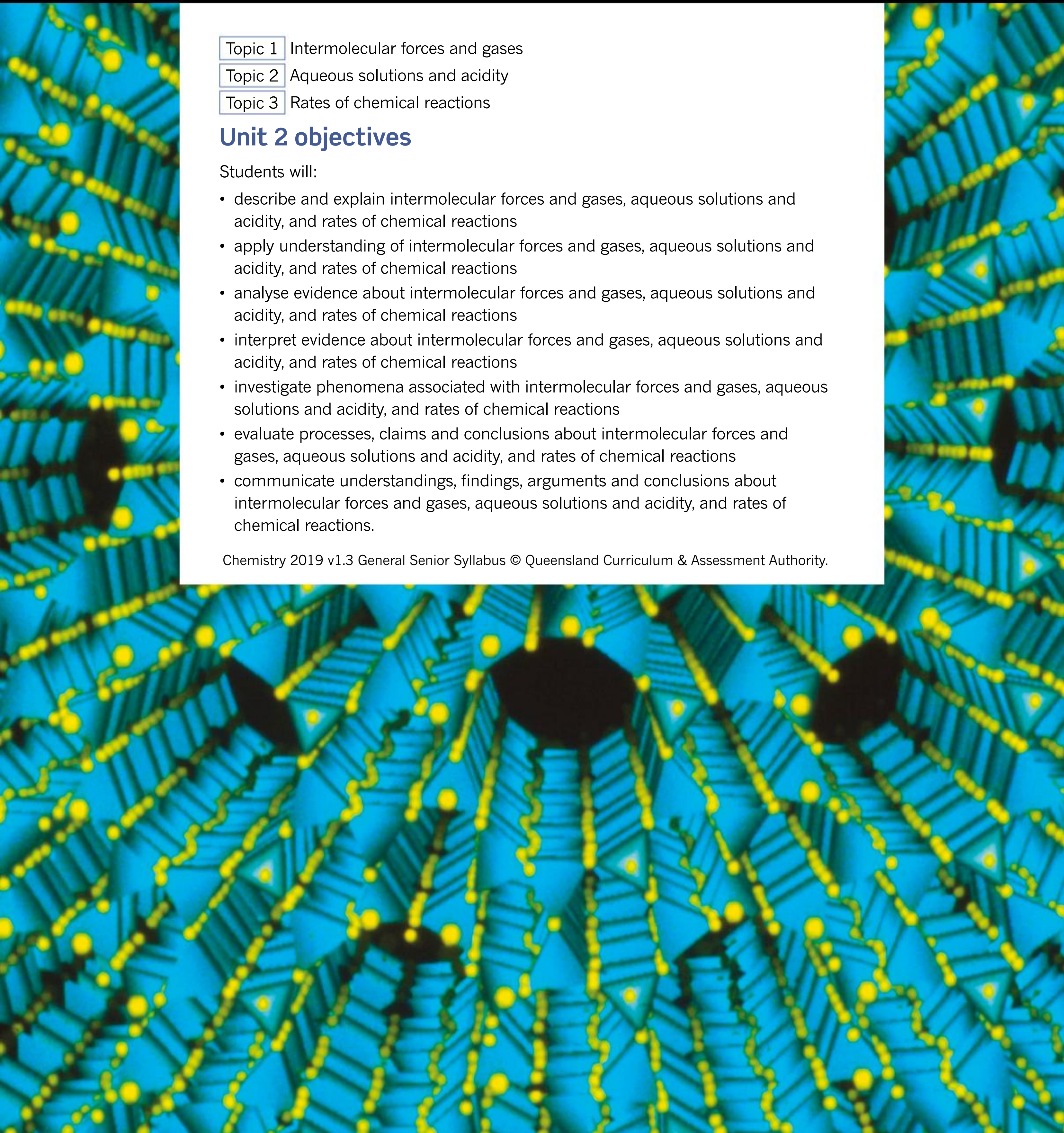
e The value for the heat of combustion of ethanol quoted in chemical data bases is 1367 kJ mol^{-1}.

- **i** Calculate the percentage experimental error of the enthalpy value determined by this experiment.
- **ii** Suggest one cause of this error and indicate how the experimental design could be improved.
- **iii** Does the cause of the error identified in part **ii** produce a random or systematic error? Explain your decision.

f In this experiment, a change in temperature was used to calculate the heat energy released by a reaction. Explain the difference between heat and temperature.

UNIT 2

Molecular interactions and reactions



Topic 1 Intermolecular forces and gases

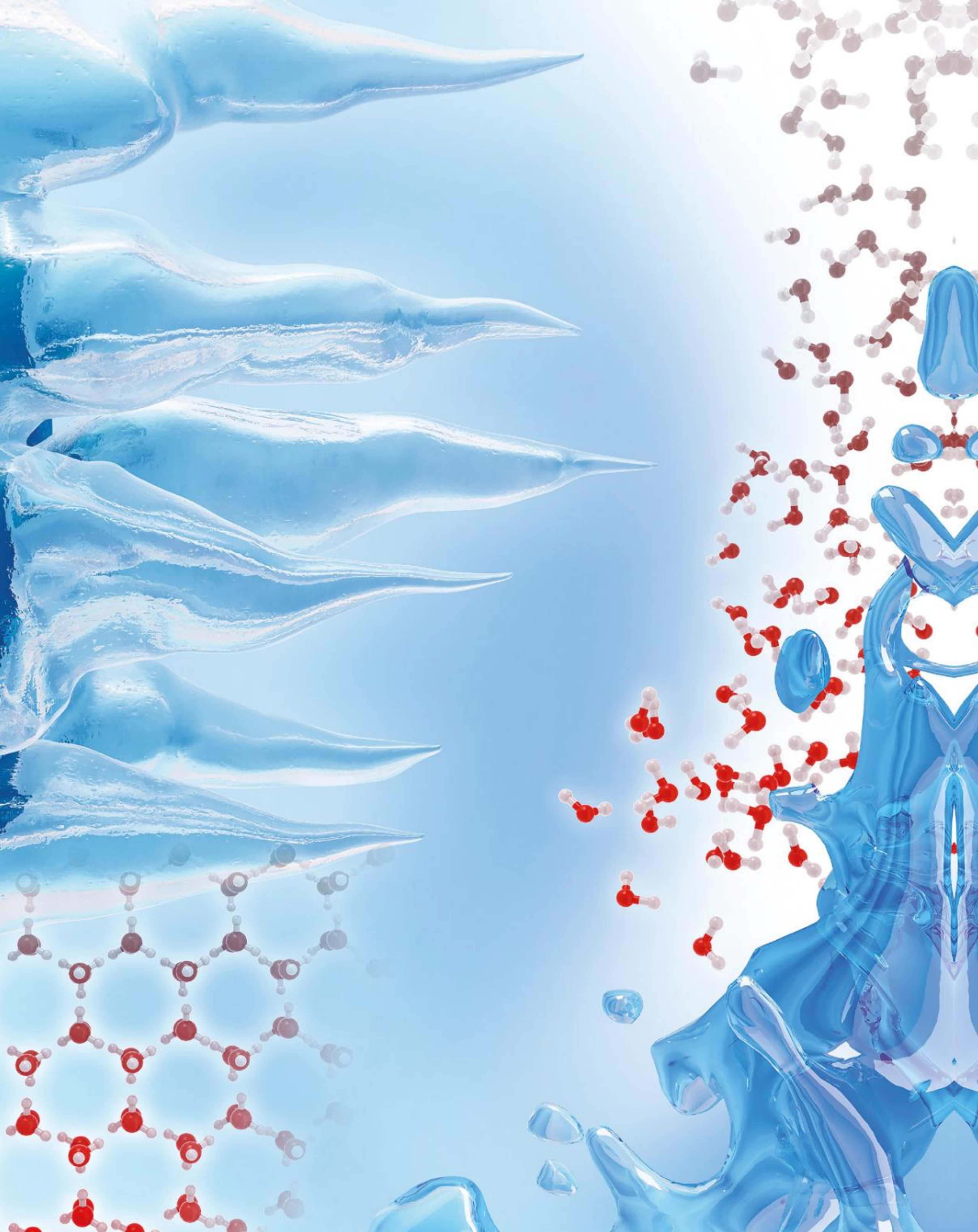
Topic 2 Aqueous solutions and acidity

Topic 3 Rates of chemical reactions

Unit 2 objectives

Students will:

- describe and explain intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
- apply understanding of intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
- analyse evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
- interpret evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
- investigate phenomena associated with intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
- evaluate processes, claims and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
- communicate understandings, findings, arguments and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.



CHAPTER 12

Intermolecular forces

Covalent molecular substances have a much greater range of properties than ionic or metallic substances. Covalent molecular substances are the only substances that can be solids, liquids or gases at room temperature. They can be hard or soft, flexible or brittle, sticky or oily—almost any consistency. This broad range of properties makes molecular substances very useful materials.

The properties of covalent molecular substances can generally be explained by the strong covalent bonds within the molecules (intramolecular bonds) and the much weaker forces of attraction between the molecules (intermolecular forces).

During this chapter, you will become familiar with the valence shell electron pair repulsion (VSEPR) theory and use it to predict the shape and polarity of molecules. Knowing the shapes of molecules and whether or not they are polar is vital when predicting and explaining the type and strength of intermolecular forces acting between the molecules.

You will learn to identify intermolecular forces including dipole–dipole forces, hydrogen bonding and dispersion forces, and consider factors that influence the strength of these forces.

Finally, you will examine how intermolecular forces determine many of the physical properties of covalent molecular substances including vapour pressure, melting and boiling points, and solubility.

WS 2.1.1

Syllabus subject matter

Topic 1 • Intermolecular forces and gases

■ INTERMOLECULAR FORCES

- apply the valence shell electron pair repulsion (VSEPR) theory to predict, draw and explain the shapes of molecules
- use molecular shape, understanding of symmetry and comparison of the electronegativity of elements to explain and predict the polarity of molecules
- explain the relationship between observable properties, including vapour pressure, melting point, boiling point and solubility, and the nature and strength of intermolecular forces, including dispersion forces, dipole–dipole attractions and hydrogen bonding within molecular covalent substances

■ SCIENCE AS A HUMAN ENDEAVOUR

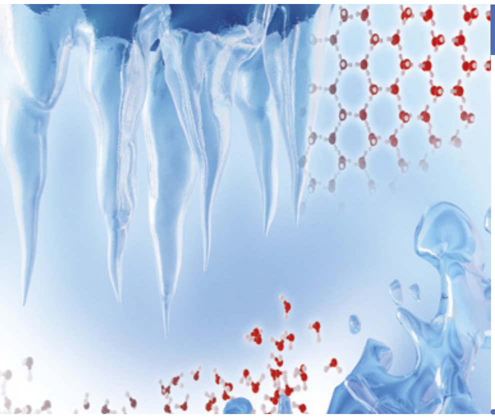
Development of VSEPR theory: Two- and three-dimensional graphical models have been developed and adopted by chemists to represent and communicate the shapes of molecules.

■ MANDATORY PRACTICAL

Construct 3D models (real or virtual) of linear, bent, trigonal planar, tetrahedral and pyramidal molecules.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

12.1 Shapes of molecules



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that the valence shell electron pair repulsion (VSEPR) theory is used to predict the shapes of molecules by assuming the valence shell electron pairs around atoms in a molecule will be arranged to minimise any repulsion between the electron pairs
- use VSEPR theory in conjunction with Lewis structures to determine shapes of molecules
- understand how the presence of lone pairs of electrons affects a molecule's bond angles and be able to deduce the approximate bond angles between atoms in various molecules.

FIGURE 12.1.1 Water (H_2O) molecules have a distinctive shape that is responsible for many of water's properties. The valence shell electron pair repulsion theory accurately predicts the shape of water molecules.

The shapes of molecules are critical in determining many physical properties of covalent molecular substances. In particular, molecular shape affects vapour pressure, melting point, boiling point and solubility. This is because the shape of a molecule determines how it interacts with other molecules.

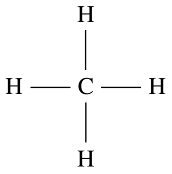
For very large molecules, such as DNA, proteins and enzymes, shape plays a key role in how they behave chemically and biologically. For example, the twisted double helix of DNA allows the molecule to coil up tightly so that it fits inside the nucleus of a cell.

To understand the shape of large molecules like DNA, scientists use complex techniques such as X-ray crystallography or powerful computer simulations. However, the shape of small molecules can be predicted using a relatively simple model known as the **valence shell electron pair repulsion (VSEPR) theory**. In this module, you will see how VSEPR theory can be used to predict the shapes of molecules such as the water molecule shown in Figure 12.1.1.

VALENCE SHELL ELECTRON PAIR REPULSION THEORY

As the name suggests, VSEPR theory uses our knowledge of the **valence electrons** in the atoms of a molecule to predict the shape of the molecule. VSEPR theory is based on the principle that negatively charged electron pairs in the outer shell of an atom repel each other. As a consequence, these electron pairs are arranged as far away from each other as possible.

Molecules with four electron pairs



Lewis structure

FIGURE 12.1.2 This Lewis structure of a methane (CH_4) molecule is made up of four C–H covalent bonds. Each covalent bond represents a pair of bonding electrons.

In Chapter 8, you saw that atoms in covalent molecules are generally most stable when they have eight electrons in their valence shell. This is known as the **octet rule**. These eight electrons are arranged into four pairs of electrons.

For example, consider the **Lewis structure** of the methane molecule shown in Figure 12.1.2 in which the carbon atom is covalently bonded to four hydrogen atoms. Each covalent bond represents a pair of electrons that the carbon atom shares with a hydrogen atom; the four pairs of electrons give the carbon atom a stable octet. Note that hydrogen is an exception to the octet rule as the $n = 1$ shell can only accommodate a maximum of two valence electrons. Each of the four hydrogen atoms in methane shares one pair of electrons with the carbon atom to give them the same electron configuration as helium, a very stable gas.

As discussed in Chapter 8, Lewis structures are often used to simplify the drawing of molecules. In a Lewis structure, lines are used to represent the two electrons in a covalent bond. One line represents one pair of **bonding electrons**.

VSEPR theory states that the electron pairs repel each other so that they are as far apart as possible. In methane, there are four **bonding pairs** of electrons that make up the four C–H covalent bonds. The repulsion between these four pairs of electrons results in the electron pairs being arranged in a **tetrahedral** shape around the central carbon atom as shown in the ball and stick model in Figure 12.1.3. The tetrahedral shape ensures that the electron pairs are as far from each other as possible with angles of 109° between all of the single bonds.

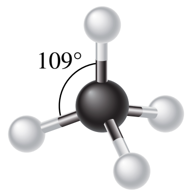


FIGURE 12.1.3 In a methane molecule (CH_4), the electron pairs in the single covalent bonds repel each other. The repulsion forces the bonds as far apart as possible, leading to a tetrahedral molecular shape with bond angles of 109°.

Lone pairs of electrons

Bonding pairs of electrons form the covalent bonds within a molecule. However, not all electron pairs are involved in bonding. Some electrons form a non-bonding pair of electrons known as a **lone pair**. In VSEPR theory, lone pairs are treated in the same way as the bonding electrons.

In the ammonia molecule shown in Figure 12.1.4, the nitrogen atom has a stable octet made up of three bonding pairs of electrons (which form three covalent N–H bonds) and one lone pair. These four pairs of electrons repel each other to give a tetrahedral **electron arrangement**. It is important to understand that VSEPR theory uses both lone pairs and bonding pairs of electrons to determine the electron arrangement and the angles between them. However, the shape of the molecule is derived by considering only the position of the bonding pairs (in other words, the position of the atoms). Therefore, in the case of ammonia, while the electron arrangement is tetrahedral, the lone pair is ignored when determining the shape of the molecule, which is **pyramidal**. The lone pair of electrons occupies slightly more space than the bonding electrons. This results in greater repulsion from the lone pairs; therefore, the covalent bonds between the nitrogen atom and the three hydrogen atoms are pushed closer together. The **bond angles** are 107°, which is slightly less than the 109° in methane.

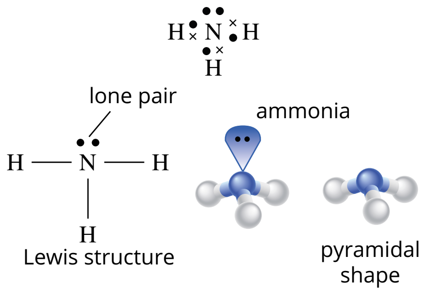


FIGURE 12.1.4 The Lewis structure of an ammonia (NH_3) molecule shows that the nitrogen atom has a total of four pairs of electrons: three bonding pairs (which form the three covalent bonds) and one lone pair. Therefore, the electron arrangement around the nitrogen atom is tetrahedral due to the repulsion of the electron pairs but the molecular shape is pyramidal.

In the water molecule shown in Figure 12.1.5, the oxygen atom has a stable octet made up of two bonding pairs of electrons (which form two covalent O–H bonds) and two lone pairs. In total, there are four pairs of electrons that repel each other to give a tetrahedral electron arrangement. To determine the shape of the molecule, the two lone pairs of electrons are ignored and the shape is described as bent. With two lone pairs in the molecule, the two single covalent bonds are pushed closer together. The bond angle around the central oxygen atom is 104.5°, which is slightly less than the 107° in ammonia, and demonstrates the effect of the extra repulsion from the second lone pair of electrons.

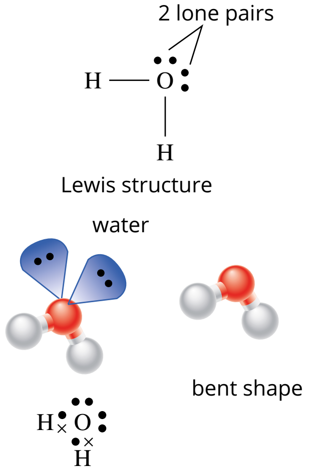
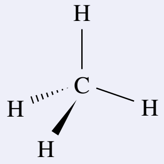
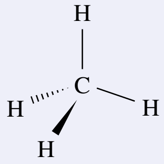
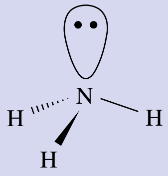
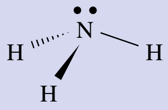
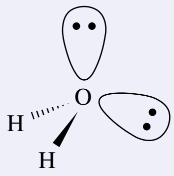
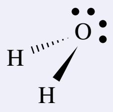


FIGURE 12.1.5 The Lewis structure of a water (H_2O) molecule shows that the oxygen atom has a total of four pairs of electrons: two bonding pairs (which form the two covalent bonds) and two lone pairs. Therefore, the electron arrangement around the oxygen atom is tetrahedral due to the repulsion of the electron pairs but the shape of the molecule is bent.

In Table 12.1.1, the electron arrangement, shapes and bonding angles of methane, ammonia and water are compared. Note that the structures of each molecule in the table are drawn using the **wedge-dash notation** to show the three-dimensional arrangement of the atoms. Using this notation, bonds drawn with lines are in the plane of the page, bonds drawn with bold wedges are coming out of the page towards you, and bonds drawn with dashed lines are going away from you, behind the page. You can see that the methane, ammonia and water molecules all have a tetrahedral electron arrangement but the shape of each molecule depends on the positions of the bonded atoms. You can also see that the bond angles decrease as the number of lone pairs of electrons around the central atom increases; the lone pairs occupy slightly more space than the bonding electrons and push the covalently bonded atoms closer together.

TABLE 12.1.1 A comparison of the electron arrangement, shapes and bonding angles in methane, ammonia and water

Molecule	Electron arrangement	Shape of molecule	Bonding pairs	Lone pairs	Bond angles (°)
methane	 tetrahedral	 tetrahedral	4	0	109
ammonia	 tetrahedral	 pyramidal	3	1	107
water	 tetrahedral	 bent	2	2	104.5

While lone pairs of electrons influence the bond angles, when describing the shape of the molecule only the atoms are considered.

In a hydrogen fluoride molecule, the fluorine atom has a stable octet made up of one bonding pair of electrons (which forms the covalent H–F bond) and three lone pairs. In total, there are four pairs of electrons surrounding the fluorine atom that repel each other to give a tetrahedral electron arrangement. To determine the shape of the molecule, the three lone pairs of electrons are ignored and the shape is described as linear, as shown in Figure 12.1.6.

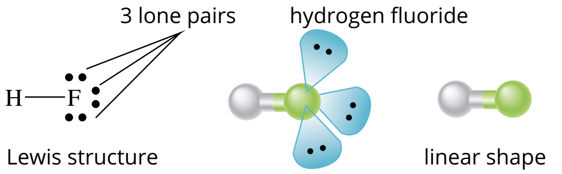
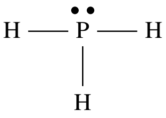


FIGURE 12.1.6 The Lewis structure of a hydrogen fluoride molecule shows that the fluorine atom has a total of four pairs of electrons: one bonding pair (which forms the covalent bond) and three lone pairs. Therefore, the electron arrangement around the fluorine atom is tetrahedral due to the repulsion of the electron pairs but the shape of the molecule is linear.

Worked example 12.1.1

PREDICTING THE SHAPE OF MOLECULES

Predict the shape of a molecule of phosphine (PH_3).	
Thinking	**Working**
Draw the Lewis structure for the molecule.	
Count the number of bonding pairs and lone pairs of electrons on the central atom and determine the total number of electron pairs.	There are three bonding pairs and one lone pair—a total of four pairs of electrons.
Determine how the pairs of electrons will be arranged to get maximum separation.	The four pairs of electrons will be distributed in a tetrahedral arrangement.
Deduce the shape of the molecule by ignoring the lone pairs and considering the arrangement of just the atoms.	The shape of the molecule is pyramidal.

➤ Try yourself 12.1.1

PREDICTING THE SHAPE OF MOLECULES

Predict the shape of a molecule of hydrogen sulfide (H_2S).

Molecules with fewer than four electron groups

As discussed above, the shape of a molecule is determined by considering the repulsion between regions of high electron density around a central atom. VSEPR theory can also be applied to molecules in which the central atom forms multiple bonds and therefore has fewer regions of electron density around the nucleus.

A double bond contains two pairs of electrons. A triple bond contains three pairs of electrons. VSEPR theory treats double and triple bonds in the same way that it treats single bonds. A double and triple bond are both considered to be one group of bonding electrons and the effect is the same as a pair of electrons in a single bond.

For example, consider the molecule in Figure 12.1.7. This represents a molecule of methanal (which is more commonly known as formaldehyde) and has the formula CH_2O. The central carbon atom has a stable octet made up of four pairs of bonding electrons. However, there are only three groups of electrons surrounding the carbon atom as the double bond (which is made up of two pairs of bonding electrons) counts as one group. These three groups of electrons repel each other to get maximum separation and are distributed in a **trigonal planar** arrangement. Furthermore, because there are no lone pairs of electrons surrounding the central carbon atom, the shape of the molecule is also trigonal planar because the atoms form a triangle in one plane. Trigonal planar molecules have bond angles of 120°.

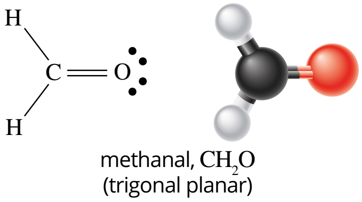


FIGURE 12.1.7 Methanal (also called formaldehyde) has a central carbon atom that forms a double bond with an oxygen atom and single bonds with two hydrogen atoms. The three groups of electrons repel each other to form a trigonal planar electron arrangement. The shape of the molecule is also trigonal planar due to the absence of lone pairs around the central atom.

The central carbon atom in the carbon dioxide molecule shown in Figure 12.1.8 is surrounded by four pairs of bonding electrons. However, there are only two groups of electrons as each double bond (which is made up of two bonding pairs of electrons) counts as one group. The two groups of electrons repel each other to get maximum separation. Since there are no lone pairs of electrons surrounding the central carbon atom, the molecule is linear. Linear molecules have bond angles of 180°.

carbon dioxide, CO_2
(linear)

FIGURE 12.1.8 In a carbon dioxide molecule, the carbon atom forms double bonds with two oxygen atoms. The two groups of electrons repel each other to form a linear electron arrangement. The shape of the molecule is also linear due to the absence of lone pairs around the central atom.

Another example of a linear molecule in which the central atom has a multiple bond is hydrogen cyanide (HCN), as shown in Figure 12.1.9. In this molecule, the central carbon atom has a single bond to the hydrogen atom and a triple bond to the nitrogen atom. Like the double bonds in carbon dioxide, the triple bond between carbon and nitrogen is treated as one group of electrons so that overall there are only two groups of electrons surrounding the central atom.

hydrogen cyanide, HCN
(linear)

FIGURE 12.1.9 The hydrogen cyanide molecule is linear. In this case, the central carbon atom forms a triple bond with the nitrogen atom and a single bond with the hydrogen atom.

ℹ Double and triple bonds count as one group of electrons.

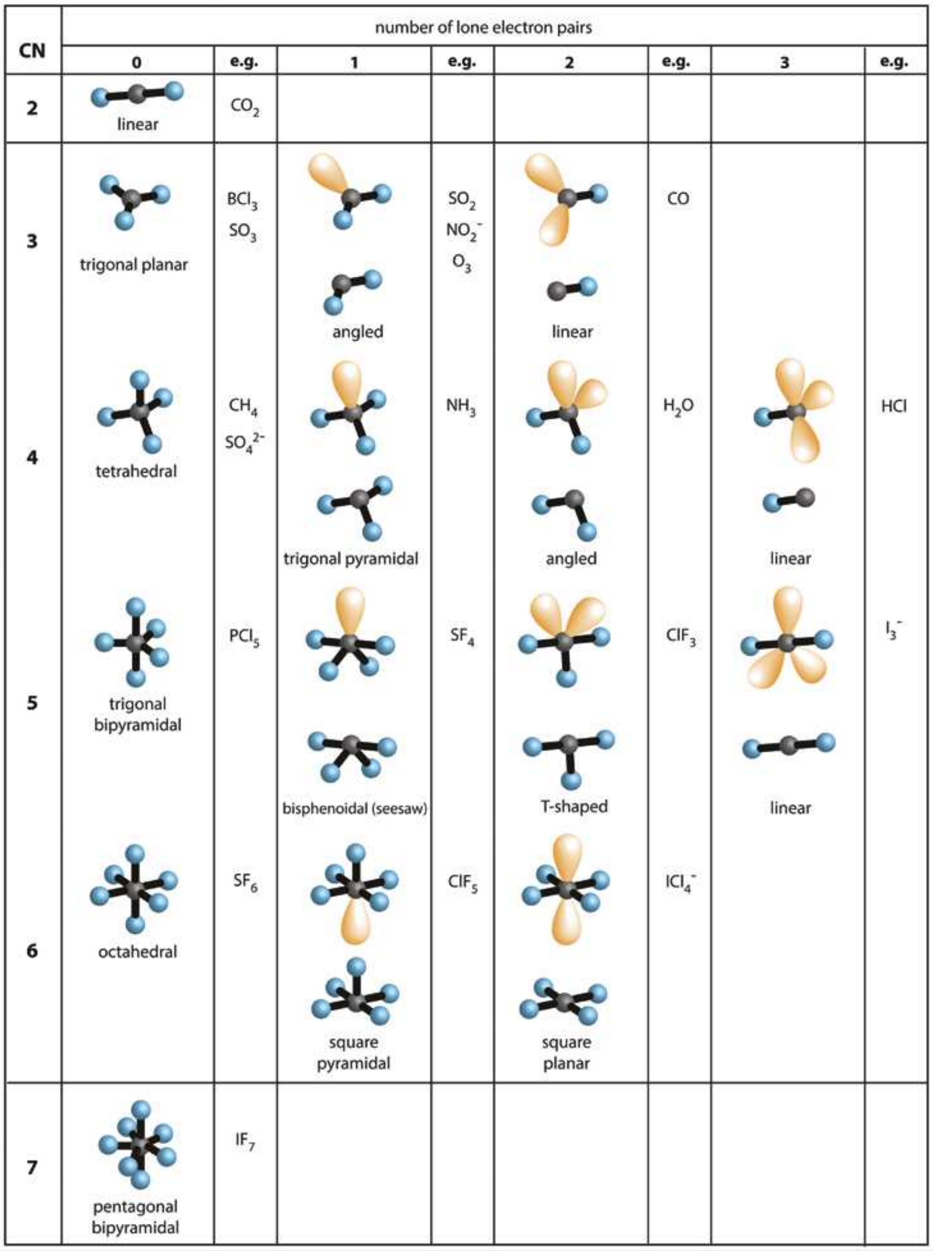
VSEPR theory can be explored further through the mandatory practical at the end of this chapter.

SCIENCE AS A HUMAN ENDEAVOUR

Development of VSEPR theory

The beginnings of VSEPR theory stemmed from the work of Nevil Sidgwick and Herbert Powell who, in 1940, showed that the shape of a molecule was dependent on the total number of electron pairs surrounding a central atom. This concept was refined into a more detailed and useful theory in 1957 by an Australian Professor of Chemistry, Ron Nyholm, and his British colleague, Professor Ronald Gillespie. Prior to the development of VSEPR theory other models were in use for predicting the shapes of molecules, but as they only took into account the positions of the bonded atoms they could not accurately predict and explain the shape of every molecule, such as ammonia and water.

VSEPR theory dealt with this problem by accounting for the lone pairs as well as the bonding pairs of electrons, as shown in Figure 12.1.10. While this solved the problem of predicting the shapes of molecules, the theory still could not explain why the bond angles in some molecules deviated from expected bond angles (such as the bond angles in ammonia and water deviating from the tetrahedral bond angles of 109°). Nyholm and Gillespie overcame this by further developing the theory to include rules about the types of electron repulsions and their relative effect; this led to a much better understanding of why and to what extent such bond angle deviations occur.



CN	number of lone electron pairs							
	0	e.g.	1	e.g.	2	e.g.	3	e.g.
2	linear	CO_2						
3	trigonal planar	BCl_3 SO_3	angled	SO_2 NO_2^- O_3	linear	CO		
4	tetrahedral	CH_4 SO_4^{2-}	trigonal pyramidal	NH_3	angled	H_2O	linear	HCl
5	trigonal bipyramidal	PCl_5	bisphenoidal (seesaw)	SF_4	T-shaped	ClF_3	linear	I_3^-
6	octahedral	SF_6	square pyramidal	ClF_5	square planar	ICl_4^-		
7	pentagonal bipyramidal	IF_7						

FIGURE 12.1.10 The VSEPR theory, refined by Nyholm and Gillespie, can be used to predict the shape of molecules.

Review

1 Discuss why VSEPR theory was an improvement on some of the earlier models for predicting shapes of molecules.

2 Identify the further additions that Nyholm and Gillespie made to the VSEPR theory to make it a better model.

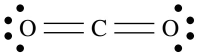
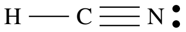
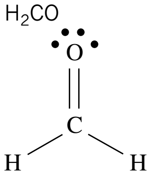
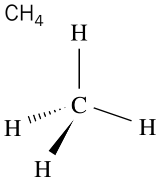
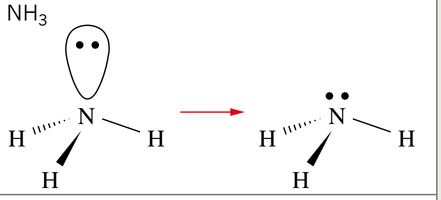
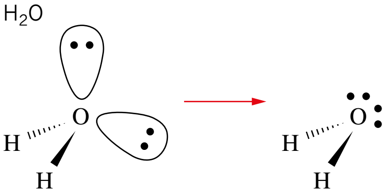
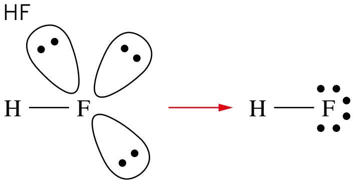
3 Using VSEPR theory, explain the unusual bond angles seen in water.

12.1 Review

SUMMARY

- The shapes of simple molecules can be predicted by the valence shell electron pair repulsion (VSEPR) theory, as shown in Table 12.1.2.
- VSEPR theory is based on the principle that electron pairs in the outer shell of an atom repel each other. As a consequence, these electron pairs are arranged as far away from each other as possible.
- Electron pairs can exist as bonding electrons (in covalent bonds) or as lone pairs of electrons.
- Lone pairs of electrons and bonding electrons repel each other due to the repulsion between clouds of negative charge.
- Lone pairs affect the bond angles in a molecule but are not considered when describing the shape of the molecule.
- Double and triple bonds count as one electron group when determining the electron arrangement around the central atom.

TABLE 12.1.2 Summary of shapes of molecules

Total electron groups	**Electron geometry**	**Bonding groups**	**Lone pairs**	**Molecular shape**	**Approximate bond angles (°)**	**Examples**
2	linear	2	0	linear	180	CO_2  HCN 
3	trigonal planar	3	0	trigonal planar	120	H_2CO 
4	tetrahedral	4	0	tetrahedral	109	CH_4 
		3	1	pyramidal	107	NH_3 
		2	2	bent	104.5	H_2O 
		1	3	linear	180	HF 

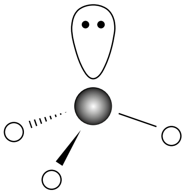
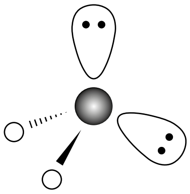
KEY QUESTIONS

Retrieval

1 **a** Identify what VSEPR stands for.
  b Define a lone pair.

Comprehension

2 Copy the table below and use VSEPR theory to identify the molecular shape in each case.

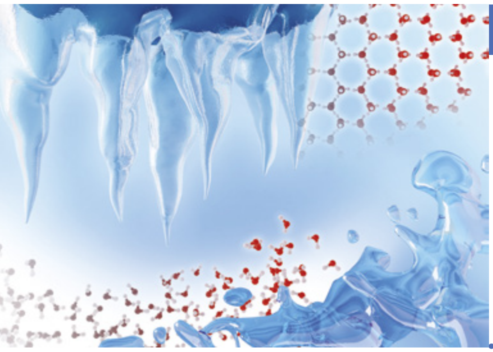
Molecule	**Bonding groups**	**Lone pairs**	**Electron arrangement**	**Molecular shape**
a	3	1		
b	2	2		
c	2	0		

3 Draw the Lewis structure and then use VSEPR theory to determine the shape of each of the following molecules.
  a CF_4
  b H_2Se
  c CS_2

Analysis

4 Consider NH_3, NH_2^- and NH_4^+. Predict the bond angles within each molecule and explain any differences between their shapes and bond angles.

12.2 Polarity of molecules



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the concept of electronegativity and that it is the key factor in determining electron distribution within molecules
- compare the electronegativity of different elements based on their positions in the periodic table
- use electronegativity to predict the polarity of diatomic molecules
- use electronegativity and molecular shape to predict the polarity of polyatomic molecules.

The shape of a molecule is critical in determining many physical properties of covalent molecular substances. In addition to the shape of the molecule, the **electron distribution** within a molecule is also critical to the physical properties of covalent molecules; the type and strength of the intermolecular forces between molecules depend strongly upon each molecule's shape and electron distribution.

Intermolecular forces are an example of **electrostatic forces**. You have already seen many other examples of how electrostatic forces form bonds at the atomic level. The attractions between positive and negative ions in an ionic crystal lattice (Chapter 7) and the attraction of shared electrons and nuclei in a covalent bond (Chapter 8) are all examples of electrostatic forces.

The electrostatic attraction between molecules works in a similar way. An intermolecular force of attraction is the electrostatic attraction between the positive charges in one molecule and the negative charges in a neighbouring molecule. These charges appear as a result of uneven electron distributions within the molecules. In this module, you will examine how the shape of the molecule and the **electronegativity** of its atoms can cause these uneven electron distributions.

ELECTRONEGATIVITY AND POLARITY

Electronegativity is a key factor in determining the electron distribution in molecules. Electronegativity is the tendency of an atom in a covalent bond to attract electrons. Electronegativity increases from left to right across the periods of the periodic table and decreases down the groups of the table, as shown in Figure 12.2.1. You will remember seeing these patterns in Chapter 4.

The **polarity** of a covalent bond is a measure of the unequal sharing of electrons and is a result of the difference in the electronegativities of the bonded atoms. A diatomic molecule contains only one covalent bond, so its molecular polarity is simply determined by its bond polarity. However, the polarity of a polyatomic molecule depends not only on the individual bond polarities but also on how those individual bonds are arranged (i.e. the shape of the molecule).

Polarity of diatomic molecules

When two atoms form a covalent bond, you can regard the atoms as competing for the electrons being shared between them. The atom that has the stronger pull on the electrons is the one with the greater electronegativity.

Non-polar diatomic molecules

If the covalent bond is between atoms of the same element, then the atoms will have identical electronegativities and the electrons will be shared equally between them. This is the case for diatomic molecules such as chlorine (Cl_2), oxygen (O_2), hydrogen (H_2) and nitrogen (N_2).

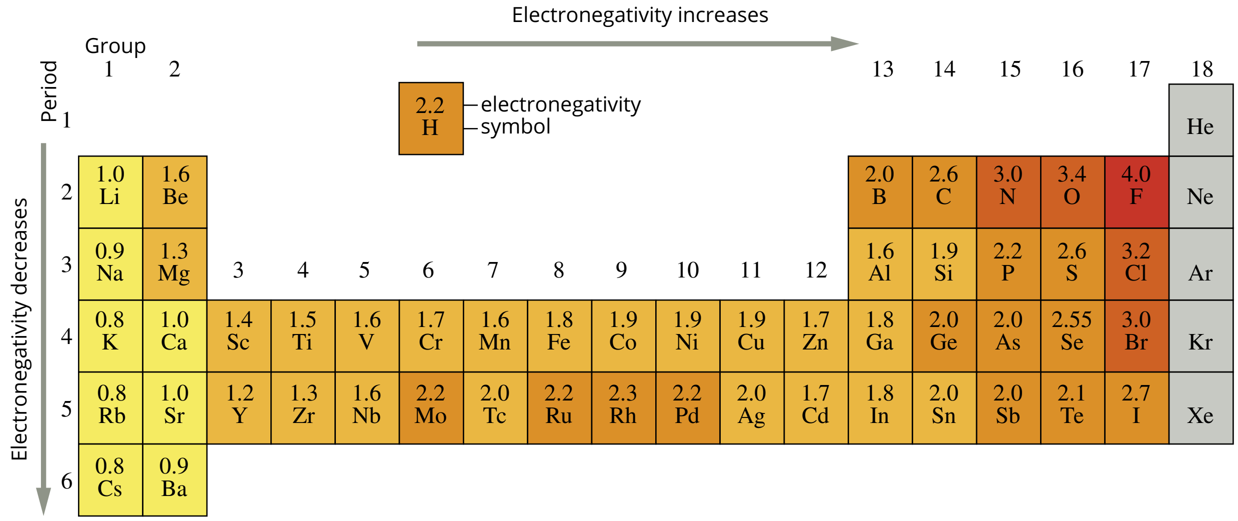
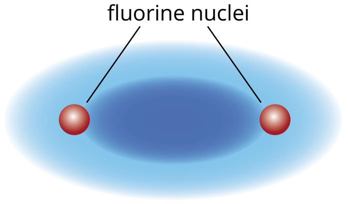


FIGURE 12.2.1 Table of electronegativity values. This periodic table shows the electronegativities of the atoms of each element. The electronegativities increase from left to right across the periods and decrease down the groups.

Bonds with an even distribution of valence (outer shell) electrons are said to be **non-polar** because there is no charge on either end of the molecule.

Figure 12.2.2 shows the electron distribution in the non-polar fluorine molecule (F_2). The molecule has a high **electron density** between the two fluorine atoms, forming the covalent bond. However, the valence electrons are distributed evenly between the two atoms, making the bond non-polar.



F_2 contains a non-polar bond.

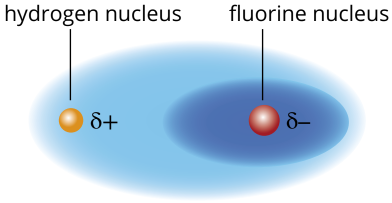
FIGURE 12.2.2 Fluorine molecules have an even distribution of electrons and are therefore non-polar.

Polar diatomic molecules

If the covalent bond is between atoms of two different elements, then the electrons will stay closer to the more electronegative atom as it has a stronger pull on the electrons in the bond. An example is the hydrogen fluoride (HF) molecule, shown in Figure 12.2.3. Bonds with an uneven distribution of electrons are said to be **polar**.

A fluorine atom is more electronegative than a hydrogen atom. Therefore, in a hydrogen fluoride molecule the shared electrons tend to stay closer to the fluorine atom. This imbalance in the electron distribution means the fluorine atom is slightly negatively charged and the hydrogen atom is slightly positively charged. A molecule with oppositely charged poles is said to be polar. The fluorine atom is described as having a partial negative charge, which is represented with the Greek letter delta as δ–. The hydrogen atom is described as having a partial positive charge, δ+. The separation of these partial positive and negative charges is known as an **electric dipole**, or simply as a **dipole**, because there are oppositely charged poles at each end of the molecule.

Hydrogen fluoride is polar because it has a permanent electric dipole created by the different electronegativities of the two atoms. All diatomic molecules that are made up of atoms with different electronegativities are polar to some extent. The level of polarity will depend on the difference in the electronegativities of the two atoms. The greater the difference, the greater the polarity of the molecule.



HF contains a polar bond; fluorine is more electronegative than hydrogen.

FIGURE 12.2.3 The electron distribution in hydrogen fluoride is uneven because of the different electronegativities of the hydrogen and fluorine atoms. Hydrogen fluoride is an example of a polar molecule.

The partial charges on polar molecules are different from the charges on ions. The atom at the partially negative end of the bond can be considered as having a larger share of the electrons (or an excess of electrons). Conversely, the atom at the partially positive end of the bond can be considered as having lost some of its share of the electrons that would have balanced the positive charge of the protons. Consequently, this end is slightly 'naked' in terms of electrons—exposing the positive charge beneath. However, this charge separation is not complete—there are no electrons lost or gained as happens in the formation of ions. The partial charges on a polar molecule will always add to give a total charge of zero so the molecule will be neutral overall. The partial charges on a polar molecule are also a lot weaker than the charges on ions.

It is not just the covalent bonds in diatomic molecules that can be polar. Covalent bonds in larger molecules can have some degree of polarity. The polarity of any covalent bond can be compared by examining the difference in the electronegativities of the atoms involved in the bond.

As the difference in electronegativities of two atoms increases, the covalent bond formed between them increases in polarity.

Worked example 12.2.1

COMPARING THE POLARITY OF DIATOMIC MOLECULES

Compare the polarity of hydrogen fluoride (HF) and carbon monoxide (CO).	
Thinking	**Working**
Use the values in Figure 12.2.1 on page 357 to find the electronegativities of the atoms in each molecule.	HF: hydrogen 2.2; fluorine 4.0 CO: carbon 2.6; oxygen 3.4
For each molecule, subtract the lowest electronegativity value from the highest value.	HF: 4.0 – 2.20 = 1.78 CO: 3.4 – 2.6 = 0.89
Compare the electronegativity differences in both molecules; the more polar molecule has the biggest difference.	HF is more polar than CO.

➤ Try yourself 12.2.1

COMPARING THE POLARITY OF DIATOMIC MOLECULES

Compare the polarity of nitrogen monoxide (NO) and hydrogen chloride (HCl).

Polarity of polyatomic molecules

Determining the polarity of **polyatomic molecules** (those with more than two atoms) is a little more complicated. This is because the polarity of polyatomic molecules depends on the shape of the molecules as well as the polarity of the covalent bonds within the molecules. As a general rule:

- **symmetrical molecules** are non-polar
- **asymmetrical molecules** are polar.

In this context, when considering the diagrams that follow, it is useful to imagine that the electrons are being pulled towards the slightly negative end of the dipoles; then to consider whether the direction and magnitude of the individual dipoles would cancel each other out. If they do, the molecule would be classified as symmetrical and be non-polar (an analogy for this would be a dead heat in a 'tug-of-war' for electrons). If the dipoles are not cancelled out, then the molecule is considered asymmetrical and the molecule would have a **net dipole** (an uneven distribution of charge over the molecule as a whole) and be polar.

Electron density is the measure of the probability of an electron being present at a particular location within an atom. In molecules, areas of electron density are commonly found around the atom and its bonds.

Non-polar polyatomic molecules

As discussed above, even molecules with polar covalent bonds can be non-polar if the molecule is symmetrical.

Figure 12.2.4a shows a molecule of methane in which the carbon atom is slightly more electronegative than the hydrogen atoms. Therefore, the carbon atom attracts the electrons more strongly than each hydrogen atom and has a partial negative charge, leaving hydrogen with a partial positive charge. However, the molecule has a perfect tetrahedral shape and is therefore symmetrical. The symmetry of the molecule means that the individual dipoles of the covalent bonds cancel each other perfectly (Figure 12.2.4b). The result is a molecule with no net dipole. Methane is an example of a non-polar molecule with slight polar covalent bonds (due to slight differences in electronegativities of C and H).

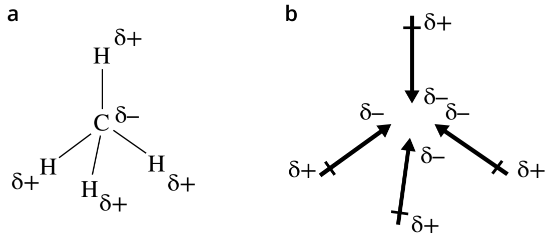


FIGURE 12.2.4 (a) Structure of a methane molecule showing the partial charges on the atoms. (b) The individual dipoles are distributed symmetrically around the molecule.

Polar polyatomic molecules

In asymmetrical molecules, the individual dipoles of the covalent bonds do not cancel each other. This results in a net dipole, making the molecule polar.

The chloromethane molecule shown in Figure 12.2.5a is an example of an asymmetrical molecule. The chlorine atom is more electronegative than the carbon atom. Therefore, the chlorine atom attracts electrons from the carbon atom while the carbon atom attracts electrons from the hydrogen atoms. The individual dipoles of the covalent bonds are shown in Figure 12.2.5b. These add to give the molecule a net dipole as shown in Figure 12.2.5c.

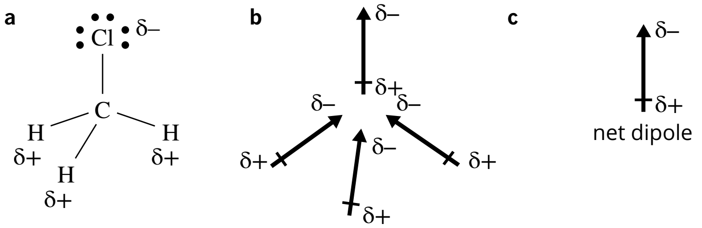


FIGURE 12.2.5 (a) Structure of a chloromethane molecule showing the partial charges on the atoms. (b) The individual dipoles are added together. (c) The result is a net dipole.

The ammonia molecule shown in Figure 12.2.6a is another example of an asymmetrical molecule. In this case, although all the bonds are identical their arrangement around the central nitrogen atom does not cancel out in all directions. The nitrogen atom is more electronegative than the hydrogen atom. Therefore, the nitrogen atom attracts electrons from the hydrogen atoms. The individual dipoles of the covalent bonds are shown in Figure 12.2.6b. These add to give the molecule a net dipole as shown in Figure 12.2.6c.

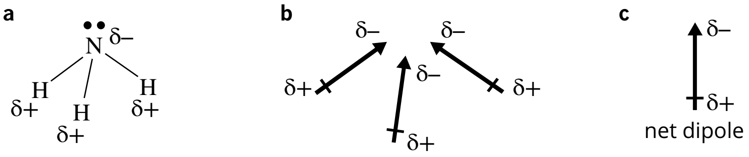
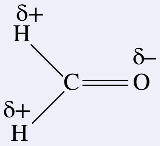
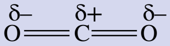
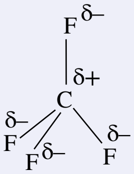
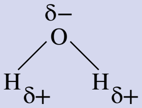
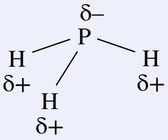


FIGURE 12.2.6 (a) Structure of an ammonia molecule. (b) An ammonia molecule showing the partial charges on the atoms. (c) The individual dipoles are added together. The result is a net dipole.

In general, symmetrical molecules with polar bonds are non-polar and asymmetrical molecules with polar bonds are polar.

Table 12.2.1 shows more examples of how symmetry determines the polarity of covalent molecules.

TABLE 12.2.1 Examples of polar and non-polar covalent molecules

Molecule	Valence structure	Symmetrical/asymmetrical	Polar/non-polar
formaldehyde (methanal)		asymmetrical	polar
carbon dioxide		symmetrical	non-polar
tetrafluoromethane		symmetrical	non-polar
water		asymmetrical	polar
phosphine		asymmetrical	polar

WS 2.1.2

12.2 Review

SUMMARY

- Electronegativity is the tendency of an atom in a covalent bond to attract electrons.
- The polarity of a covalent bond is the result of the difference in the electronegativities of the bonded atoms.
- As the electronegativity difference between two atoms increases, a covalent bond increases in polarity.
- Diatomic molecules containing the same type of atom are non-polar.
- The polarity of a polyatomic molecule depends upon the shape of the molecule and the individual bond dipoles.
- In general, symmetrical molecules with polar bonds are non-polar and asymmetrical molecules with polar bonds are polar.

KEY QUESTIONS

Retrieval

1 Define each of the following.

  a non-polar bond

  b polar bond

2 Use the electronegativity values given in Figure 12.2.1 on page 357 to list the following diatomic molecules in order of increasing polarity: HCl, N_2, HF, CO.

Comprehension

3 Determine whether each of the following molecules is polar or non-polar.

  a CCl_4

  b $CHCl_3$

  c CH_2Cl_2

  d CH_3Cl

  e CH_4

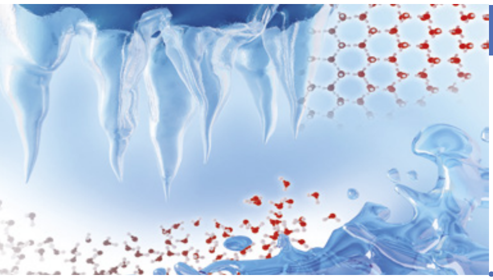
4 For each of the following pairs of molecules, determine which is the polar molecule. On the polar molecule, indicate the polarities of the bonds and the direction of the net dipole.

  a CHF_3 or CF_4

  b CO_2 or SO_2

  c NH_3 or CH_4

12.3 Intermolecular forces



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the three types of intermolecular forces: dipole–dipole forces, hydrogen bonding and dispersion forces
- identify the types of intermolecular forces present between specific molecules in the liquid and solid states
- compare the strengths of the three types of intermolecular forces.

FIGURE 12.3.1 These bubbles are held together by intermolecular forces.

FIGURE 12.3.2 The surface tension created by intermolecular forces allows insects to walk on water.

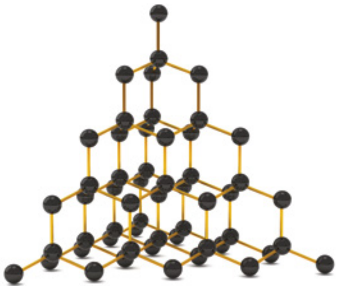


FIGURE 12.3.3 Diamond is a covalent network solid made up of a lattice of carbon atoms that are bonded with strong covalent bonds.

The molecules in covalent molecular substances are held together in the liquid and solid states by intermolecular forces. Intermolecular forces are the electrostatic attractive forces that exist *between* molecules and their effects are seen everywhere. If you have ever struggled to flatten out a sheet of cling wrap, then you have experienced the effects of intermolecular forces. Glue is an example of how intermolecular forces can be used to stick things together. The bubbles shown in Figure 12.3.1 are held together by intermolecular forces. Intermolecular forces are also responsible for creating the **surface tension** that allows insects to walk on water, as shown in Figure 12.3.2.

Intermolecular forces are 10–100 times weaker than strong intramolecular bonds such as ionic, metallic and covalent bonds. For example, as shown in Figure 12.3.3, diamond is a **covalent network solid** made up of a huge lattice of carbon atoms that are bonded with strong covalent bonds. This extremely strong lattice is the reason why diamond has a very high melting point and is the hardest natural substance on Earth.

The atoms in covalent molecules such as water (H_2O) and carbon dioxide (CO_2) are also held together by strong covalent bonds. However, covalent molecular substances tend to have much lower melting and boiling points than ionic, metallic and covalent network substances. This is because the forces between the molecules are much weaker and it is these weak intermolecular forces that are broken when a covalent molecular substance is converted from a solid to a liquid or from a liquid to a gas. Figure 12.3.4 shows the strong covalent bonds that hold the hydrogen and oxygen atoms together in a water molecule and the much weaker (and more easily broken) intermolecular forces that exist between the water molecules.

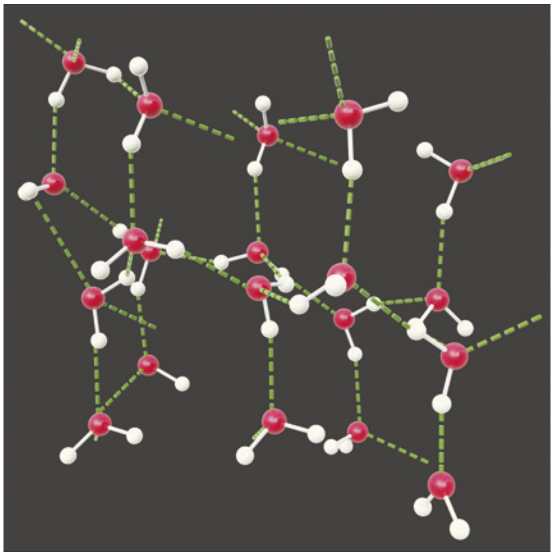


FIGURE 12.3.4 The intermolecular forces between the water molecules are much weaker than the strong covalent bonds that hold the hydrogen and oxygen atoms together within the molecules.

TYPES OF INTERMOLECULAR FORCES

Many factors determine the type and strength of intermolecular forces present in covalent molecular substances, including the size, shape and polarity of the molecules.

There are three main types of intermolecular forces:

- dipole–dipole forces
- hydrogen bonding
- dispersion forces.

In this module, you will examine the nature of these types of intermolecular forces and begin to understand their role in determining the physical properties of covalent molecular substances.

Dipole—dipole forces

Dipole–dipole forces only occur between polar molecules. These forces result from the attraction between the positive and negative ends of polar molecules, as shown in Figure 12.3.5.

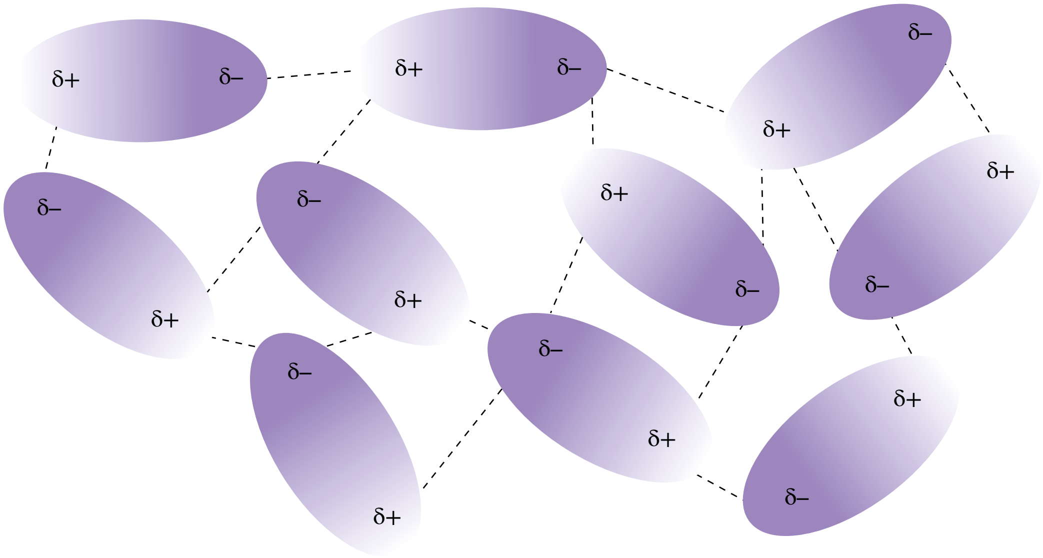


FIGURE 12.3.5 The positive and negative ends of polar molecules attract each other.

Dipole–dipole forces are relatively weak since the partial charges (the δ+ and δ–) on the molecules are small. However, the more polar a molecule is, the stronger the dipole–dipole forces are. The polarity will be larger when there is a large difference in the electronegativities of the atoms and a high degree of asymmetry in the shape of the molecule.

Dipole–dipole forces in a molecular substance directly affect its melting and boiling points. A substance that contains stronger dipole–dipole forces has a higher melting and boiling point than substances of a similar size that contain either weaker dipole–dipole forces or no dipole–dipole forces at all. This is because it takes more energy (i.e. higher temperatures) to break the stronger dipole–dipole forces when the substance changes from a solid to a liquid or a liquid to a gas.

For example, compare methanal (also called formaldehyde, CH_2O) and ethane (CH_3CH_3), as shown in Figure 12.3.6. Methanal molecules are asymmetrical and polar. The molecules form dipole–dipole forces in the liquid state, which is the reason why methanal has a relatively high boiling point of –19°C. On the other hand, ethane molecules are symmetrical and non-polar. They do not form dipole–dipole forces and ethane has a much lower boiling point of –88.5°C.

Molecules are more polar when there is a large difference in the electronegativities of the atoms and a high degree of asymmetry in the shape of the molecules. Molecules that are more polar form stronger dipole–dipole forces.

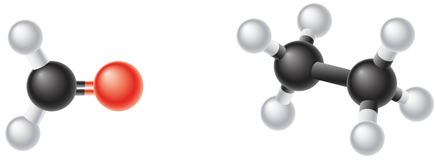


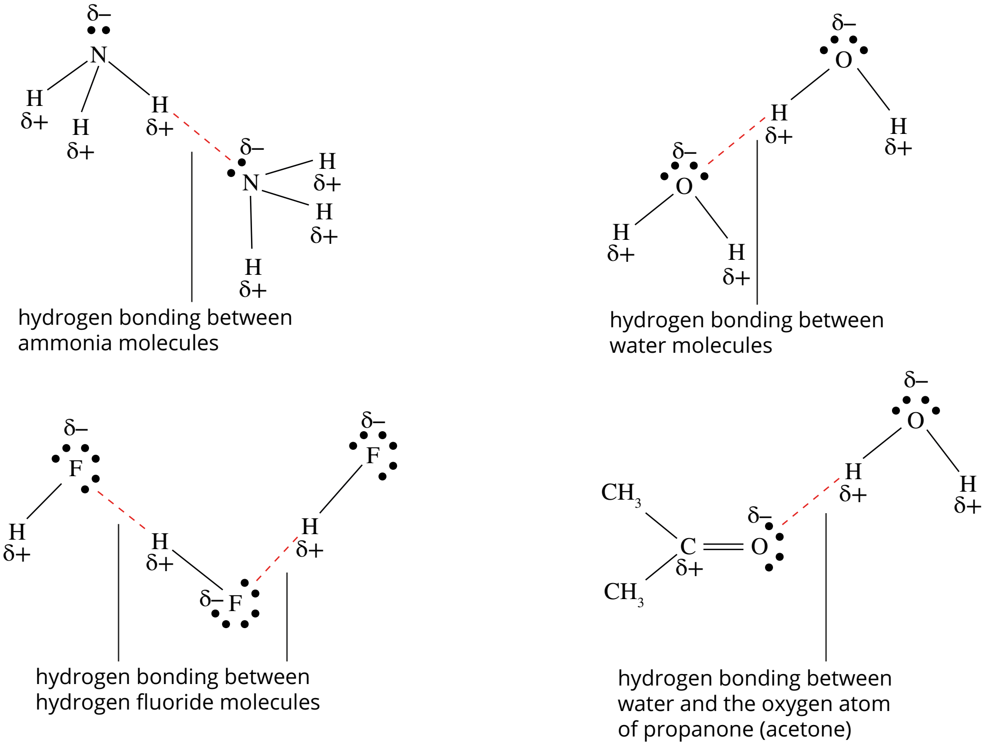
FIGURE 12.3.6 Molecular structures of methanal and ethane. Dipole–dipole forces exist between molecules of methanal because methanal is polar.

Hydrogen bonding

Hydrogen bonding is a particularly strong form of dipole–dipole force. Hydrogen bonding only occurs between highly polar molecules in which a hydrogen atom is covalently bonded to an oxygen, a nitrogen or a fluorine atom.

Oxygen, nitrogen and fluorine atoms are small and highly electronegative. Therefore, when bonded with a hydrogen atom, they strongly attract the electron pair in the covalent bond and a large dipole results. Remember that hydrogen only has one electron and it is pulled towards the highly electronegative atom in the bond. The hydrogen nucleus (a proton) is therefore left exposed and is attracted to the lone pairs of electrons on the nitrogen, oxygen or fluorine atom of a neighbouring molecule. The small size of the hydrogen atom allows the neighbouring molecule to closely approach and the resulting attractive force is relatively strong.

This intermolecular force is known as a hydrogen bond. It is approximately ten times stronger than a dipole–dipole force but about one-tenth the strength of an ionic or a covalent bond. Figure 12.3.7 shows examples of hydrogen bonding among molecules that all contain a hydrogen atom covalently bonded to either an N, O or F atom.



hydrogen bonding between ammonia molecules

hydrogen bonding between water molecules

hydrogen bonding between hydrogen fluoride molecules

hydrogen bonding between water and the oxygen atom of propanone (acetone)

FIGURE 12.3.7 Examples of hydrogen bonding between molecules containing a hydrogen atom covalently bonded to either an N, O or F atom.

The presence of hydrogen bonds results in much higher melting and boiling points. Figure 12.3.8 demonstrates the effect of hydrogen bonding on boiling point by comparing methanol to methanal and ethane. Recall that ethane contains non-polar molecules that do not form dipole–dipole forces. Methanal contains polar molecules that attract each other through dipole–dipole forces but not hydrogen bonds. Methanol, however, contains polar molecules that attract each other through hydrogen bonds. This is because each methanol molecule contains a hydrogen atom attached to an oxygen atom; as there is a large difference in electronegativities, this part of the methanol molecule is highly polar and can hydrogen bond with neighbouring methanol molecules. As a result, the boiling point of methanol (64.7°C) is significantly higher than the boiling points of methanal (–19°C) and ethane (–88.5°C).

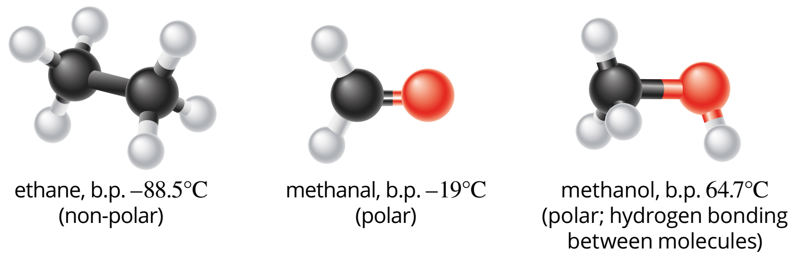


FIGURE 12.3.8 Hydrogen bonding exists between molecules of methanol giving it a much higher boiling point than methanal and ethane.

There are two key requirements for hydrogen bonding:

1 a hydrogen atom covalently bonded to an oxygen, a nitrogen or a fluorine atom
2 a lone pair of electrons on the nitrogen, oxygen or fluorine atoms of neighbouring molecules.

You may wonder why other highly electronegative atoms do not form hydrogen bonds. For example, chlorine atoms have a high electronegativity but because they are larger atoms, the electron density is more spread out and less concentrated. This results in weaker dipole–dipole forces with the hydrogen atoms on neighbouring molecules.

ℹ Hydrogen bonding only occurs in highly polar molecules where a hydrogen atom is bonded to an oxygen, a nitrogen or a fluorine atom.

Hydrogen bonding in water

Water is a very common substance but it has some unique properties, which are covered in more detail in Chapter 15. One such property is that ice floats because it is less dense than liquid water. In comparison, for most other substances the solid form is denser than the liquid.

The fact that ice is less dense than liquid water can be explained by hydrogen bonding. Water has two hydrogen atoms attached to an oxygen atom in a bent shape. As shown in Figure 12.3.9, hydrogen bonds can form between the hydrogen atoms on one molecule and the lone pairs of electrons on the oxygen atoms of neighbouring molecules. In this way, a water molecule can form hydrogen bonds with four other water molecules.

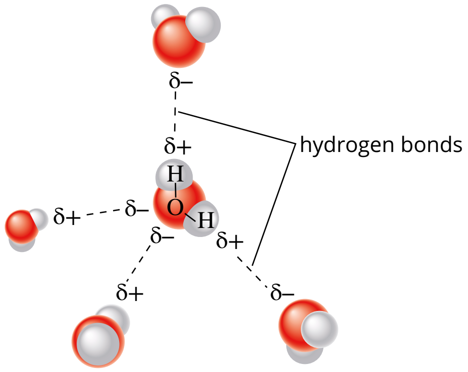


FIGURE 12.3.9 Each water molecule can form hydrogen bonds with four other water molecules.

The hydrogen bonding holds the water molecules in ice in a regular crystal lattice. In this lattice, the molecules are held further apart than in liquid water as shown in Figure 12.3.10 on page 366. As a result, ice is less dense and therefore floats in liquid water.

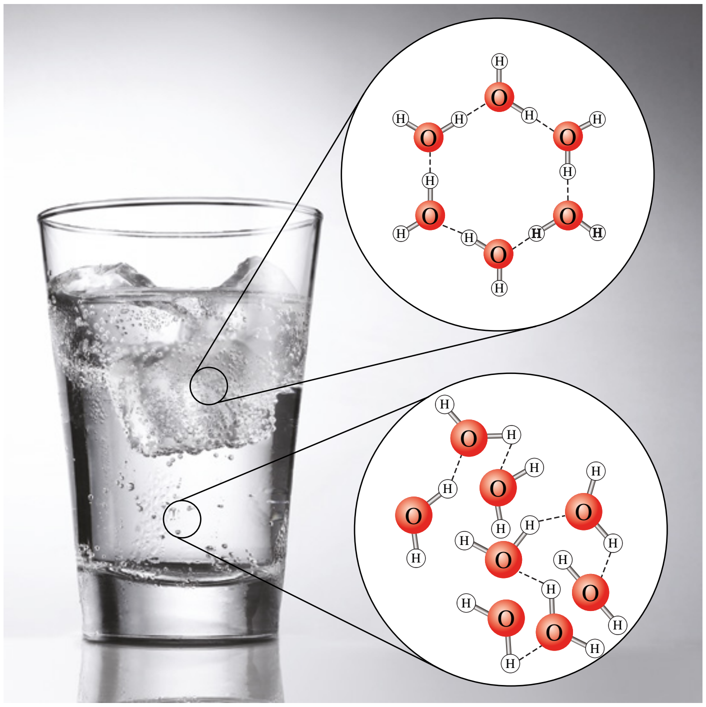


FIGURE 12.3.10 Ice floats because the water molecules form a crystal lattice in which they are spaced more widely apart than in liquid water. This arrangement means ice is less dense than liquid water.

When ice melts and forms liquid water, the density of the water increases rapidly. This is because the open crystal lattice collapses and the water molecules pack together more tightly.

However, as the temperature of the water increases further, water molecules in the liquid move and vibrate more rapidly. The movement causes the molecules to spread further apart. As the molecules move further apart, the liquid becomes less dense.

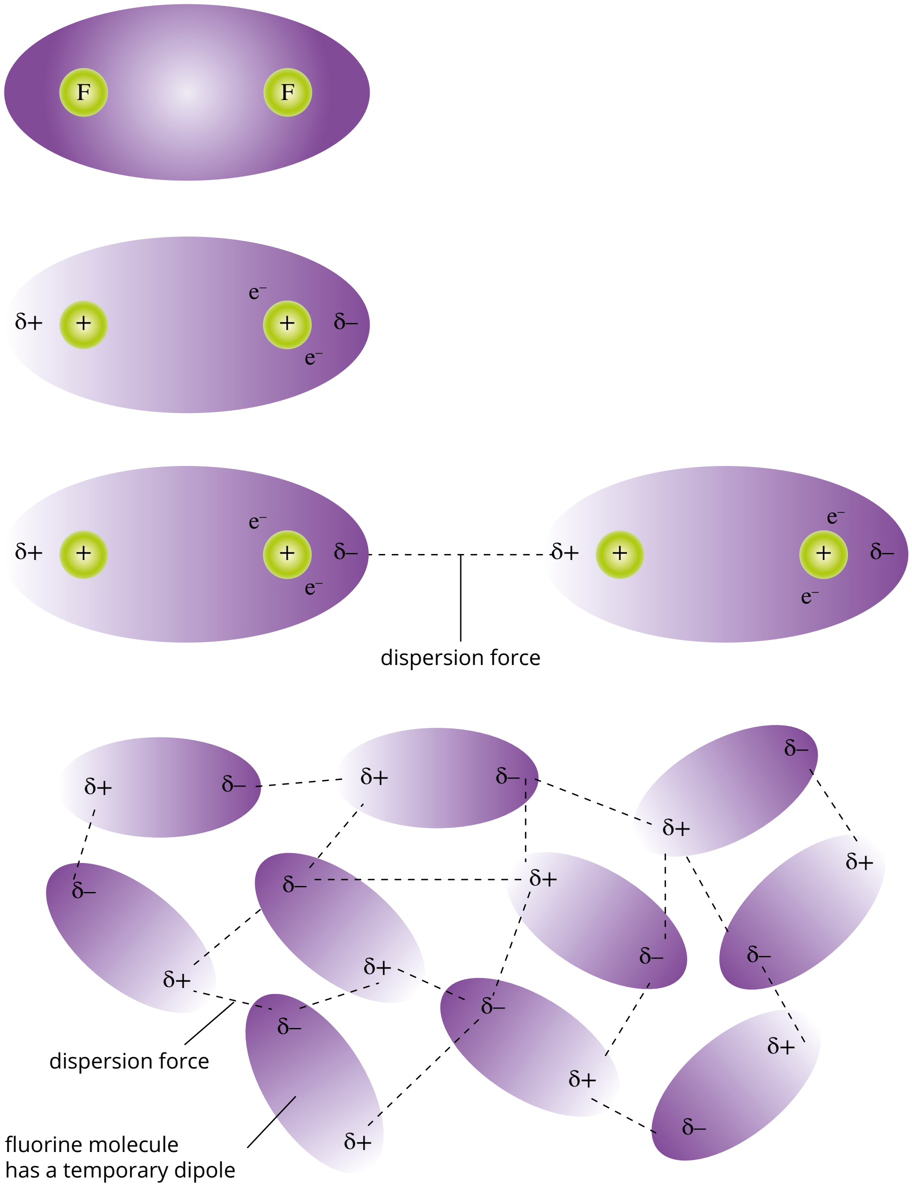
Dispersion forces

Dipole–dipole forces and hydrogen bonding explain the intermolecular forces in polar substances. However, they do not explain the existence of intermolecular forces in non-polar substances.

We know that intermolecular forces are present in non-polar substances, otherwise they could not exist as liquids and solids. Without intermolecular forces, there would be nothing to hold the molecules together and non-polar substances would only exist as gases. However, there are many non-polar compounds that are liquids at room temperature (such as vegetable oil) and even non-polar solids (such as candle wax). In fact, all non-polar substances form liquids or solids if cooled to low enough temperatures. Even hydrogen liquefies below −259°C.

The forces of attraction between non-polar molecules are known as **dispersion forces**. Dispersion forces are caused by **temporary dipoles** in the molecules that are the result of random movement of the electrons surrounding the molecules. These temporary dipoles are also known as **instantaneous dipoles**.

Dispersion forces are always present between molecules, no matter whether they are polar or non-polar, as electrons are constantly in motion within atoms. Consider the fluorine molecule in Figure 12.3.11a. F_2 is non-polar so the electrons are distributed evenly, as represented by the even shading in the figure. However, this is only true over an average period of time. At any one time the electrons might move towards one end of the molecule and produce a temporary partial negative charge at that end and leave a temporary partial positive charge on the opposite end, as shown in Figure 12.3.11b. The molecule is said to be **polarised** and the temporary electrical distortion is called a temporary dipole. These temporary dipoles can then induce (create) dipoles in neighbouring molecules, which subsequently induce dipoles in their neighbours and so on, as shown in Figures 12.3.11c and d.



a In a molecule, the electrons are constantly moving. In the case of a non-polar molecule, such as fluorine, the electrons spend an equal amount of time around each atom.

b Occasionally, the electrons gather more closely together at one end of the molecule, causing one end of the molecule to become partially negative and the other end to become partially positive. This is known as a temporary dipole.

c These temporary dipoles can then induce (create) dipoles in the neighbouring molecules.

d The neighbouring molecules then induce dipoles in their neighbours and so on. The temporary dipoles attract each other to create the intermolecular forces known as dispersion forces.

FIGURE 12.3.11 How dispersion forces form within non-polar substances. Diagrams a–d show dispersion forces forming between fluorine molecules.

Strength of dispersion forces

The strength of dispersion forces between molecules increases as the size of the molecules increase. This is because larger molecules contain more electrons that are, on average, further from the nuclei than those in smaller molecules. Because the electrons are further from the nuclei, they are less tightly held and more easily distorted. Larger molecules are therefore described as being more **polarisable** because they can form temporary dipoles more easily. As a result, larger molecules form stronger dispersion forces.

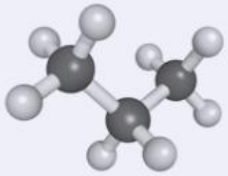
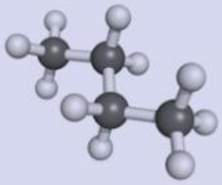
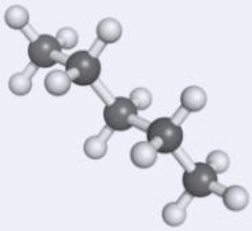
Table 12.3.1 shows the boiling points of the halogens (group 17), which all form non-polar, diatomic molecules. The only forces between their molecules are dispersion forces. You can see that as molecular mass increases, the boiling point also increases. The comparison between increasing molecular mass and boiling point is a useful one. However, it is only used because, in general, as molecular mass increases, the size of the molecules and the number of electrons also increase. Consequently, the molecules become more polarisable and the dispersion forces become stronger, resulting in higher boiling points.

TABLE 12.3.1 The effect of dispersion forces on the boiling points of the halogens

Molecule	Molecular mass (amu)	Number of electrons	Boiling point (°C)
fluorine (F_2)	38.00	18	–188
chlorine (Cl_2)	70.90	34	–35
bromine (Br_2)	159.80	70	59
iodine (I_2)	253.80	106	184

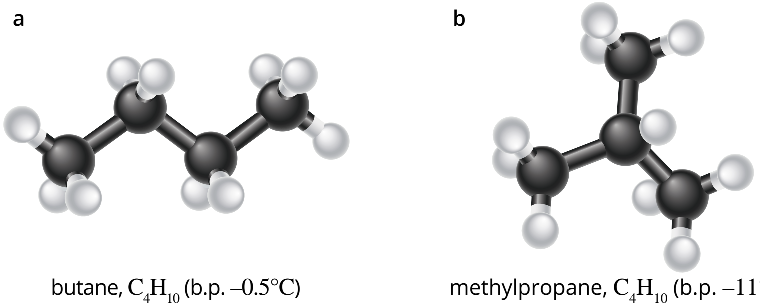
The chain length of a molecule also influences the strength of the dispersion forces. Molecules that form long chains will tend to have stronger dispersion forces than molecules with smaller chains. For example, compare the boiling points of the alkanes propane, butane and pentane, which are shown in Table 12.3.2. As the carbon chain increases in size, the boiling point also increases. This is because the contact area over which the dispersion forces can act with neighbouring molecules increases; the dispersion forces become stronger and the boiling points of the substances increase.

TABLE 12.3.2 The effect of chain length on the boiling points of the alkanes propane, butane and pentane

Molecule	Molecular Mass (amu)	Structure	Boiling point (°C)
propane (C_3H_8)	44.1		–42.2
butane (C_4H_{10})	58.1		–0.5
pentane (C_5H_{12})	72.1		36.3

Finally, the shape of a molecule affects the strength of the dispersion forces. Molecules that form long chains will tend to have stronger dispersion forces than more compact molecules with similar numbers of electrons.

For example, butane and methylpropane (Figure 12.3.12) have the same molecular mass and the same number of electrons because they both contain four carbon atoms and ten hydrogen atoms. The boiling point of butane is –0.5°C, while the boiling point of methylpropane is –11°C. The higher boiling point of butane is because of the different shapes of the two molecules; butane is a long molecule, while methylpropane is compact. Being less compact and long means butane has more contact area to interact with its neighbouring molecules and form stronger dispersion forces.



butane, C_4H_{10} (b.p. –0.5°C)   methylpropane, C_4H_{10} (b.p. –11°C)

FIGURE 12.3.12 Butane and methylpropane have different boiling points because their molecules are different shapes.

Comparing the strengths of intermolecular forces

It is important to remember that dispersion forces occur between all molecules, those that are polar as well as those that are non-polar. However, when comparing molecules that have a similar molecular mass (and are therefore a similar size), the dispersion forces are generally weaker than dipole–dipole forces and much weaker than hydrogen bonds. Recall that in Figure 12.3.8 the boiling points of ethane, methanal and methanol were compared and found to be quite different. These covalent molecular substances have similar molecular masses so the strengths of their dispersion forces are also similar. The differences in boiling point are therefore attributed to the different types of intermolecular forces that are present in each substance. Ethane is a non-polar substance and only contains dispersion forces; it has the lowest boiling point of –88.5°C. Methanal is a polar substance and contains dipole–dipole forces as well as dispersion forces; therefore, it has a higher boiling point of –19°C. Methanol contains hydrogen bonds as well as dispersion forces; hydrogen bonds are the strongest of the intermolecular forces and result in methanol having the highest boiling point of 64.7°C.

ℹ When comparing molecular substances with similar molecular masses, the strength of the intermolecular forces is in the order of
dispersion < dipole–dipole
< hydrogen bonding.

When comparing molecules of very different molecular masses, the dispersion forces in the heavier molecules can dominate over the dipole–dipole forces. Table 12.3.3 shows the boiling points of three hydrogen halides: hydrogen chloride (HCl), hydrogen bromide (HBr) and hydrogen iodide (HI). Hydrogen chloride is the most polar molecule and therefore forms the strongest dipole–dipole forces. However, hydrogen iodide has the highest boiling point. This is because it has the largest molecular mass and can form particularly strong dispersion forces, which outweigh the effects of the dipole–dipole forces in hydrogen chloride.

TABLE 12.3.3 Comparison of the boiling points of three polar molecules

Substance	Molecular mass	Number of electrons	Boiling point (°C)
hydrogen chloride (HCl)	36.46	18	–85.1
hydrogen bromide (HBr)	80.91	36	–66.8
hydrogen iodide (HI)	127.91	54	–35.4

The boiling point of hydrogen fluoride (19.5°C) is much higher than any of these other compounds. This is because the hydrogen bonding between hydrogen fluoride molecules is much stronger than both the dispersion forces and the dipole–dipole forces between the other molecules listed in Table 12.3.3.

ℹ When comparing molecular substances of very different molecular masses, the dispersion forces can often outweigh the dipole–dipole forces. However, hydrogen bonds are generally the strongest of the three intermolecular forces.

12.3 Review

SUMMARY

- There are three main types of intermolecular forces: dipole–dipole forces, hydrogen bonds and dispersion forces.
- Dipole–dipole forces are only present between polar molecules and are the result of the attraction between the partially positive and partially negative ends of polar molecules.
- Hydrogen bonding occurs between highly polar molecules in which hydrogen atoms are covalently bonded to an oxygen, a nitrogen or a fluorine atom.
- A hydrogen bond is formed between the hydrogen atom in one molecule and lone pairs of electrons in an oxygen, a nitrogen or a fluorine atom in another molecule.
- Hydrogen bonding is the reason why ice floats in liquid water. As a consequence of the hydrogen bonds between the water molecules, ice forms a crystal lattice in which the molecules are spaced further apart than in liquid water. The greater spacing of molecules makes ice less dense than liquid water and so it floats on water.
- Dispersion forces occur between polar and non-polar molecules.
- Dispersion forces are the result of attractions between temporary dipoles that form in molecules.
- Temporary dipoles are due to random fluctuations in the distribution of electrons in molecules.
- Dispersion forces are stronger between molecules with larger molecular masses because it is easier to create temporary dipoles in molecules with a larger number of electrons.
- In molecules of similar molecular masses, the strength of the intermolecular forces increases in the following order: dispersion forces < dipole–dipole forces < hydrogen bonding.
- In molecules of very different molecular masses, dispersion forces can often outweigh the effects of dipole–dipole forces.
- Hydrogen bonds are generally the strongest of the three types of intermolecular forces.

KEY QUESTIONS

Retrieval

1 Identify the types of intermolecular forces of attraction that exist between:

- **a** polar molecules
- **b** non-polar molecules

Comprehension

2 Use your knowledge of the polarity of molecules and the different types of intermolecular forces to identify which of the following substances would contain dipole–dipole forces: bromine (Br_2), hydrogen chloride (HCl), methane (CH_4), tetrachloromethane (CCl_4), chloromethane (CH_3Cl).

3 Use your knowledge of the features of a molecule required for hydrogen bonding to occur to identify which of the following substances would contain hydrogen bonds.

- **a** ammonia (NH_3)
- **b** chloroform ($CHCl_3$)
- **c** chloromethane (CH_3Cl)
- **d** difluorine oxide (F_2O)
- **e** hydrogen bromide (HBr)
- **f** hydrogen sulfide (H_2S)
- **g** hydrogen fluoride (HF)
- **h** water (H_2O)
- **i** hydrogen (H_2)

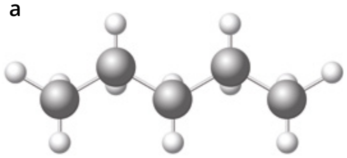
Analysis

4 Predict which substance in each of the following pairs would contain the strongest intermolecular forces in the liquid state and justify your choice.

- **a** F_2 or I_2
- **b** CH_3Cl or CH_3OH
- **c** HCl or F_2

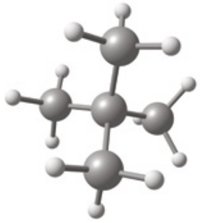
5 Pentane and neopentane, shown in the figure below, have identical molecular masses and only contain dispersion forces. However, their boiling points are quite different. Analyse the structures of both molecules and explain the difference in the boiling points of these compounds.

a



pentane, C_5H_{12} (b.p. 36°C)

b

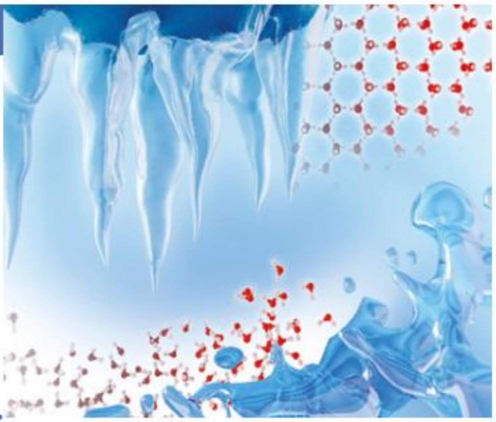


neopentane, C_5H_{12} (b.p. 9.5°C)

12.4 Properties of covalent molecular substances

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe how intermolecular forces influence some of the properties of covalent molecular substances, such as vapour pressure, melting point, boiling point and solubility
- compare the vapour pressures of various covalent molecular substances and explain any differences
- compare the melting and boiling points of various covalent molecular substances and explain any differences
- describe the process of dissolving and explain why some covalent molecular substances are soluble or insoluble in particular solvents.



The physical properties of covalent molecular substances are primarily determined by the strength of the intermolecular forces within each substance. This module examines the relationship between the observable properties of covalent molecular substances and their intermolecular forces. Properties to be examined include vapour pressure, melting point, boiling point and solubility.

VAPOUR PRESSURE

When water is left in an open beaker it will completely evaporate. **Evaporation** occurs when the particles at the surface of the liquid are able to overcome the intermolecular forces of attraction within the liquid and escape. In the case of water, the H_2O molecules are kept together in the liquid state with hydrogen bonds. Some of the molecules at the surface of the liquid have enough energy to overcome these hydrogen bonds and move away from each other and from the surface—they evaporate. Given enough time, all of the water molecules will eventually escape from the surface of the liquid. Figure 12.4.1 shows this process. It is important to note that during evaporation the water molecules themselves remain intact; they move further away from each other as they transition from the **liquid phase** to the **gaseous phase**, but their structure does not change.

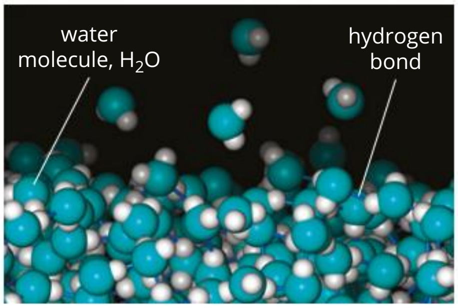


FIGURE 12.4.1 When water evaporates, the hydrogen bonds between the water molecules at the surface of the liquid are broken; however, the water molecules themselves remain intact (their structure does not change).

Now consider what will happen when water is put into a closed container. In this situation, not all of the water will evaporate. There is constant evaporation from the surface of the liquid but because the water particles that escape are moving around in the space above the liquid, some of them will hit the surface of the liquid and re-form hydrogen bonds—they condense. Eventually an **equilibrium** (or balance) will exist in which the rate of evaporation equals the rate of **condensation**; at this point the amount of liquid water and the amount of water vapour above the surface of the liquid will remain constant. This is depicted in Figure 12.4.2.

The water vapour exerts a pressure on the container due to the collisions of the water molecules with the container wall. This pressure is known as the **vapour pressure**.

Liquids with stronger intermolecular forces have lower vapour pressures because the molecules are held together more tightly, making it harder for them to escape from the surface of the liquid. As a consequence, the rate at which molecules leave the surface of the liquid to form the vapour is much lower. Because there are fewer gas particles above the surface of the liquid, the vapour pressure is lower.

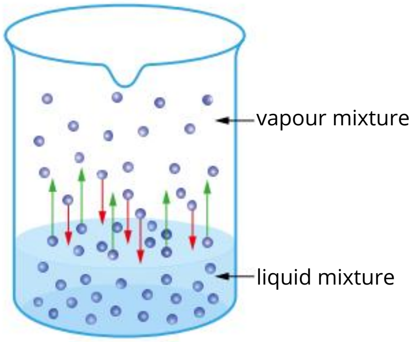


FIGURE 12.4.2 Vapour pressure is the pressure that the gaseous molecules exert on the closed container walls, where the rates of evaporation and condensation are equal.

For example, water has a vapour pressure at room temperature of 3.17 kPa (kPa stands for **kilopascals**, one of many units used for pressure). On the other hand, butane, which is the liquid used as a lighter fluid, has a much higher vapour pressure of approximately 200 kPa. The difference in their vapour pressures can be explained by considering the intermolecular forces present in each of the liquids. As discussed above, water contains hydrogen bonds and, as you saw in Module 12.3, hydrogen bonds are the strongest of the intermolecular forces. Consequently, the water molecules are held together very tightly in the liquid and only a small proportion will be able to escape from the surface of the liquid to form water vapour. The vapour pressure of water is therefore relatively low. Butane (C_4H_{10}) is a non-polar substance so the only intermolecular forces holding the molecules together in the liquid are dispersion forces. These dispersion forces are much weaker than the hydrogen bonds in water and more molecules of butane are able to escape from the surface of the liquid to form a vapour. The vapour pressure of butane is therefore much higher than the vapour pressure of water.

Covalent molecular substances with weaker intermolecular forces have higher vapour pressures.

Worked example 12.4.1

COMPARING THE VAPOUR PRESSURES OF COVALENT MOLECULAR SUBSTANCES

Predict which would have the highest vapour pressure: carbon disulfide (CS_2) or ethanol (C_2H_5OH).

Thinking	**Working**
Determine the intermolecular forces present in each of the substances.	Carbon disulfide is non-polar and only contains dispersion forces. Ethanol is polar and contains hydrogen bonds (due to the presence in each molecule of an oxygen atom attached to a hydrogen atom).
Determine which substance contains the strongest intermolecular forces.	Ethanol contains the strongest intermolecular forces (hydrogen bonds are stronger than dispersion forces).
Determine which substance will have the highest vapour pressure (this will be the substance with the weakest intermolecular forces).	Carbon disulfide will have the highest vapour pressure as it contains the weakest intermolecular forces.

➤ Try yourself 12.4.1

COMPARING THE VAPOUR PRESSURES OF COVALENT MOLECULAR SUBSTANCES

Predict which would have the highest vapour pressure: bromine (Br_2) or carbon disulfide (CS_2).

Vapour pressure and temperature

When a liquid is heated, it shows a greater tendency to evaporate. This is because its molecules have an increased kinetic energy, which allows them to more easily overcome the intermolecular forces holding them together. Therefore, as the temperature increases, the vapour pressure of the liquid also increases as more and more particles escape from the surface of the liquid and move into the gaseous phase. Figure 12.4.3 shows the vapour pressures of four different liquids increasing with temperature. When a liquid's vapour pressure becomes equal to the atmospheric pressure of the surroundings, boiling occurs. At this point, bubbles of vapour are observed forming throughout the liquid. At sea level, atmospheric pressure is 101.325 kPa (or 1 atm). At this pressure, the boiling point is known as the normal boiling point. The normal boiling points of the four liquids in Figure 12.4.3 can be determined by identifying the temperature at which the vapour pressure is equal to 101.325 kPa.

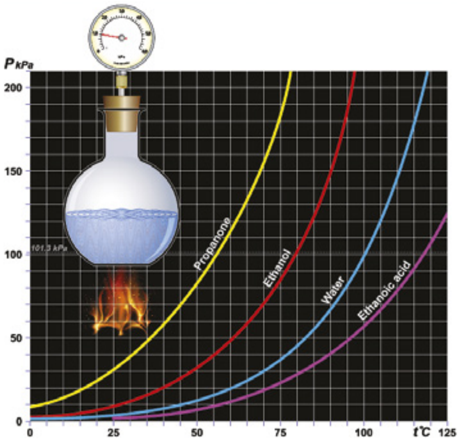
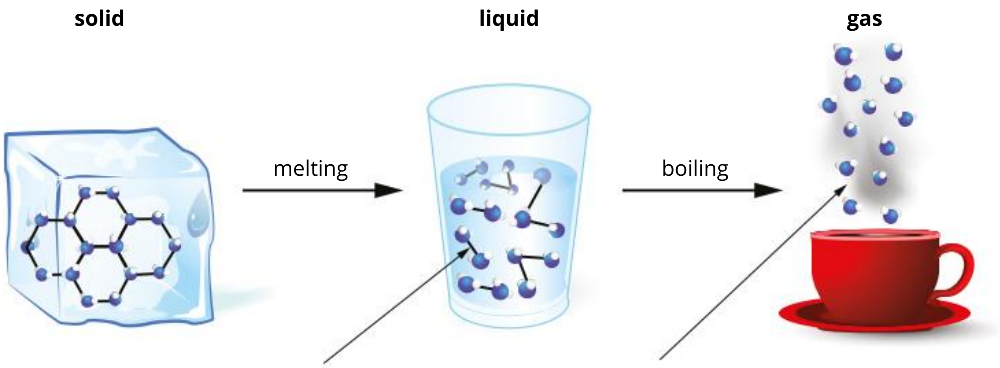


FIGURE 12.4.3 As temperature increases, the vapour pressure also increases. The boiling points of liquids can be determined by identifying the temperature at which the vapour pressure is equal to the atmospheric pressure. Here, at an atmospheric pressure of 101.325 kPa, ethanol has a boiling point of about 80°C and water about 100°C.

MELTING AND BOILING POINTS

Covalent network solids (such as diamond, shown in Figure 12.3.3, page 362) and ionic compounds (such as sodium chloride) have very high melting and boiling points, which range from several hundred degrees Celsius to several thousand degrees Celsius. This is because they contain strong intramolecular covalent or ionic bonds that are broken when these substances are melted or boiled. Breaking these intramolecular bonds requires large amounts of energy.

However, when covalent molecular substances are melted or boiled, only intermolecular forces are broken; the intramolecular covalent bonds holding the atoms together within each molecule remain intact. Intermolecular forces are much weaker than strong intramolecular bonds and covalent molecular substances have much lower melting and boiling points as a result. Figure 12.4.4 depicts what happens to water molecules as water changes state from a solid to a liquid and then to a gas. As you can see, it is the hydrogen bonds between the water molecules that are broken and not the covalent O–H bonds within the molecules.



Some hydrogen bonds are broken when water changes from a solid to a liquid.

The remaining hydrogen bonds are broken when water changes from a liquid to a gas.

FIGURE 12.4.4 When water changes from a solid to a liquid and then from a liquid to a gas, it is the hydrogen bonds between the water molecules that are broken and not the O–H covalent bonds within the molecules.

In Module 12.3, you looked at the different types of intermolecular forces and how their strengths affect the melting and boiling points of covalent molecular substances. In this part of Module 12.4, you will revise the main points from Module 12.3 and consolidate your understanding with specific examples.

In general, the strength of intermolecular forces increases as shown in Figure 12.4.5. Hydrogen bonding is the strongest intermolecular force, whereas dispersion forces are the weakest. The melting and boiling points of covalent molecular substances often depend on the different types of intermolecular forces present in each substance.

dispersion forces < dipole–dipole forces < hydrogen bonding

→

increasing strength of intermolecular forces

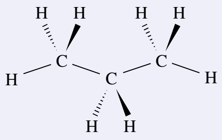
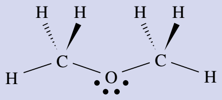
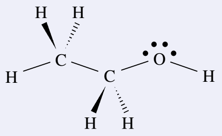
FIGURE 12.4.5 Increasing strength of the different types of intermolecular forces

However, you will recall that in the liquid and solid states, dispersion forces exist between all covalent molecular substances and that the strength of the dispersion forces tends to increase with an increase in molecular mass. Consequently, in substances that have large molecular masses, the contribution by the dispersion forces can be quite significant and often outweighs the dipole–dipole forces. The following examples illustrate the above points.

Comparing covalent molecular substances of similar molecular masses

Consider the following covalent molecular substances: propane ($CH_3CH_2CH_3$), methoxymethane (CH_3OCH_3) and ethanol (CH_3CH_2OH). They all have similar molecular masses, which means that, in general, the strength of their dispersion forces will also be similar. However, as a result of different intermolecular forces being present in each substance, the melting points for the three vary significantly as do the boiling points. This is shown in Table 12.4.1. As you can see from the shapes of the molecules in Table 12.4.1, propane is non-polar so the only intermolecular forces present in the substance are dispersion forces. Propane therefore has the lowest melting and boiling points. Methoxymethane is polar, which means that in addition to dispersion forces, the substance also contains dipole–dipole forces. These dipole–dipole forces are stronger than the dispersion forces and results in methoxymethane having higher melting and boiling points than propane. Ethanol molecules are polar and have a highly electronegative oxygen atom attached to a hydrogen atom, which means that the molecules can form hydrogen bonds. Because hydrogen bonding is the strongest intermolecular force, ethanol has the highest melting and boiling points of the three substances.

TABLE 12.4.1 Comparison of the melting and boiling points of covalent molecular substances with similar molecular masses

Substance	Molecular shape	Polarity	Relative molecular mass (amu)	Melting point (°C)	Boiling point (°C)
propane		non-polar	44.11	–188	–42
methoxymethane		polar	46.08	–142	–25
ethanol		polar	46.08	–117	78

Worked example 12.4.2

COMPARING THE MELTING AND BOILING POINTS OF MOLECULAR SUBSTANCES WITH SIMILAR MOLECULAR MASSES

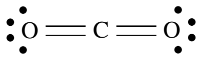
Predict which would have the highest melting and boiling points: CH_4 or NH_3.

Thinking	**Working**
Determine which intermolecular forces are present in each substance.	CH_4 is non-polar so only dispersion forces will be present. NH_3 is polar and can form hydrogen bonds because it contains a highly electronegative nitrogen atom covalently bonded to a hydrogen atom.
Determine which type of intermolecular force is the strongest.	Hydrogen bonding is generally stronger than dispersion forces.
Determine which substance has the highest melting and boiling point based on the intermolecular forces present.	NH_3 will have the highest melting and boiling point because of its stronger intermolecular forces.

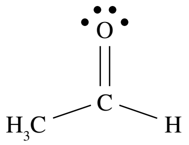
➤ Try yourself 12.4.2

COMPARING THE MELTING AND BOILING POINTS OF MOLECULAR SUBSTANCES WITH SIMILAR MOLECULAR MASSES

Predict which would have the highest melting and boiling points: CO_2 or CH_3CHO.



CO_2



CH_3CHO

Comparing covalent molecular substances that have very different molecular masses

Table 12.4.2 shows the melting and boiling points of some of the group 16 hydrides. Ignoring water for the moment and just focusing on the remaining hydrides, you can see that as the molecular mass increases, the melting and boiling points also increase (even though the polarities of the molecules, and consequently the dipole–dipole forces, decrease). This is due to the dispersion forces becoming stronger and being the dominant force between the heavier molecules.

Water (H_2O) has the smallest molecular mass of the group 16 hydrides and therefore the weakest dispersion forces. However, it melts at 0°C and boils at 100°C, which is considerably higher than the melting and boiling points of any of the other group 16 hydrides. This marked deviation from the general trend is attributed to the hydrogen bonding that exists between water molecules in the solid and liquid states. Hydrogen bonding is the strongest intermolecular force and can often outweigh very large dispersion forces; it is the reason why many substances containing hydrogen bonds between their molecules have particularly high melting and boiling points.

TABLE 12.4.2 Comparison of the melting and boiling points of some of the group 16 hydrides

Hydride	Molecular mass (amu)	Melting point (°C)	Boiling point (°C)
H_2O	18.02	0	100
H_2S	34.08	–82	–60.7
H_2Se	80.99	–66	–41.5
H_2Te	129.62	–49	–2.2

ℹ Hydrogen bonding is the strongest intermolecular force and can often outweigh even large dispersion forces.

Worked example 12.4.3

COMPARING THE MELTING AND BOILING POINTS OF MOLECULAR SUBSTANCES WITH VERY DIFFERENT MOLECULAR MASSES

Predict which would have the highest melting and boiling points: CH_4 or CF_4.

Thinking	**Working**
Compare the molecular masses of the molecules.	CF_4 is a much heavier molecule than CH_4.
Determine the intermolecular forces present in each substance.	Both are non-polar so dispersion forces are the only intermolecular forces present.
Determine which substance has the highest melting and boiling point based on the strength of the intermolecular forces.	The dispersion forces will be stronger in CF_4 than CH_4 because CF_4 has the largest molecular mass (and is therefore the larger molecule with more electrons; this makes it more polarisable). CF_4 has the highest melting and boiling points.

➤ Try yourself 12.4.3

COMPARING THE MELTING AND BOILING POINTS OF MOLECULAR SUBSTANCES WITH VERY DIFFERENT MOLECULAR MASSES

Predict which would have the highest melting and boiling points: O_2 or S_8.

SOLUBILITY

The intermolecular forces present in a covalent molecular substance have a significant impact on its **solubility**. In general, a **solute** will dissolve in a **solvent** that has similar intermolecular forces. You can remember this with the saying 'like dissolves like'.

Pentane (C_5H_{12}), for example, readily dissolves in octane (C_8H_{18}). Both are non-polar liquids and only contain dispersion forces; they mix easily because of their similar intermolecular forces. During the dissolving process, energy is required to break the dispersion forces in the pure liquids and separate the molecules. New dispersion forces then form between the pentane and octane molecules, as shown in Figure 12.4.6. The energy that is required to break the pentane–pentane and octane–octane dispersion forces in the pure liquids is compensated for by the energy that is released with the formation of the new pentane–octane dispersion forces.

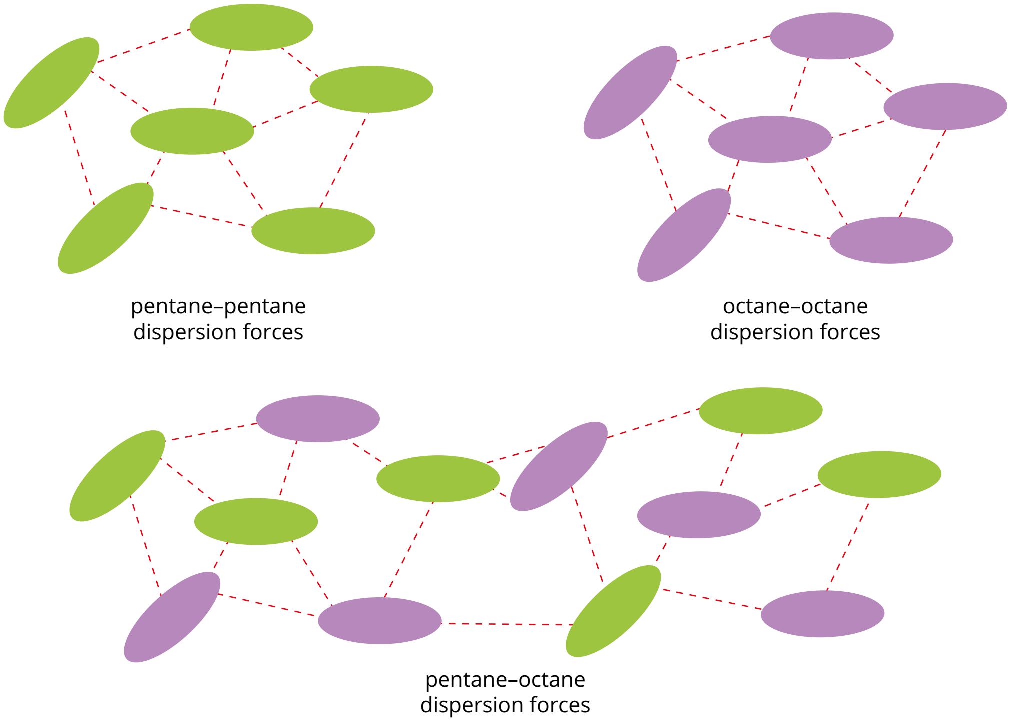


FIGURE 12.4.6 Pentane dissolves in octane. Both molecules are non-polar and are attracted to each other by dispersion forces.

While non-polar solutes dissolve in non-polar solvents, they will not dissolve in polar solvents such as water. This is because the intermolecular forces in the pure liquids are very different. For example, if pentane were to dissolve in water, because it is non-polar it could only form dispersion forces with the water molecules. The energy released on the formation of these dispersion forces would not be enough to compensate for the much larger amount of energy that would be required to break the hydrogen bonds between the water molecules in the first place. As a result, when pentane is added to water, two separate layers are visible with the pentane floating on top of the water; this is similar to what is observed when oil is added to water (oil is also non-polar and does not dissolve in water).

Solutes will generally only dissolve in solvents if their intermolecular forces are similar.

Lava lamps, such as the one shown in Figure 12.4.7, make use of the fact that non-polar substances are insoluble in water. They are made from a mixture of water and several different non-polar substances, such as oil and wax. When the lamps are turned on, the movement of the different layers throughout the mixture is very attractive.

Polar solutes dissolve in polar solvents because of their similar intermolecular forces. For example, methanol (CH_3OH) is very soluble in water because both substances contain hydrogen bonds. While the energy required to break the hydrogen bonds in the pure liquids is relatively large, the release of energy when new hydrogen bonds form between the methanol and water molecules compensates for this. Figure 12.4.8 shows the hydrogen bonds that form when methanol dissolves in water.

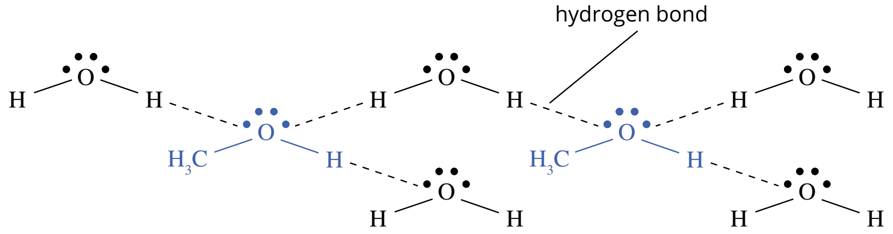


FIGURE 12.4.8 Hydrogen bonding between methanol and water molecules

FIGURE 12.4.7 Lava lamps make use of the fact that non-polar substances do not dissolve in polar solvents like water.

Worked example 12.4.4

COMPARING THE SOLUBILITIES OF COVALENT MOLECULAR SUBSTANCES

Predict whether the following substances (the solutes) would be soluble in hexane (a non-polar solvent) or water.

a bromine (Br_2)

b urea ($CO(NH_3)_2$)

Thinking	**Working**
Determine the intermolecular forces present in each of the solutes.	Bromine is non-polar so only contains dispersion forces. Urea is polar and contains hydrogen bonds (due to the presence of electronegative nitrogen atoms attached to hydrogen atoms).
Determine the intermolecular forces present in each of the solvents.	Hexane is non-polar so only contains dispersion forces. Water is polar and contains hydrogen bonds.
A solute will dissolve in a solvent that has similar intermolecular forces. Therefore, pair up each solute with the solvent that closely matches its intermolecular forces.	Bromine will be soluble in hexane, whereas urea will be soluble in water.

➤ Try yourself 12.4.4

COMPARING THE SOLUBILITIES OF COVALENT MOLECULAR SUBSTANCES

Predict which would have the highest solubility in hexane: iodine (I_2) or hydrogen iodide (HI).

12.4 Review

SUMMARY

- Many of the physical properties of covalent molecular substances are determined by intermolecular forces. These properties include vapour pressure, melting point, boiling point and solubility.
- When covalent molecular substances change state, it is the intermolecular forces that are broken and not the intramolecular covalent bonds.
- Intermolecular forces are much weaker than strong intramolecular bonding such as ionic and covalent bonds.
- Vapour pressure is the pressure exerted by a gas above the surface of its liquid.
- Liquids with stronger intermolecular forces have lower vapour pressures.
- A liquid's vapour pressure increases with temperature because the increased kinetic energy of the molecules enables them to overcome the intermolecular forces holding them together.
- Boiling occurs when a liquid's vapour pressure becomes equal to the atmospheric pressure of the surroundings.
- Melting and boiling points of covalent molecular substances are significantly lower than those of covalent network substances and ionic compounds.
- As the strength of the intermolecular forces increases, melting and boiling points also increase.
- Intermolecular forces present in a molecular substance have a significant impact on its solubility.
- Solutes will dissolve in solvents that have similar intermolecular forces.

KEY QUESTIONS

Retrieval

1 **a** Define the normal boiling point of a liquid.
  b State the normal boiling point of water.

2 Identify what attractive forces are broken when the following substances are melted or boiled.
  a covalent molecular substances
  b covalent network solids
  c ionic substances

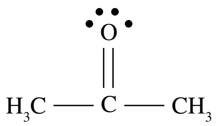
3 Indicate what is meant by the saying 'like dissolves like' when referring to a solute dissolving in a solvent.

Comprehension

4 The following statements refer to the vapour pressures of covalent molecular substances. Use your understanding of intermolecular forces to explain each statement.
  a A liquid with strong intermolecular forces has a low vapour pressure.
  b An increase in temperature leads to an increase in a liquid's vapour pressure.

5 **a** Describe the changes that occur at the molecular level when a covalent molecular liquid is heated to boiling point.
  b Explain why covalent molecular substances have lower melting and boiling points than covalent network solids and ionic compounds.

6 Use your understanding of how intermolecular forces affect solubility to determine whether each of the following covalent molecular substances would be soluble in a polar solvent, such as water, or a non-polar solvent, such as hexane.
  a methane (CH_4)
  b ethanol (C_2H_5OH)
  c bromine (Br_2)
  d propanone (CH_3COCH_3) (the structure of propanone is shown below)



Analysis

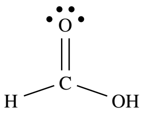
7 Pentane (C_5H_{12}) and octane (C_8H_{18}) are both non-polar and liquids at room temperature. Predict which will have the highest vapour pressure at this temperature and explain your answer.

8 Bromine (Br_2) and iodine monochloride (ICl) are both liquids at room temperature. Predict which will have the highest vapour pressure at this temperature and explain your answer.

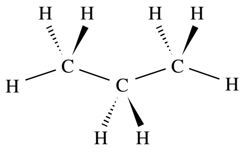
9 For each pair of covalent molecular substances, analyse their structures and determine the reason for the differences in their boiling points.

- **a** F_2, boiling point = –188°C; HCl, boiling point = –85°C
- **b** PH_3, boiling point = –88°C; NH_3, boiling point = –33°C
- **c** CH_4, boiling point = –161°C; CCl_4, boiling point = 77°C

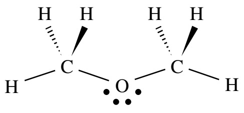
10 **a** Categorise the following covalent molecular substances in order of increasing boiling point.



methanoic acid



propane



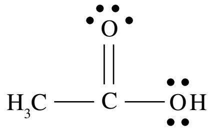
methoxymethane

b Explain the order in part **a**.

11 Determine why pentane (C_5H_{12}) is soluble in carbon tetrachloride (CCl_4), whereas ethanol (C_2H_5OH) is insoluble in carbon tetrachloride.

12 **a** Predict which one of the following covalent molecular substances is soluble in water: ethanoic acid (shown below), bromine (Br_2) or carbon tetrachloride (CCl_4).

b Explain your choice.

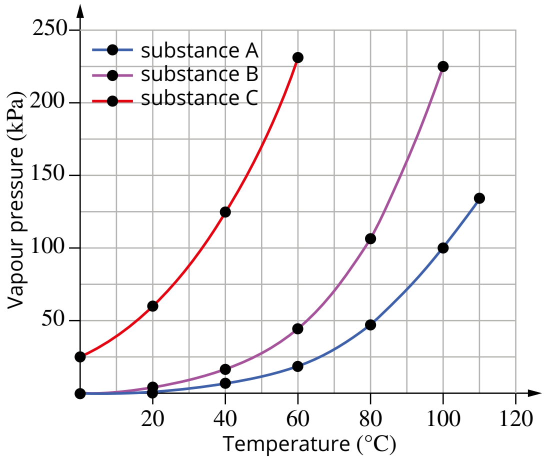


$\mathrm{H_3C{-}C({=}O){-}OH}$

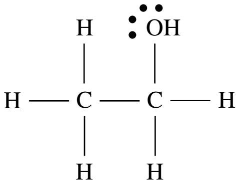
ethanoic acid

13 The figure below shows the vapour pressure curves for three liquids that are covalent molecular substances. Use the graph to answer the questions that follow.

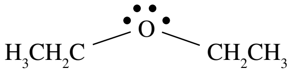
Vapour pressure curves for three liquids



- **a** Conclude which substance (A, B or C) contains the strongest intermolecular forces.
- **b** Conclude which substance (A, B or C) contains the weakest intermolecular forces.
- **c** Knowing that the three substances are ethanol, water and diethyl ether (the structures of ethanol and diethyl ether are shown) and that water is substance A, determine the identities of substances B and C. Explain your answer.



ethanol



diethyl ether

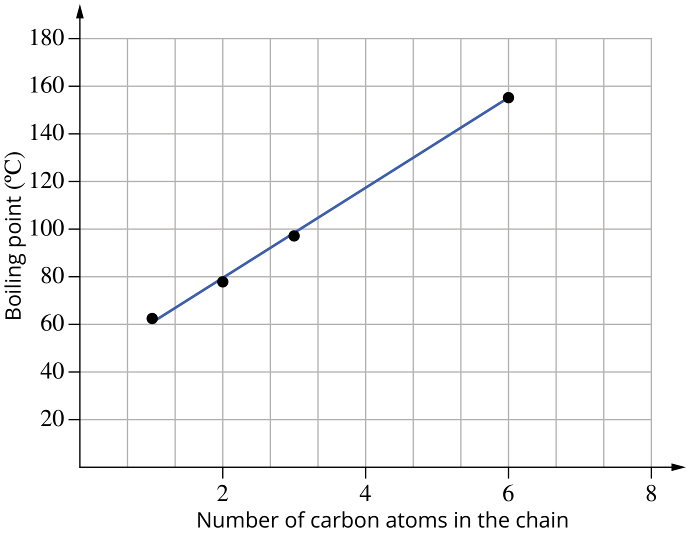
- **d** Deduce from their structures, why there is a difference in vapour pressures between substance A and substance B.

14 The boiling points of a series of organic alcohols are given in the table below and shown on the graph following. You will notice that as you move down the series, the chain length of the alcohols increases by one carbon each time.

- **a** Use the graph to predict the boiling point of the alcohols butan-1-ol and pentan-1-ol.
- **b** Determine which type of intermolecular force is responsible for the increase in boiling points within this series of alcohols.
- **c** Explain why the boiling points increase as the carbon chain length increases.

Alcohol	**Carbon chain length**	**Boiling point (°C)**
methanol (CH_3OH)	1	64.5
ethanol (CH_3CH_2OH)	2	78.5
propanol ($CH_3CH_2CH_2OH$)	3	97.2
butanol ($CH_3CH_2CH_2CH_2OH$)	4	?
petan-1-ol ($CH_3CH_2CH_2CH_2CH_2OH$)	5	?
hexan-1-ol ($CH_3CH_2CH_2CH_2CH_2CH_2OH$)	6	157

Relationship between carbon chain length and boiling point in a series of organic alcohols



MANDATORY PRACTICAL 3

Making molecular models

Research and planning

Aim

To demonstrate the bonding and shape of a number of simple covalent molecules

Rationale (scientific background to the experiment)

Making three-dimensional representations of covalent molecules provides a physical demonstration of the bonding and structure that is within a molecule. Models enable students to understand the concepts of molecular shape, symmetry, bond angles and polarity more clearly as they can rotate the structure and view it from different perspectives.

However, there are limitations to what these models can show, and, at all times while building these molecular models, students should be questioning what the differences are between the model and what they know of the actual molecule they are trying to represent.

Timing

60 minutes

Materials

- commercially produced molecular model-building kit or lumps of different colours of plasticine and different length straws
- If you are using straws or sticks to show bonds rather than a modelling kit, the lengths listed below should be used to represent the bonds of different types.

Bond type	Length of straw or stick (cm)
single covalent	5.0
double covalent	3.5
triple covalent	2.0
dispersion forces and hydrogen bonds	6.0

Method

Risk assessment

Consideration of risks includes chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and Assessment book or download it from your eBook.

1 Before building models of the molecules listed in Table 12.5.1, complete column two (Valence structure of the molecule to be built) and three (Lewis structure of the molecule to be built) of Table 12.5.1.

2 Using the appropriately coloured ‘atoms’, construct a model of each molecule listed in Table 12.5.1.

3 Draw each structure with a neat diagram that shows a three-dimensional representation of your model.

4 Repeat steps 1–3 for Table 12.5.2 then Table 12.5.3.

Analysing

Raw data

1 Record your results in Table 12.5.1, Table 12.5.2 and Table 12.5.3.

TABLE 12.5.1 Single covalent bonds

Molecule	Valence structure	Lewis structure	3D structure
methane (CH_4)			
ammonia (NH_3)			
water (H_2O)			
hydrogen chloride (HCl)			
phosphine (PH_3)			
butane (C_4H_{10})			

TABLE 12.5.2 Double covalent bonds

Molecule	Valence structure	Lewis structure	3D structure
oxygen (O_2)			
carbon dioxide (CO_2)			
ethene (C_2H_4)			
methanal (CH_2O)			

TABLE 12.5.3 Triple covalent bonds

Molecule	Valence structure	Lewis structure	3D structure
nitrogen (N_2)			
ethyne (C_2H_2)			

Processed data

2 Consider each molecular model. Complete Table 12.5.4 by identifying each molecule's shape, whether it is polar or non-polar and the type/s of intermolecular force you would expect to exist between molecules for each of these substances.

TABLE 12.5.4 Molecule shape, polarity and intermolecular force

Molecule	Shape	Polarity	Intermolecular force/s
methane (CH_4)			
ammonia (NH_3)			
water (H_2O)			
hydrogen chloride (HCl)			
phosphine (PH_3)			
butane (C_4H_{10})			
oxygen (O_2)			
carbon dioxide (CO_2)			
ethene (C_2H_4)			
methanal (CH_2O)			
nitrogen (N_2)			
ethyne (C_2H_2)			

Reflect and check that your data analysis demonstrates these characteristics

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.
☐ Accurate application of algorithms, visual and graphical representations of data is demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

Analysis

3 Name the polar molecules that you have built and explain why they are polar.

4 State at least two limitations of the models you have built in how well they account for the properties of the compounds named in Question **3**.

5 Name one or more elements that can form stable single, double and triple bonds with its own atoms.

6 Name two elements whose atoms never form double bonds.

7 Explain why the bonding pairs in CH_4 are arranged in a tetrahedral shape.

Reflect and check that your analysis demonstrates these characteristics

☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.
☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.
☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Interpreting and communicating

Conclusion

1 **a** State your conclusion.
  b Describe the evidence you have collected that leads you to this conclusion.

Evaluation

2 Considering your analysis and conclusion, evaluate whether the experiment provided an effective and efficient method of modelling the bonding and shape of various simple covalent molecules.

Improvements

3 If you were to repeat the experiment, identify the steps that you would do differently. You should include these points your answer.
  a Explain how you would change the methodology and how this might improve the results.
  b Consider how well you performed the tasks and the skills that you need to improve on in your technique.
  c Explain how the collection of data could be improved.

Reflect and check that your evaluation demonstrates these characteristics

☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis).
☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respect to the error analysis and limitations or sufficiency of the data.
☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment, which are logically derived from the analysis of the evidence.

Chapter review

12

KEY TERMS

asymmetrical molecule
bond angle
bonding electrons
bonding pairs
condensation
covalent network solid
dipole
dipole–dipole forces
dispersion forces
electric dipole
electron arrangement
electron density
electron distribution
electronegativity
electrostatic force
equilibrium
evaporation
gaseous phase
hydrogen bonding
instantaneous dipole
kilopascal (kPa)
Lewis structure
liquid phase
lone pair
non-polar
octet rule
polar
polarisable
polarised
polarity
polyatomic molecule
pyramidal
solubility
solute
solvent
surface tension
symmetrical molecule
temporary dipole
tetrahedral
trigonal planar
valence electrons
valence shell electron pair repulsion (VSEPR) theory
vapour pressure
wedge-dash notation

KEY QUESTIONS

Retrieval

1 Identify the correct bond angle between atoms in a molecule that adopts a tetrahedral shape.

- **A** 180°
- **B** 109°
- **C** 104.5°
- **D** 120°
- **E** 90°

2 Select the correct definition for the electronegativity of an atom.

- **A** the amount of energy required for an atom to lose an electron
- **B** the ability of an atom to form a hydrogen bond
- **C** the ability of an atom to form an ion
- **D** the ability of an atom to attract electrons to itself

3 Indicate the correct word from the options below to complete the following sentence: Water molecules are predominantly held together in the liquid state by:

- **A** covalent bonds
- **B** hydrogen bonds
- **C** dispersion forces
- **D** dipole–dipole forces

4 Identify the type of intermolecular force that is generally the strongest.

- **A** covalent bonds
- **B** hydrogen bonds
- **C** dispersion forces
- **D** dipole–dipole forces

5 Identify whether each of the following statements about covalent molecular substances is true or false.

- **a** The vapour pressure of a substance is the pressure exerted by a gas in equilibrium with its liquid in a closed container.
- **b** When a substance changes state from a solid to a liquid, the covalent bonds within the molecules are broken.
- **c** A solute will only dissolve in a solvent if their intermolecular forces are very different.
- **d** When a substance changes state from a liquid to a gas, its intermolecular forces of attraction are completely broken.

6 State the three main types of intermolecular forces.

7 For hydrogen bonding to occur between molecules, state what group of atoms each molecule must contain.

Comprehension

8 Use your knowledge of VSEPR theory to determine the shape of a molecule that contains a central atom surrounded by four bonding groups of electrons and no lone pairs.

- **A** linear
- **B** trigonal planar
- **C** tetrahedral
- **D** pyramidal
- **E** bent

9 Use your knowledge of VSEPR theory to determine the shape of a molecule that contains a central atom surrounded by three bonding groups of electrons and one lone pair.

- **A** linear
- **B** trigonal planar
- **C** tetrahedral
- **D** pyramidal
- **E** bent

10 Use the electronegativity values in Figure 12.2.1 on page 357 to calculate the electronegativity difference between the atoms in each of the following covalent bonds and determine which bond is the most polar.

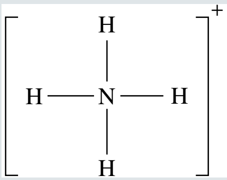
N–N, C–H, H–F, O–F, F–F

11 Calculate the electronegativity differences between the atoms and determine which of the following molecules contains the most polar bond.

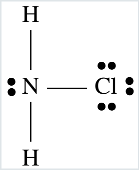
CO_2, H_2O, H_2, H_2S, NH_3

12 All of the following molecules have four pairs of electrons around the central atom. For each one, use your understanding of VSEPR theory to determine whether the molecular shape is tetrahedral, pyramidal or bent.

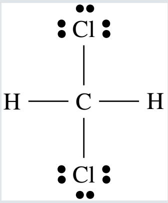
a



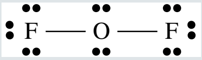
b



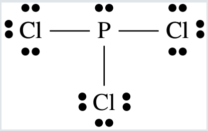
c



d



e



13 Use your understanding of electronegativity and polarity of molecules to determine which of the molecules in Question **12** are polar.

14 In the liquid state, explain what intermolecular forces of attraction exist between molecules of hydrogen chloride (HCl). Draw a representation of these intermolecular forces.

15 Use your understanding of intermolecular forces to identify the type of intermolecular force which must be overcome in each of the following processes.

- **a** the melting of ice
- **b** the conversion of solid iodine to gaseous iodine
- **c** dissolving ethanoic acid (CH_3COOH) in water to make vinegar

16 The following substances increase in boiling point in the order shown: $CH_3F < CH_3Cl < CH_3Br < CH_3I$. Explain this order.

17 Diethyl ether and ethanol are both liquids at room temperature. Explain why ethanol has a lower vapour pressure than diethyl ether at this temperature.

diethyl ether

ethanol

Analysis

18 Draw the Lewis diagrams of the following molecules and apply your understanding of VSEPR theory to determine their molecular shape.

- **a** PF_3
- **b** O_3
- **c** CO_3^{2-}
- **d** NH_3

19 Determine the bond angles in each of the following molecules using your understanding of VSEPR theory.

- **a** CF_4
- **b** CO_2
- **c** H_2O
- **d** NH_3

20 Predict the shape of each of the following molecules.

- **a** CS_2
- **b** H_2S
- **c** SiH_4
- **d** H_3O^+

21 Ozone (O_3) and water (H_2O) both have a central oxygen atom bonded to two atoms and a bent molecular shape. However, the bond angles in ozone are larger than those in water. Compare the structures of both molecules and determine the reason for the difference in these bond angles.

22 CBr_4 and NO_3^- are both non-polar molecules even though they contain polar bonds. Explore why this is so using your understanding of molecular shape, bond polarities and net dipoles.

23 Consider the polarity of the bonds and the shapes of the following molecules and predict which of the following oxides of carbon are polar. Explain your choice/choices.

- **a** CO
- **b** CO_2
- **c** CO_3^{2-}

24 Consider the polarity of the bonds and the shapes of the following molecules and predict which molecules are polar.

- **a** CH_2Cl_2
- **b** NF_3
- **c** Cl_2
- **d** O_3
- **e** PCl_4^+
- **f** HCl
- **g** Cl_2

25 Predict which molecule would be the most polar: NH_3 or PH_3. Explain your choice.

26 Consider the two compounds CH_4 and CCl_4. CH_4 is a gas at room temperature, whereas CCl_4 is a liquid.

- **a** Determine which substance has the stronger intermolecular forces of attraction.

b Compare the structure of the molecules in each substance and use this comparison to explain the difference in the strengths of the intermolecular forces.

27 Classify the following compounds in order of increasing strength of the intermolecular forces of attraction present in each: F_2, HF, HCl. Explain your order.

28 Consider the two compounds OF_2 and CF_4. OF_2 has a boiling point of –145°C and CF_4 has a boiling point of –128°C. From their boiling points, determine which compound contains the strongest intermolecular forces of attraction. Compare the structure of the molecules in each compound and discuss why their boiling points differ.

29 Hydrogen bonding occurs in both NH_3 and H_2O. Predict the substance in which the hydrogen bonding would be stronger. Explain your answer and use a diagram to illustrate your point.

30 Fluorine (F_2) is a gas at room temperature, whereas iodine (I_2) is a solid. Identify which factor best explains this difference using your understanding of intermolecular forces.

31 Predict which of the following substances would have the highest boiling point: CO_2 or NO_2. Explain your choice.

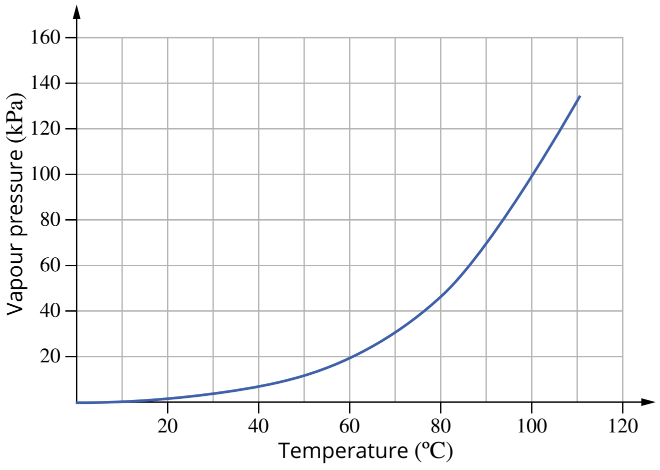
32 Ammonia (NH_3) boils at –33°C, whereas hydrazine ($H_2N–NH_2$) boils at 114°C. Explore a possible reason for the very large difference in boiling points.

33 Predict which substance would have the lower vapour pressure at a given temperature: CH_2F_2 or CH_4. Explain your choice.

34 Predict which gas would have the highest solubility in water: N_2, Cl_2, CO_2, NH_3.

35 The vapour pressure curve for water is shown in the graph below. Use this graph to answer the questions that follow.

Vapour pressure curve for water



a Determine why the vapour pressure of water increases with temperature.

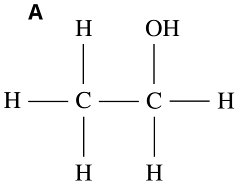
b State the normal boiling point of water.

c Predict the boiling point of water on the summit of Everest where the atmospheric pressure is only about 35 kPa.

d Explore why the boiling point of water is so much lower on the summit of Everest than at sea level.

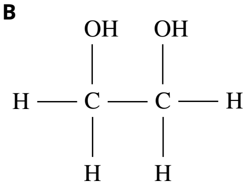
36 A student recorded the boiling points of four different covalent molecular substances but forgot to write down the substance that corresponded with each boiling point. The student tested ethanol, ethane-1,2-diol, methanol and propan-1-ol (Lewis diagrams shown below). Analyse the molecular structure of each substance and use your understanding of intermolecular forces to match each substance with its correct boiling point.

A



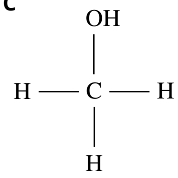
ethanol

B



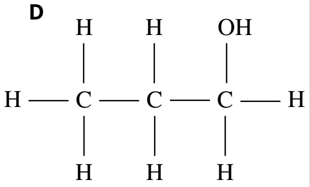
ethane-1, 2-diol

C



methanol

D



propan-1-ol

Substance	Boiling point (°C)
?	64.5
?	78.5
?	97.2
?	198

37 Analyse the data in the table below and explore the evidence that allows you to conclude that methylamine, CH_3NH_2, has stronger intermolecular forces of attraction than chloromethane, CH_3Cl.

Substance	Molecular formula	Melting point	Boiling point	Solubility in water
methylamine	CH_3NH_2	–92.5	–6.3	very soluble
chloromethane	CH_3Cl	–97.7	–23.8	sparingly soluble

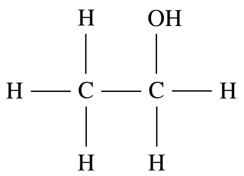
Knowledge utilisation

38 The molecule H_2O_2 has a net dipole.

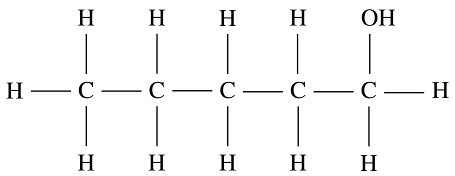
- **a** Determine whether it is possible for the shape of the molecule to be linear. Discuss.
- **b** Predict a possible structure for the molecule.

39 VSEPR theory recognises the significant effect that lone pairs of electrons have on the shapes of molecules such as water. If the lone pairs of electrons *were not* taken into account, predict what might be expected for the shape of a water molecule. Consider this predicted shape for a water molecule again: discuss how the observed properties of water would confirm that this shape is incorrect.

40 Consider the structures of ethanol and pentan-1-ol as shown below. Both substances are polar and contain hydrogen bonds in the liquid state. Ethanol is highly soluble in water (which you would expect for a substance that contains hydrogen bonds), whereas pentan-1-ol is only very slightly soluble. Explore a possible reason for this difference.



ethanol



pentan-1-ol

41 When water freezes at 0°C, the ice it forms can float on the liquid water. However, when benzene (C_6H_6) freezes at 5.5°C, the frozen benzene cannot float on the liquid benzene; instead, it sinks to the bottom. Propose a possible reason for this difference.

42 Consider the molecular models that were built for the Mandatory Practical. Examine the ways in which bonding and molecular shape determine the polarity of a molecule. Use diagrams of the models and show the net dipole to help answer this question.

CHAPTER 13

Chromatography

The term ‘chromatography’ refers to a set of techniques that can be used to separate and analyse the components in a mixture. Separation is performed by passing the mixture through a medium that allows the components of the mixture to move through at different rates. By choosing the medium and managing the conditions of the separation chemists can qualitatively and quantitatively answer the question ‘What is in this sample?’

You will see how simple laboratory experiments using chromatography paper, thin-layer chromatography and column chromatography demonstrate key principles that are also used in complex automated separations used by chemists in industry, research and forensics. Extremely sensitive and rapid analysis of the components in a mixture is performed in the gas phase (gas chromatography, GC) or in the liquid phase (high-performance liquid chromatography, HPLC).

In each of the techniques described in this chapter you will see how interactions between the moving sample and the chromatography medium allow separation of individual components based on their physical properties. You will then examine how these basic principles are used in GC and HPLC for identifying compounds (qualitative identification) and determining their concentration (quantitative identification).

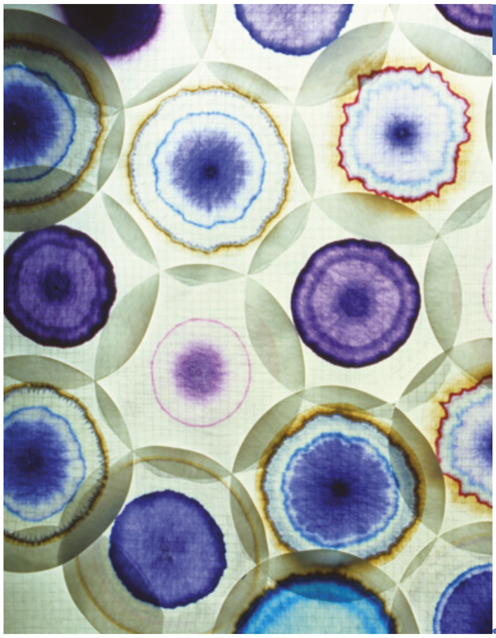
Syllabus subject matter

Topic 1 • Intermolecular forces and gases

CHROMATOGRAPHY TECHNIQUES

- recognise that chromatography techniques, including paper, thin layer, gas and high-performance liquid chromatography, can be used to determine the composition and purity of substances
- describe and explain how variations in the strength of the interactions between atoms, molecules or ions in the mobile and stationary phases can be used to separate components
- analyse, interpret and evaluate data from chromatographs to determine the composition and purity of substances, including calculating R_f values.

13.1 Principles of chromatography



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that chromatography refers to a range of techniques, using common principles, that are used to separate and identify the components in a mixture
- recognise simple forms of this technique such as paper chromatography, thin-layer chromatography and simple column chromatography
- understand that separation of components of a mixture is due to differences in affinity for the mobile and stationary phases
- understand that components separate during chromatography because of how strongly they adsorb to the stationary phase and desorb back into the mobile phase
- understand that in paper and thin-layer chromatography, the components in a mixture can be qualitatively identified by comparison with known standards and determination of R_f values
- recognise that simple column chromatography techniques are the basis for more sensitive automated techniques.

Chromatography is a technique used to separate the components in a mixture. It can be used for both **qualitative analysis** and **quantitative analysis** to identify numerous inorganic and organic substances such as contaminants in water, toxic gases in air, impurities in food and drugs present in blood. A feature of chromatography is that it can be performed on very small samples and with solutions of low concentrations. When separated, the various components often appear as distinct bands or as a **chromatogram**, as seen in Figure 13.1.1.

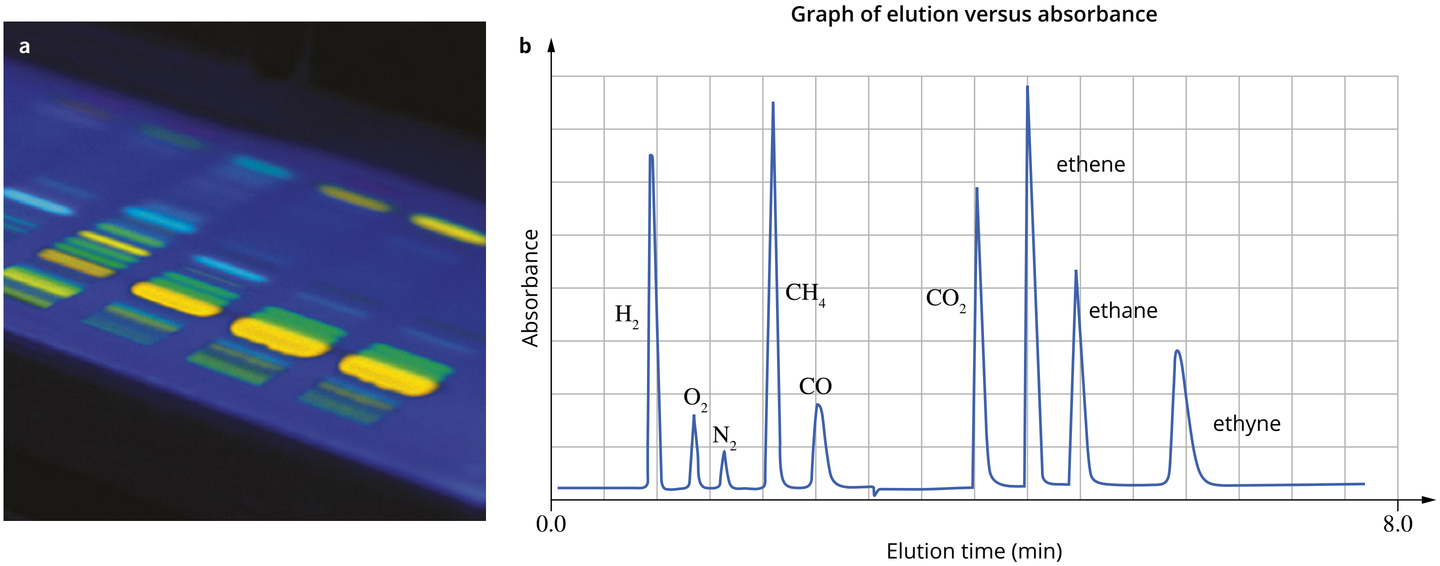


FIGURE 13.1.1 (a) Bands viewed under fluorescent lighting are individual chemical components of a mixture that has been separated by chromatography. (b) Peaks registered by gas chromatography indicate dissolved gases in insulating oil in a transformer. High gas levels indicate impending transformer failure. Elution time (horizontal axis) is a measure of the time taken for each compound to travel through the chromatography column.

In this module, you will learn about the underlying principles of **chromatography** and about two simple forms of the technique called **paper chromatography** and **thin-layer chromatography**. Then, you will see how these two-dimensional tests can be expanded to a three-dimensional structure in **column chromatography.** These basic techniques are compared to the highly developed and sensitive instruments that are explained in Modules 13.2 and 13.3.

HOW CHROMATOGRAPHY WORKS

You can perform a simple chromatography experiment by putting spots of ink from felt-tip pens near one end of a strip of filter paper. The filter paper can then be suspended so the end closest to the ink spots is just immersed in water.

The original position of the ink spots on the paper is called the **origin**. The inks are a mixture of different-coloured chemicals. As the water soaks up the paper, the components of the inks are carried upwards at different rates and separate into spots of different colours. When the paper is removed from the water, the positions of the individual components are analysed, as seen in Figure 13.1.2.

FIGURE 13.1.2 Paper chromatography can be used to separate mixtures of inks into their separate components.

Each spot contains one of the components that are present in the ink mixture. The pattern of bands or spots is called a chromatogram. In the laboratory, paper chromatography is performed with high-quality absorbent paper, similar to filter paper.

All methods of chromatography have a:

- **stationary phase**
- moving or **mobile phase**.

In the simple experiment using inks and paper, the stationary phase is a thin layer of water adsorbed to the paper and the mobile phase is the water.

As the components in the ink are swept upwards over the stationary phase by the **solvent**, they undergo a continual process of **adsorption** (onto the stationary phase), followed by **desorption** and dissolving (into the mobile phase).

The amount that each component moves mainly depends on:

- how strongly the component adsorbs onto the stationary phase
- how readily the component dissolves in the mobile phase.

The components of the mixture separate because they undergo these two processes to a different degree.

When both **solute** and solvent contain molecules or ions with charge or charged areas, dissolving occurs (as described in Chapter 15). Water is a **polar solvent**. If water is used as the mobile phase, the more polar components will form dipole–dipole interactions and hydrogen bonds. These interactions result in the polar components being highly soluble in a polar solvent such as water. They will stay in the mobile phase and will tend to move further up the paper than the **non-polar** components.

Adsorbing onto the stationary phase and desorbing into the mobile phase are opposing processes that occur simultaneously. The processes can be likened to a 'tug of war' with the solute component being pulled in two directions. The degree to which a component undergoes these processes is determined by its physical properties, such as size, polarity and charge, and so its rate of movement through the medium, using a given solvent and temperature, will be unique. To maximise chromatographic separation of particular compounds of interest, chemists can alter both the stationary phase medium and the mobile phase solvent.

The term 'adsorption' describes the adhesion of molecules or substances to the surface of a solid or liquid. This is why the stationary phase in chromatography is either a solid or a liquid on a solid support. The mobile phase is a liquid or gas state.

Comparing chromatographic techniques

This chapter will describe different chromatographic techniques that can be used for analysis: paper chromatography, thin-layer chromatography, **gas chromatography (GC)** and **high-performance, or high-pressure, liquid chromatography (HPLC)**. These techniques use the same principles.

The particular technique chosen depends on a number of factors, including the:

- properties of the components being separated (e.g. molecular mass)
- amount of sample available for analysis
- concentration of the component in the sample
- sensitivity of the data required
- whether qualitative or quantitative data is required
- time taken for analysis
- cost of equipment.

Some of the important features of the different chromatographic techniques are listed in Table 13.1.1. Chromatography is widely used for the drug testing of blood and urine samples, monitoring water and air pollution, and testing food quality.

TABLE 13.1.1 Features of different chromatographic techniques

Technique	Typical substances tested	Typical samples	Advantages	Disadvantages	Comments
paper and thin-layer chromatography	• polar, water-soluble substances	• drug detection • dyes in foodstuffs	• very cheap, only basic laboratory equipment needed • easy to perform • ideal for fast qualitative analysis	• poor precision and accuracy • quantitative data is of low accuracy	• samples need to be coloured or visible under UV light, otherwise, samples can be made visible using stains
high-performance liquid chromatography	• medium to high molecular mass organic compounds, e.g. proteins, pesticides, enzymes	• wide range of foods, drugs, biological samples	• high quantitative sensitivity and precision • small sample size • readily automated	• moderately expensive instrument • trained technician needed to operate	• samples must be soluble in suitable solvent • suitable detecting device (e.g. mass spectrometer) is used
gas chromatography	• low molecular mass compounds (less than 300 g mol^{-1}) that vaporise without decomposing	• organic and inorganic analytical chemistry	• very high quantitative sensitivity • small sample size • readily automated	• expensive instrument • trained technician needed to operate	• samples must be a gas at the column temperature • suitable detecting device (e.g. mass spectrometer) is used

PAPER AND THIN-LAYER CHROMATOGRAPHY

In the laboratory, paper chromatography (Figure 13.1.2) is performed on high-quality absorbent paper, similar to filter paper, as the stationary phase. Thin-layer chromatography (TLC) is very similar to paper chromatography. In this case, the stationary phase is a thin layer of a fine powder, such as alumina (aluminium oxide) or silica, spread on a glass or plastic plate. Both techniques are useful for qualitative analysis. They can indicate which chemicals are present in the mixture but cannot give their specific concentration.

In both paper and thin-layer chromatography, a small spot of the solution of the sample to be analysed is placed on one end of the chromatography paper or plate. The position of this spot is called the origin. The paper or plate is then placed in a container with solvent. The origin must be a little above the level of the solvent so that the components can be transported up the paper or plate and not dissolve into the liquid in the container. As the solvent rises up the paper or plate, the components of each sample separate, as shown in Figure 13.1.3. The paper or plate is removed from the solvent before the liquid reaches the end, and the final position (distance from the origin) of each component and the **solvent front** are recorded.

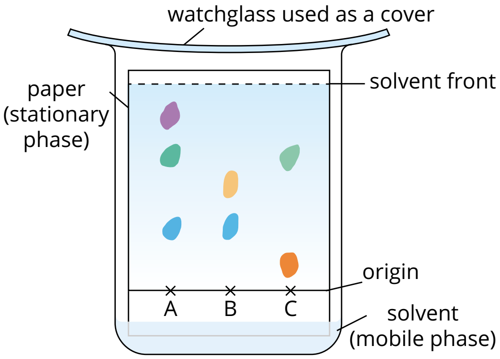


FIGURE 13.1.3 Paper chromatography of three different food colours (A, B and C)

Although the components of inks can be seen easily in chromatograms, most compounds are colourless and must be made visible. Many organic compounds **fluoresce** and appear blue when viewed under ultraviolet light. Another way to make colourless components visible is to spray them with a chemical that reacts to form coloured or fluorescent compounds.

The choice between paper and thin-layer chromatography depends upon the sample being analysed. Advantages of each method are given in Table 13.1.2.

TABLE 13.1.2 A comparison of paper and thin-layer chromatography

Paper chromatography	Thin-layer chromatography
is cheap	detects smaller amounts
requires little preparation	achieves better separation of less polar compounds
is more efficient for polar and water-soluble compounds	can use corrosive materials such as organic solvents
is easy to handle and store	has a wide range of stationary phases

Chromatography results—using R_f values and standards

The components in a mixture separated by chromatography can be identified in two ways:

1 by calculating the **retardation factor (R_f)** of the sample

2 by including **standards** of known chemicals on the same chromatogram as the unknown sample.

Method 1: Calculating R_f values

Each component of a mixture will travel a specific distance up the stationary phase. By comparing the distance travelled by the component to the position of the edge of the solvent (solvent front) after it has finished moving, a retardation factor can be calculated.

The retardation factor (R_f) for a component is given by:

$$R_f = \frac{\text{distance the component travels from the origin}}{\text{distance the solvent front travels from the origin}}$$

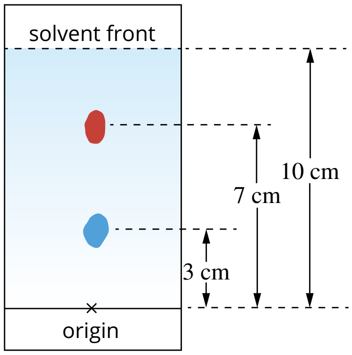


FIGURE 13.1.4 A thin-layer chromatogram of a sample with two components. The distances from the origin enable you to calculate R_f.

You can see from Figure 13.1.4 that:

- R_f values are always less than one
- the component most strongly adsorbed onto the stationary phase moves the shortest distance and has the lowest R_f value
- the component most soluble in the mobile phase moves the furthest and has the largest R_f value.

Each component has a characteristic R_f value for the conditions under which the chromatogram was obtained. By comparing the R_f values of components of a particular mixture with the R_f values of known substances determined under identical conditions, you can identify the components present in a mixture.

In this method, the distance moved by the solvent front is not critical. R_f values are a ratio and have no units. The proportion of the distance moved from the origin (the R_f value) stays the same provided the conditions under which the chromatogram is obtained are the same. This means the R_f values of unknown spots can be compared against a table of R_f values of common materials.

However, changes in the temperature, the type of stationary phase, the amount of water vapour (around the plate or paper) and the type of solvent all change the R_f value for a particular chemical. For the R_f value of a component to be comparable to a set of standard R_f values, they must be determined under identical conditions.

Worked example 13.1.1

CALCULATING R_f VALUES

Using Figure 13.1.4, calculate the R_f value of the blue component.	
Thinking	**Working**
Record the distance the component has moved from the origin and the distance the solvent front has moved from the origin.	The distance the blue component has moved from the origin is 3 cm. The distance the solvent front has moved from the origin is 10 cm.
$R_f = \frac{\text{distance the component travels from the origin}}{\text{distance the solvent front travels from the origin}}$	R_f (blue component) $= \frac{3}{10}$ $= 0.3$

➤ Try yourself 13.1.1

CALCULATING R_f VALUES

Using Figure 13.1.4, calculate the R_f value of the red component.

Method 2: Using standards

In this method, you need to have an understanding of the chemicals that might be present in the sample. The sample is tested on the same chromatogram as pure samples of the possible components in the unknown. These pure samples are known as standards. An example of this strategy is testing a lolly for the presence of glucose and fructose. If spots from the lolly sample move the same distance from the origin as the spots from the fructose and glucose standards, then the lolly is likely to contain the sugars (Figure 13.1.5).

Standards can also be used in paper and thin-layer chromatography as an indication of the amount of a substance present. By applying the standard in a range of concentrations, a visual comparison is made with the sample of interest. This technique of quantitative analysis, while quick and informative, is not as accurate as GC and HPLC (refer to Module 13.4).

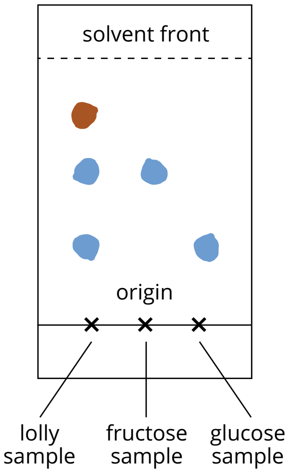


FIGURE 13.1.5 Identification of sugars in a lolly using paper or thin-layer chromatography. The lolly is likely to contain fructose and glucose, as well as a brown-coloured component.

Chromatography of amino acids

Amino acids are the monomers that combine to make all proteins. There are 20 common types, each one containing an acid group (–COOH) and an amine group ($-NH_3$). The generic structure of an amino acid is shown in Figure 13.1.6. These molecules contain similar polar areas and variable non-polar regions. They can be separated using thin-layer chromatography due to these differences. Although not coloured, their position on the chromatography plate is revealed using a ninhydrin stain (Figure 13.1.7).

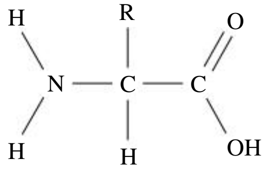


FIGURE 13.1.6 The 20 common amino acids have the same basic structure. The R group is a branched, straight chain, aromatic or other functional group.

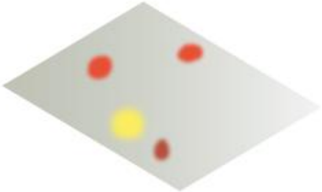


FIGURE 13.1.7 Amino acids separated using silica thin-layer chromatography. Each spot contains a 5 μg sample and the plate was eluted with an *n*-butanol:ethanoic acid water 3:1:1 (by volume) mixture followed by ninhydrin stain.

COLUMN CHROMATOGRAPHY

Column chromatography uses the same principles as paper and thin-layer chromatography to separate a mixture but the stationary phase is a three-dimensional porous column structure rather than two-dimensional paper or plate. Two key examples are gas chromatography and high-performance liquid chromatography—these are explained in Modules 13.2 and 13.3.

In column chromatography, shown in Figure 13.1.8, the stationary phase is small particles loosely packed into a glass or metal column. The sample mixture is applied carefully to one end of the column, and a solvent, which acts as the mobile phase, is dripped onto or pumped through the column from a reservoir above. Separated components leave the column at the other end with the solvent, which is now called the **eluent**.

As the components are carried down the column, they move between the particles of the stationary phase and are repeatedly adsorbed to the stationary phase and desorbed back into the mobile phase (Figure 13.1.8). The components of the sample undergo the processes of adsorption and desorption to different degrees based on their physical characteristics, so they move at different rates downwards with the eluent.

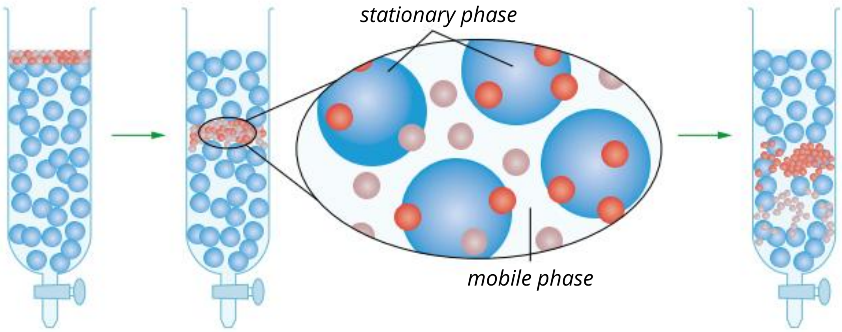


FIGURE 13.1.8 Column chromatography is used to separate the components in a mixture. Interaction with the stationary phase of the column causes the red-coloured components of the mixture to move more slowly. Components of the mixture that are highly soluble in the eluent and do not interact with the stationary phase (grey colour) will move more quickly, and can be collected first.

In paper, thin-layer and column chromatography, charge interactions between components of the mixture and the mobile phase and stationary phase cause different rates of movement. This allows components of the sample to separate. However, if the physical properties (e.g. polarity) of two substances are similar they will travel at similar speeds. In this situation, the molecular mass of the substances may allow them to separate. This would occur in cases where there is a large difference in molecular mass.

13.1 Review

SUMMARY

- Chromatography is a technique commonly used to separate and identify the components in a mixture.
- Paper and thin-layer chromatography are simple forms of this technique.
- All chromatographic techniques involve a mobile phase and a stationary phase.
- Components separate during chromatography because of how strongly they adsorb to the stationary phase and desorb back into the mobile phase.
- Components in a mixture have differing affinities for the mobile and stationary phases.
- In paper and thin-layer chromatography, the components in a mixture are identified by comparison with known standards and determination of $R_f = \frac{\text{distance the component travels from the origin}}{\text{distance the solvent front travels from the origin}}$ values.
- Column chromatography also uses these principles and processes in a cylindrical three-dimensional structure.
- When two compounds have very similar interactions with the mobile and stationary phase, they can be separated based on different molecular weights.
- A component with similar polarity to the mobile phase (compared to the stationary phase) will have a high R_f value.

KEY QUESTIONS

Retrieval

1 Identify the correct description for each term.

Term	Description
adsorption	the breaking of the attraction between a substance and the surface to which the substance is adsorbed
desorption	the different compounds in the mixture, which can be separated by chromatography
components	the components of a mixture undergo adsorption to this phase
polar molecule	the attraction of one substance to the surface of another
mobile phase	a molecule that acts as a dipole; it has one or more polar covalent bonds, with the charge being distributed asymmetrically
stationary phase	the solvent that moves through the stationary phase in chromatography

2 Use the following terms to complete the sentences about paper chromatography. Terms can be used more than once.

stationary	mobile	components
origin	phase	sample
above	below	phases

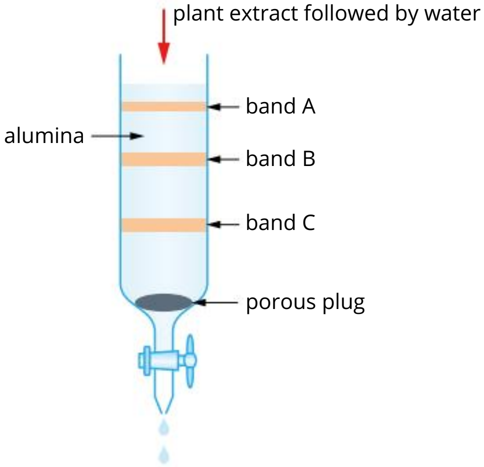
In paper chromatography, the paper acts as the ________ phase. A small spot of a solution is placed at one end of the paper, called the ________. The sample solution contains a number of different coloured compounds, the ________. The paper is suspended so that the end with the spot is ________ the surface of the solvent. The solvent or ________ phase moves up the ________ phase. Different coloured spots are observed at various places on the paper, due to the separation of different ________.

Comprehension

3 Phenacetin was once an ingredient in analgesic drugs, but it is not used now because it can cause liver damage. It is soluble in chloroform. A chemist wishes to analyse a brand of analgesic by thin-layer chromatography to determine whether it contains phenacetin. Describe the steps in the analysis. (Assume that a sample of pure phenacetin is available to the chemist.)

4 The components in a plant extract were separated by simple column chromatography. The mobile phase used was water and the column was packed with alumina powder. Use the diagram below and the following terms to complete the sentences.

band A	band B	band C
used coarser alumina powder	increased the column length	decreased the column length
smaller	larger	



The component most strongly adsorbed to the stationary phase is ________________.

As band B begins to emerge from the column, it appears to separate into two bands. These two components would be more successfully separated if you ________________.

GC and HPLC could provide clearer results because the particles in the stationary phase are ________________.

5 For each of the following types of chromatography, determine the:

- **i** mobile phase
- **ii** stationary phase

a paper chromatography

b column chromatography

c thin-layer chromatography

6 Explain two reasons why the level of the solvent must be below the level of the origin where the samples are placed in paper chromatography.

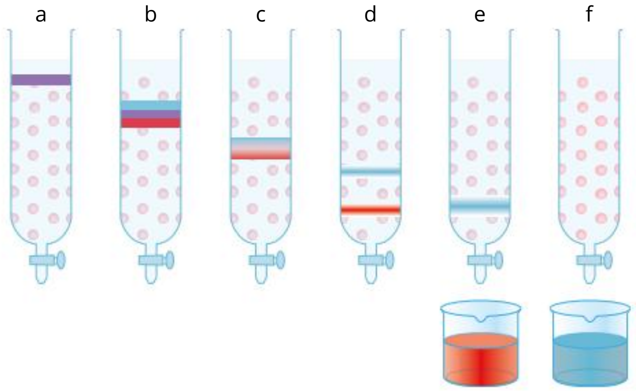
7 Thin-layer chromatography showed that the black dye used in a brand of writing ink contained blue, red, orange and yellow components. The R_f values of these substances in ethanol as solvent are 0.59 (blue), 0.32 (red), 0.80 (orange) and 0.19 (yellow).

a Determine the distance between the blue and yellow components after the solvent front had moved 8.0 cm from the origin.

b Calculate how far the orange component would have travelled when the red component had travelled 6.0 cm from the origin.

c Draw the chromatogram of the ink to scale after the solvent front had moved 15 cm from the origin.

8 Purple ink is separated using a column containing a polar support, as shown in the figure below. Describe the process depicted in the figure.



Analysis

9 Each chromatography technique has disadvantages associated with it. Identify the problem or problems associated with each technique in the table below using the options provided. Options may be used more than once.

A requires relatively large amounts of solvents to operate

B difficult or impossible to obtain quantitative data

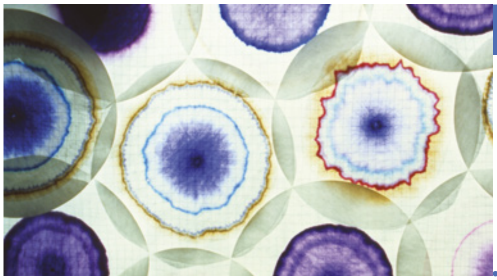
C expensive equipment needed

D samples must be able to be dissolved in solvent

Technique	Problem
paper or thin-layer chromatography	
high-performance chromatography and gas chromatography	

10 Ethanol is a polar molecule. In comparison to ethanol, propanone is less polar and ethanoic acid is much less polar. Determine the expected order of elution of these three when a mixture is applied to a chromatography column that has a non-polar stationary phase and a polar mobile phase.

13.2 Gas chromatography



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the structure of a gas chromatograph and the chromatography principles it uses
- understand the extreme sensitivity of the technique in detecting compounds in mixtures
- understand the importance of retention time in compound identification.

The most sensitive of the column chromatographic techniques is gas chromatography (GC). GC is capable of detecting as little as 10^{-12} g of a compound in a mixture. However, it is limited to compounds that can be readily vaporised without decomposing. Such compounds usually have relative molecular masses less than 300 g mol^{-1}. GC requires an expensive instrument and trained operators.

The extreme sensitivity of GC makes it ideal for the analysis of **trace contaminants** in samples or for the detection of tiny amounts of very potent compounds. Gas chromatography is used extensively in forensic science, arson investigations, drug testing in sport, paint chip analysis and toxicology cases to identify and quantify various biological specimens and crime-scene evidence.

HOW GAS CHROMATOGRAPHY WORKS

There are two types of GC: **gas–liquid chromatography (GLC)** and **gas–solid chromatography (GSC)**. Both GC techniques use chromatography principles described in Module 13.1. A mobile phase moves through a stationary phase and separated components are identified as they leave the column. A generalised diagram of the equipment used in GC is shown in Figure 13.2.1.

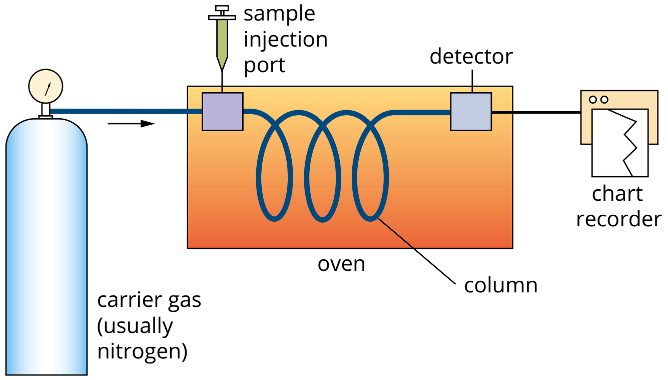


FIGURE 13.2.1 A gas chromatograph

Gas chromatography has the following features.

- The mobile phase is a gas, generally nitrogen, called the **carrier gas**.
- A small amount of sample is injected into the top of the column through an injection port.
- The injection port is heated to a temperature sufficient to instantly vaporise the sample, which is then swept into the column by the carrier gas.
- The column is a loop, or series of loops, of glass that has an internal diameter of about 4 mm and is 2–3 m long in total. In GLC, the column is packed with a porous solid that has been coated with a liquid hydrocarbon or ester with a high boiling point. This liquid acts as a liquid stationary phase. In GSC, the column is packed with an adsorbent solid such as silica gel or alumina. The solid acts as a solid stationary phase.

- The column is mounted in an oven and heated.
- The components of the sample repeatedly adsorb onto and desorb from the stationary phase. The least soluble components are swept out first by the gas into the detector (Figure 13.2.1). The detector is often a mass spectrometer. In Chapter 5, you learnt how mass spectrometry is used to analyse organic compounds.

The detector forms a trace (or peak) as each substance exits the column and is recorded. The resulting trace is called a chromatogram—it shows the response of the detector against time. The response of the solvent is constant and gives a flat baseline. The sample is injected into the column at time zero. Different chemicals show different strengths of attraction for the stationary phase and the mobile phase, and so pass through the column at different rates. The time a component takes to pass through the column is called the **retention time (R_t)**. Retention time is similar to the retention factor (R_f) obtained using paper chromatography or thin-layer chromatography. The same compound gives the same retention time if the conditions (temperature, mobile phase, stationary phase, flow rate, pressure etc.) remain the same.

ⓘ Retention time (R_t) is used to identify the components associated with the peaks on a chromatogram. It is the time that a component takes to traverse the column.

Drug testing in sport

Since competitive sports began, some participants have sought to gain an advantage by whatever means available. For example, athletes in ancient Greece and Rome consumed specially prepared lizard meat in the hope that it would improve their performance and give them a competitive edge.

Today, drug problems in sport usually result from the inappropriate use of therapeutic substances and manipulation of substances such as hormones and electrolytes that occur naturally in an animal's or athlete's system. In horseracing, the use of illegal 'go-fast' or 'stopper' drugs is now relatively rare because they are easily detected by the combined analytical technique of **gas chromatography–mass spectrometry**, or **GC–MS**.

Many of these substances are so potent that they are used in miniscule amounts and the extreme sensitivity of gas chromatography is required. Urine samples are routinely taken from athletes competing in major events to ensure that the athletes are not benefiting from the use of illegal performance-enhancing drugs (Figure 13.2.2). These samples are analysed by gas chromatography. The combined technique, GC–MS, uses a **mass spectrometer** as the detector in the gas chromatography.

FIGURE 13.2.2 Drug testing in sport—a chemist analyses athletes' samples for illegal substances.

13.2 Review

SUMMARY

- Gas chromatography (GC) is a very sensitive technique and is used for qualitative and quantitative analysis.
- In GC, the mobile phase is a hot gas, and the stationary phase can be a liquid or a solid.
- Retention time is used to identify components in a mixture, and components are often identified, when exiting the column, by mass spectrometry (GC–MS).
- A component with similar polarity to the mobile phase (compared to the stationary phase) will have a low R_t value.

KEY QUESTIONS

Retrieval

1 Describe why the chromatography described in this module is called 'gas' chromatography.

2 Identify what detector is often used to identify compounds as they are eluted from the column.

Comprehension

3 Compounds A and B are equally soluble in nitrogen gas, but A is slightly more strongly adsorbed on silica particles. Compound C has a lower solubility in nitrogen than A and B. A sample containing A, B and C is injected into a GC instrument, which uses nitrogen as the solvent and silica as the stationary phase.

List the order of the components as they are likely to emerge from the column.

4 Some herbicides have an extremely potent effect on plants. Gas chromatography can be used to detect very small amounts in natural water samples.

a By rearranging the following steps in the procedure, describe how a herbicide is detected in creek water.

A Compare the water sample and the spiked water sample to identify the presence of the herbicide.

B Perform GC chromatography on the spiked creek water sample.

C Obtain a pure sample of the herbicide of interest.

D Spike the creek water sample with the herbicide of interest.

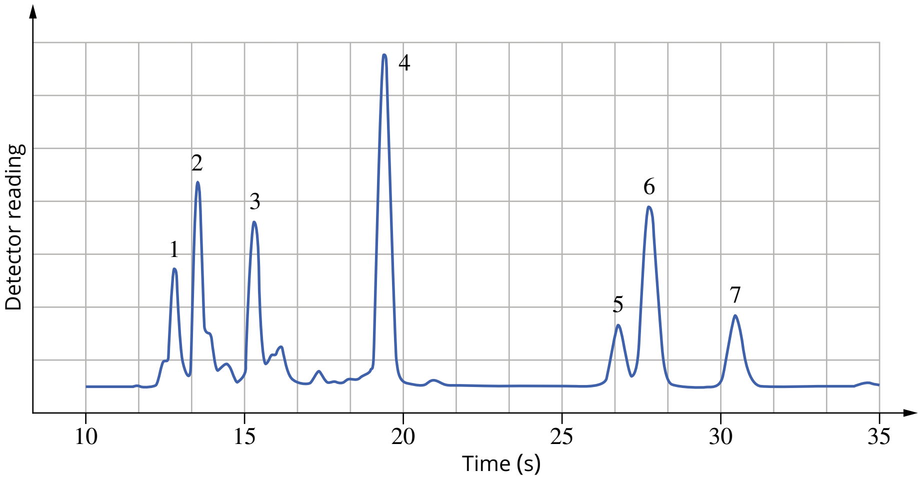
E Perform GC chromatography on the creek water sample.

b Describe how R_t could be used.

5 Describe the advantages of GC for the analysis of drugs compared to an analysis technique based on column chromatography.

Analysis

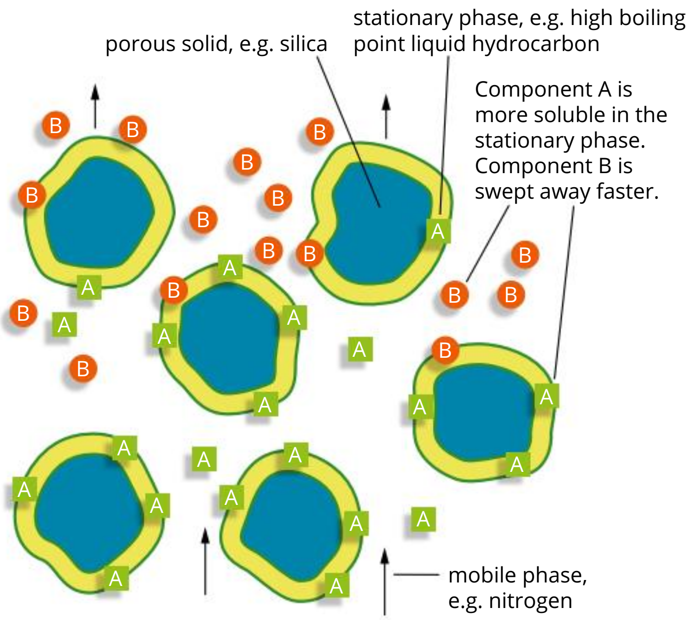
6 GC analysis of petrol is shown below. Petrol is composed of octane and other larger hydrocarbons and aromatic compounds. Component 1 passed through the GC column most quickly as shown by time of about 13 seconds.



a Identify the peak relating to octane.

b This GC column has a non-polar stationary phase. Deduce the polarity of the remaining components relative to that of peak 4.

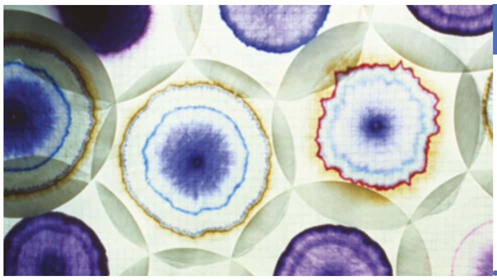
7 The figure below shows a diagrammatic representation of the interaction between the components of a sample (a mixture of A and B) and the stationary phase of a GC column. These particles of the stationary phase are silica coated with a high boiling point liquid hydrocarbon.



Consider the interactions occurring in the column, and order of elution of A and B, and summarise these in a short paragraph using the following terms.

stationary phase
elute
non-polar interactions
adsorption
mobile phase
retention time
polar molecules
desorb

13.3 High-performance liquid chromatography



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the structure of an HPLC column and the chromatography principles it uses
- understand the sensitivity of the technique in detecting compounds in a mixture
- understand the importance of retention time in compound identification.

High-performance (or high-pressure) liquid chromatography (HPLC) is an instrumental chromatographic technique based on column chromatography that allows scientists to perform sensitive analyses of a wide range of mixtures. It is commonly used for the separation and identification of complex mixtures of similar compounds, such as contaminants that are soluble in water, drugs in blood and hydrocarbons in oil samples.

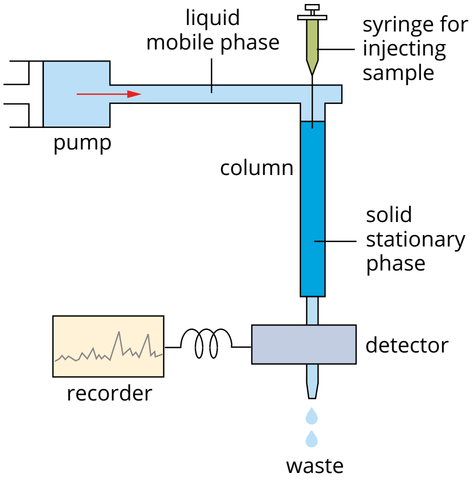


FIGURE 13.3.1 The construction of a high-performance liquid chromatograph

HOW HPLC WORKS

HPLC is now used routinely for environmental, pharmaceutical and industrial analyses. For example, it can be used to analyse for the presence and concentration of dioxins, insecticides, pesticides and oil spills in water. It is also used to determine the presence of pesticides in food or to detect the presence of drugs in blood.

Figure 13.3.1 shows the apparatus used for HPLC. The basic principles of HPLC are the same as for chromatography described earlier. The particles of the stationary phase are very small, resulting in good separation of compounds. However, the small particles result in smaller spaces between the particles and reduced flow rates. Eluting liquids (mobile phase) must be pumped through the column under high pressure.

A range of solids is available for use in HPLC columns, some with chemicals specially bonded to their surfaces to improve the separation of particular classes of compounds. Figure 13.3.2 shows resin beads that often form the stationary phase of the column.

In HPLC, the components are usually detected by passing the eluent stream through a beam of ultraviolet (UV) light. Many organic compounds absorb UV light, so when an organic compound passes in front of the beam of light, a reduced signal is picked up by a detector. The amount of light received by the detector is recorded on a chart that moves slowly at a constant speed or sent to a computer, which produces the chromatogram.

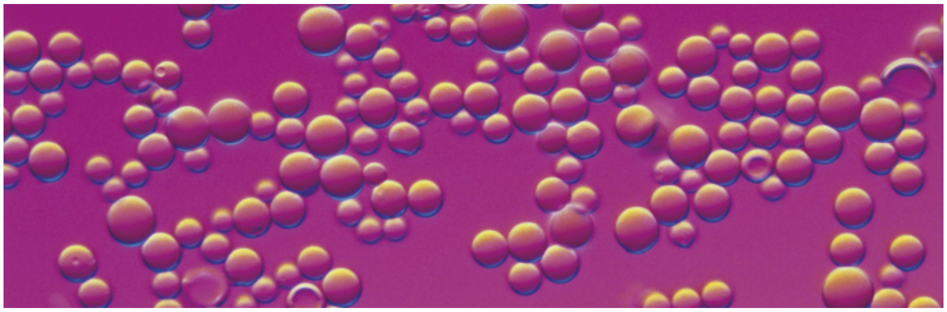


FIGURE 13.3.2 Light micrograph of resin beads used in high-performance liquid chromatography. The beads are coated with a hydrophobic, waxy substance that has a long molecular backbone of 18 carbon atoms, and are packed into a column, usually a steel tube, through which the sample is pumped under pressure.

Figure 13.3.3 indicates how the chromatogram is produced as the components pass through the instrument.

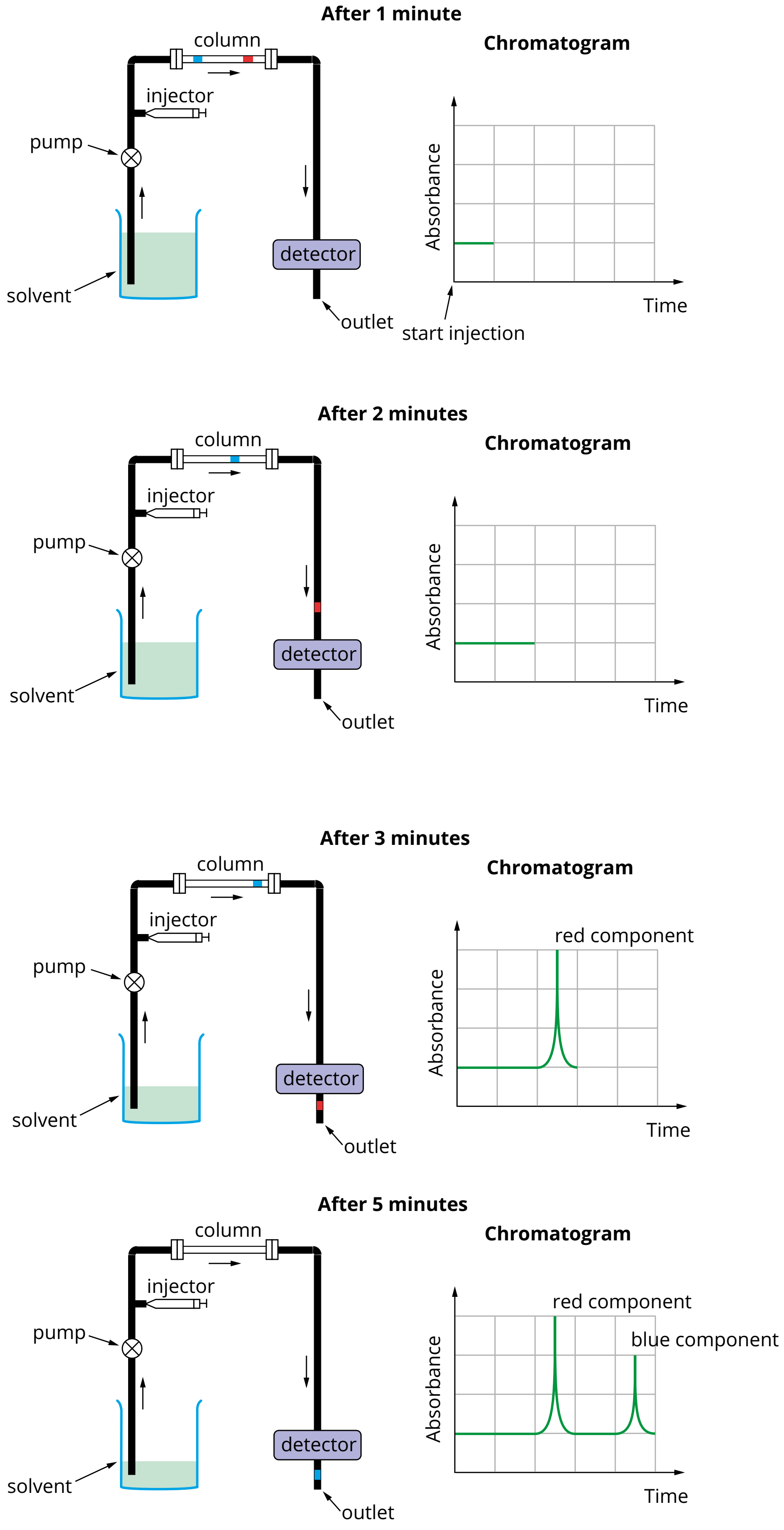


FIGURE 13.3.3 Separation of components of a sample and generation of peaks on an HPLC chromatogram

To identify a particular substance of interest a solution of a pure compound that is thought to be one of the components is injected into the HPLC under the same conditions. The chromatogram of the standard is then compared with the chromatogram of the sample.

The same compound will have the same retention time if the conditions (temperature, mobile phase, stationary phase, flow rate, pressure etc.) remain the same. This is a qualitative result.

Detecting pesticides using HPLC

Pesticides can be used to prevent mould in food such as the lettuce used in salads. HPLC can be used to test wastewater samples from a farm for the presence of two pesticides: vinclozolin and carbendazim.

The high-performance liquid chromatogram of a water sample is shown in Figure 13.3.4. The two component pesticide peaks, marked 1 and 2, are identified by obtaining chromatograms of known pure compounds under exactly the same conditions as the sample.

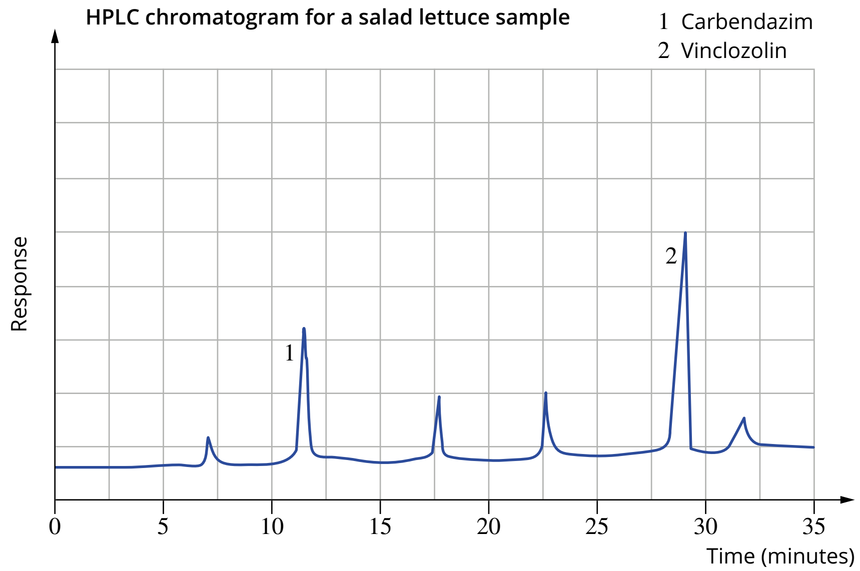


FIGURE 13.3.4 High-performance liquid chromatogram of a water sample from a farm

A compound can also be identified by adding a known compound to the sample (a procedure called **spiking**). The chromatogram in Figure 13.3.5 was produced when the water sample was spiked with carbendazim. Note that peak 1 is much bigger in the spiked sample. There are no extra peaks, indicating that carbendazim must have been component 1 in the sample.

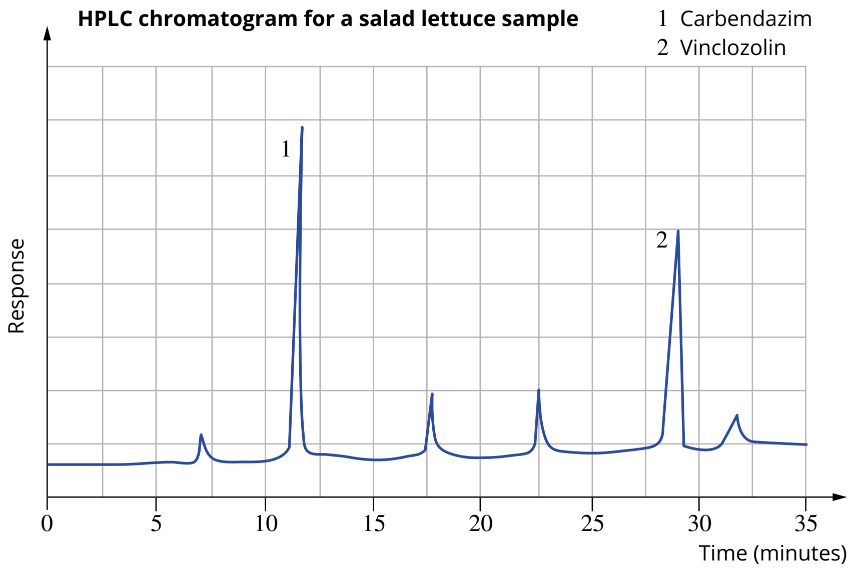


FIGURE 13.3.5 High-performance liquid chromatogram for a carbendazim-spiked water sample

In Figure 13.3.6, you can see an example of a chromatogram on a computer display of HPLC-separated components. Scientists view the eluted substances detected by the monitor as they exit the column.

WS
2.1.3

FIGURE 13.3.6 A research chemist uses HPLC to analyse the components in a mixture for medical research.

Making the grade

Gourmet cooks value olive oil (Figure 13.3.7) for its distinctive taste and aroma. Nutritionists favour it because it is rich in monounsaturated triglycerides, which are believed to lower blood cholesterol levels and reduce the risk of heart disease. But some of the companies that process and distribute olive oil succumb to the temptation to mix their olive oil with less expensive oils such as corn, peanut and soybean oils.

A European company sent a shipment of olive oil to the United States. A routine sample was taken and forwarded to chemist Richard Flor at the US Customs Service laboratory in Washington, DC. The sample looked and poured like olive oil, but it did not taste quite right. Because taste is a subjective test and does not always stand up in court, Flor and his colleague Le Tiet Hecking developed an analytical test based on olive oil's unique composition.

Flor used HPLC to separate the oils into their component triglycerides. An example of the chromatographs obtained can be seen in Figure 13.3.8. Flor and Hecking found that the oil in the suspect sample had too much of the polyunsaturated triglycerides, confirming their suspicion that the imported sample was diluted with corn oil.

FIGURE 13.3.7 Pure olive oil is highly regarded for its aroma and taste, and because it is thought to reduce blood cholesterol levels.

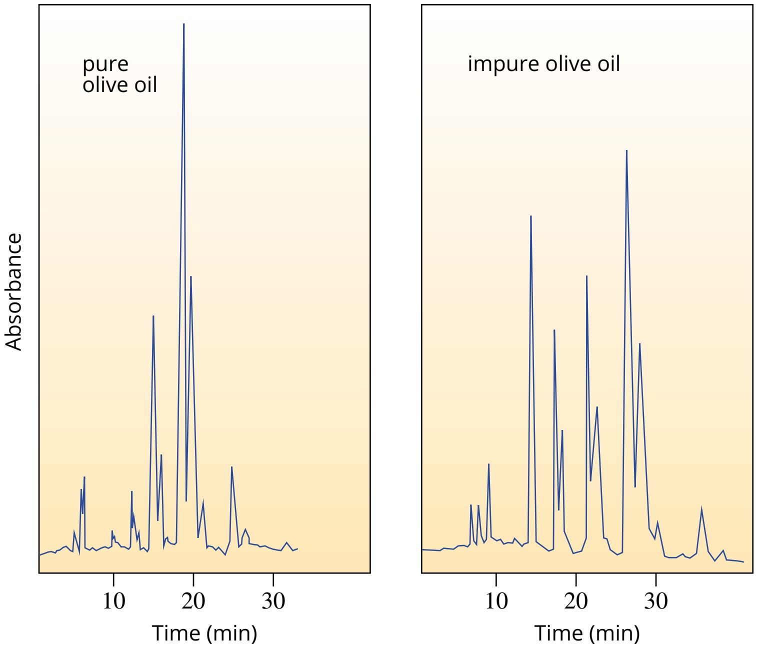


FIGURE 13.3.8 Peaks in the chromatogram of the suspect oil indicate the presence of compounds found in corn oil, rather than olive oil.

HPLC can be combined with mass spectroscopy. Together, the techniques allow chemists to determine smaller and smaller quantities and identify the chemical structure of the compounds in the sample, based on the fragmentation pattern of each component.

13.3 Review

SUMMARY

- High-performance liquid chromatography (HPLC) is a very sensitive column chromatography technique used for qualitative and quantitative analysis.
- In HPLC, the mobile phase is liquid under pressure.
- The retention time is used to identify components in a mixture.
- As components of the mixture are tested in liquid form, HPLC can analyse a wider variety of substances than GC.
- A component with similar polarity to the mobile phase (compared to the stationary phase) will have a low R_t value.

KEY QUESTIONS

Retrieval

1 Identify what HPLC stands for.

2 Describe two variables of the HPLC separation conditions that can be changed to alter the retention time of a compound.

3 Describe why HPLC can test a wide variety of substances.

4 State which feature of a chromatogram obtained during a HPLC analysis is:
  a used to identify the components of a mixture
  b least affected by a change in concentration of the sample components
  c used to determine the concentration of each component in a mixture

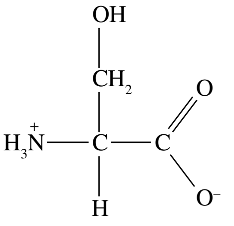
Comprehension

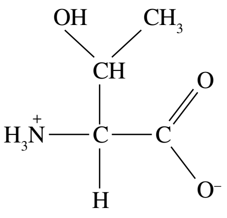
5 An energy drink powder is to be analysed by HPLC. Explain briefly how the components are separated by the HPLC technique.

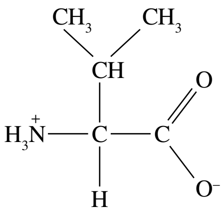
Analysis

6 A chemist working for an environmental agency wishes to analyse the levels of a herbicide in a country stream using HPLC. Sort the following steps in the analytical process into the correct order.
  A Measure the herbicide peak areas from the chromatograms of the standard solutions and the stream water sample.
  B Obtain chromatograms of the stream water and of standard solutions of the herbicide.
  C Construct a calibration curve and mark the herbicide peak area from the stream water sample on it.
  D Use R_t values to identify the herbicide peak on the chromatogram of stream water.
  E Determine the herbicide concentration in the stream water sample.

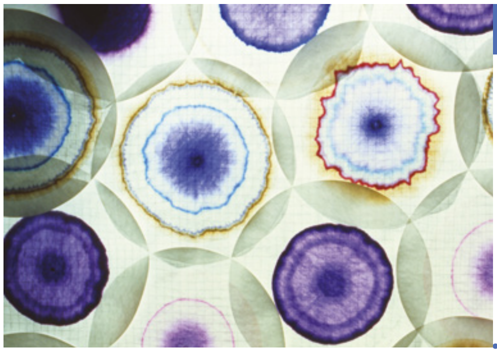
7 Reversed-phase HPLC uses a polar mobile phase and a non-polar stationary phase. A mixture of the amino acid molecules serine, threonine and valine (the structures of which are shown below) was analysed by reversed-phase HPLC. Organise the amino acids in the order that they would reach the detector.


serine (Ser)


threonine (Thr)


valine (Val)

13.4 Quantitative analysis using chromatography



BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the difference between qualitative and quantitative analysis
- understand how HPLC and GC are used to determine the concentration of a substance in a mixture
- construct a calibration curve (calibration graph) of the area under peaks of a range of standards
- use the calibration curve to determine the concentration of a compound of interest using the area under its peak on the chromatogram.

The retention time (in GC and HPLC) and the retention factor (in paper and thin-layer chromatography) provide information about the presence of a particular compound in a mixture. Using standards and/or R_f values the component can be identified. This means that scientists can report the presence or absence of substances of interest—this is a qualitative result. Many laboratory tests, however, are conducted to determine how much of each chemical is present and are searching for instances where a component may be outside an acceptable concentration range. These tests are quantitative and are vital for health and safety monitoring and pharmaceutical quality-assurance testing. Some examples are:

- organic pollutants in the environment
- crime analysis such as fire debris
- detection of active agents in medicines
- environmental analysis
- detection of metabolic products in medicine.

QUANTITATIVE HPLC ANALYSIS

To determine the concentration of an individual component in a mixture, its peak area is compared with the peak areas of samples of the same chemical at known concentrations. A solution with an accurately known concentration is called a **standard solution**.

By plotting the peak areas against the concentrations of the standard solutions, a **calibration curve** can be drawn and used to determine unknown concentrations. A mass spectrometer trace from HPLC is shown in Figure 13.4.1. The peak area is shaded.

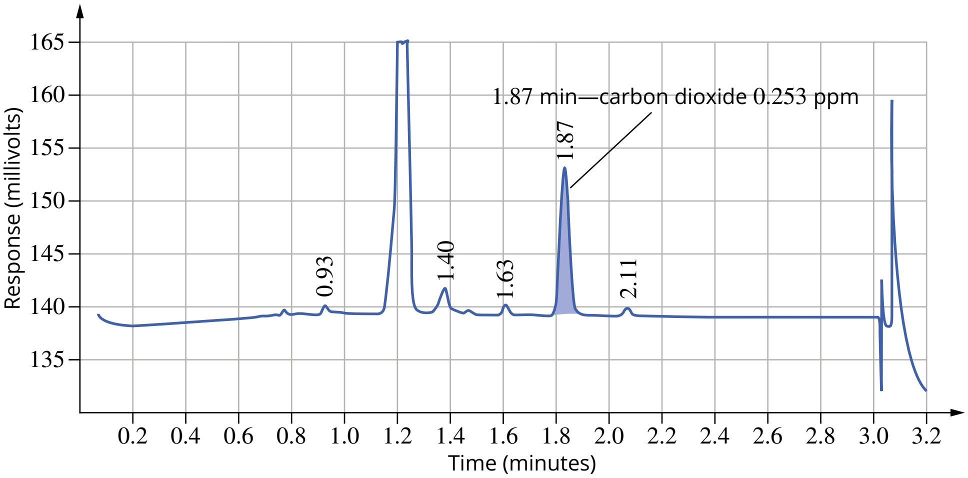


FIGURE 13.4.1 HPLC chromatogram showing the area under the carbon dioxide peak that corresponds to its concentration in the sample (0.253 ppm)

Worked example 13.4.1 shows how HPLC can be used to find the concentration of a component in a mixture.

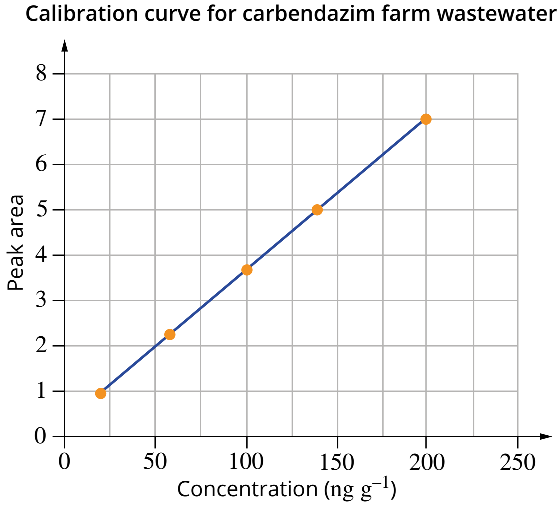
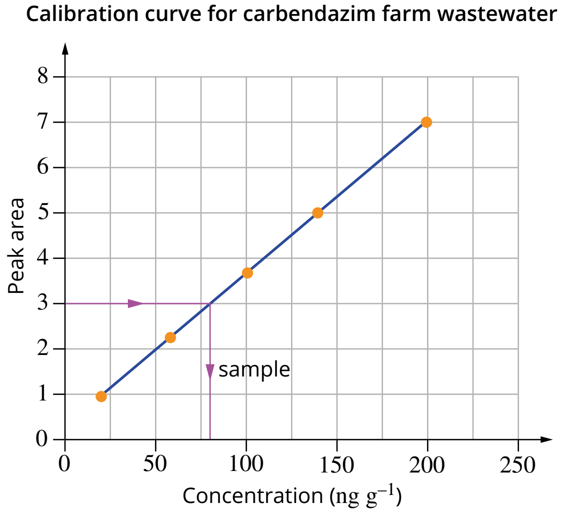
Worked example 13.4.1

CONCENTRATION OF A PESTICIDE

The concentration of the pesticide carbendazim in a sample of wastewater from a farm was determined by HPLC. The chromatograms of a series of standards with accurately known concentrations of carbendazim were obtained under the same conditions as the sample.

The results from the chromatograms of the sample and the standards are shown in the following table. Determine the concentration of carbendazim in the sample of wastewater.

	Peak area	**Concentration ($ng\,g^{-1}$) 1 ng = 10^{-9} g**
standard 1	0.95	20
standard 2	2.20	60
standard 3	3.60	100
standard 4	5.00	140
standard 5	7.00	200
sample	3.00	?

Thinking	**Working**
Construct a calibration curve by plotting peak area versus concentration.	
Mark where the peak area of the sample lies on the calibration curve by tracing a horizontal line to the curve and then a vertical line down to the *x*-axis.	
Determine the concentration of the sample by reading the concentration value for the unknown solution from the *x*-axis.	The concentration of carbendazim in the wastewater sample is 80 $ng\,g^{-1}$.

➤ Try yourself 13.4.1

CONCENTRATION OF A PESTICIDE

Procymidone is a pesticide used to prevent disease in oranges. Tests were performed by HPLC to determine the concentration of this pesticide in a sample of orange juice.

The chromatograms of a series of standards with accurately known concentrations of procymidone were obtained under the same conditions as the sample.

The results from the chromatograms of the sample and the standards are shown in the following table.

Determine the concentration of procymidone in the sample of orange juice.

	Peak area	**Concentration (mg kg^{-1})**
standard 1	10	0.5
standard 2	20	1.0
standard 3	30	1.5
standard 4	40	2.0
standard 5	50	2.5
sample	15	?

Most HPLC equipment contains software programs that will process data from standards and provide automatic calculation of peak area and amount of each component in the sample.

QUANTITATIVE GC ANALYSIS

As for HPLC, the concentration of an individual component in a mixture can be determined by comparing its peak area with the peak areas of samples with known concentrations of the same chemical. In most modern GC–MS systems, computer software is used to draw and integrate peaks, and match MS spectra to library spectra.

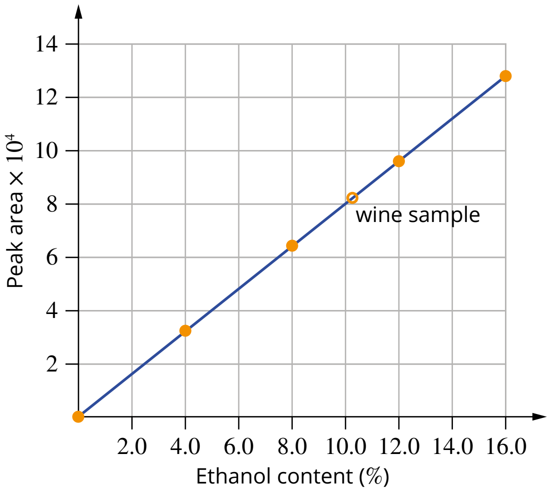
Worked example 13.4.2

CONCENTRATION OF A COMPONENT IN A MIXTURE

Australian wines are routinely tested for ethanol content using GC. The peak areas on a chromatogram produced by a sample of wine and a number of standard solutions of ethanol are shown below.

Determine the concentration of ethanol in the wine sample.

	Ethanol content (%)	**Relative ethanol peak area**
wine	?	82 400
standard 1	4.00	31 200
standard 2	8.00	62 900
standard 3	12.00	94 200
standard 4	16.00	125 700

Thinking	**Working**
Construct a calibration curve by plotting peak area versus ethanol concentration.	**Calibration curve for concentration of ethanol in wine** 
Determine the concentration of the sample by reading the concentration value for the unknown solution from the *x*-axis.	The concentration of ethanol in the wine sample is 10.2%.

➤ Try yourself 13.4.2

CONCENTRATION OF A COMPONENT IN A MIXTURE

Trinitrotoluene (TNT) is often blended with other explosives to suit particular applications. A sample from a solution of an explosive mixture was injected onto a GC column using a water/methanol mobile phase. The peak corresponding to TNT had an area of 8.8 mm^2. The peak areas for three standard solutions were also measured, as shown in the table.

Determine the concentration of TNT in the sample solution in $\mu\text{g mL}^{-1}$.

Peak areas from GC analysis of an explosive mixture and standards

TNT standards ($\mu\text{g mL}^{-1}$)	Peak area (mm^2)
2.0	6.0
4.0	11.6
6.0	17.6

PA
2.1.1

13.4 Review

SUMMARY

- Chromatographic techniques such as GC and HPLC are able to provide quantitative data (how much is present) as well as qualitative data (what compounds are present).
- Quantitative determination first requires a series of at least three samples (standards) of known concentration to be analysed.
- The areas under the peaks of the standards are plotted against concentration on a calibration curve.
- The area under the peaks of the unknown assayed under the same conditions will then provide its concentration (using the graph).

KEY QUESTIONS

Retrieval

1 Describe why scientists need quantitative chromatographic data.

2 Describe the role that computer software has in quantitative GC and HPLC.

Comprehension

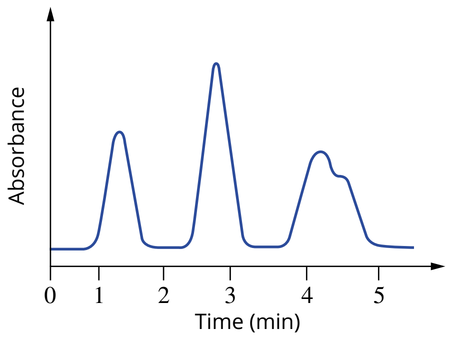
3 The maximum level of exposure to dioxin in water in the United States is 3.0×10^{-8} mg L^{-1}. Calculate the maximum mass of dioxin permitted in the following volumes.

a 1 mL of water

b a 600 mL bottle of water

c a 750 ML reservoir (ML = megalitre = 106 L)

4 A herbal tea extract was analysed by HPLC. The chromatogram obtained is shown in the figure below.



a Explain what information chemists can obtain from this chromatogram.

b State how many components are evident.

c Determine which component has the highest concentration.

Analysis

5 Determine why paper or thin-layer chromatography might be used as the initial investigation of a mixture rather than HPLC or GC. Identify data that paper and thin-layer chromatography do not provide.

6 A chemist working for a sports anti-doping agency analyses the testosterone levels of an athlete's blood by HPLC.

The chemist obtained the results shown in the table below for four testosterone standards and the athlete's blood sample. The concentrations of the standards were measured in nanograms per litre. Plot these values on a graph and determine the testosterone concentration in the athlete's blood.

Results of blood analysis

Concentration (ng L^{-1})	**Peak area**
20.0	0.12
40.0	0.22
60.0	0.36
80.0	0.46
blood sample	0.28

Chapter review

KEY TERMS

adsorption
amino acids
calibration curve
carrier gas
chromatogram
chromatography
column chromatography
desorption
eluent
fluoresce
gas chromatography (GC)
gas chromatography–mass spectrometry
gas–liquid chromatography (GLC)
gas–solid chromatography (GSC)
high-performance, or high-pressure, liquid chromatography (HPLC)
mass spectrometer
mobile phase
non-polar
origin
paper chromatography
polar solvent
qualitative analysis
quantitative analysis
retardation factor (R_f)
retention time (R_t)
solute
solvent
solvent front
spiking
standard
standard solution
stationary phase
thin-layer chromatography
trace contaminants

KEY QUESTIONS

Retrieval

1 Identify which of the following is the basis for separation in chromatography.

A interaction of chemicals with both mobile and stationary phases
B movement of chemicals through a column using a carrier gas
C passage of chemicals through a heated column under pressure
D small molecules becoming trapped in gaps between solid particles in a packed column

2 State which of the following can be determined by retention time in GC and HPLC.

A the identity of the chemical
B the concentration of a chemical
C the amount of a chemical in a sample
D the amount and concentration of a chemical sample

3 Determine which measured value most accurately determines the concentration of a substance in GC and HPLC.

A peak area
B peak height
C retention time
D R_f value

4 **a** Use the terms ‘adsorbed’ or ‘absorbed’ correctly in the sentences below.

i Water was ________________ by the towel as the wet swimmer dried himself.
ii A thin layer of grease ________________ onto the cup when it was washed in the dirty water.

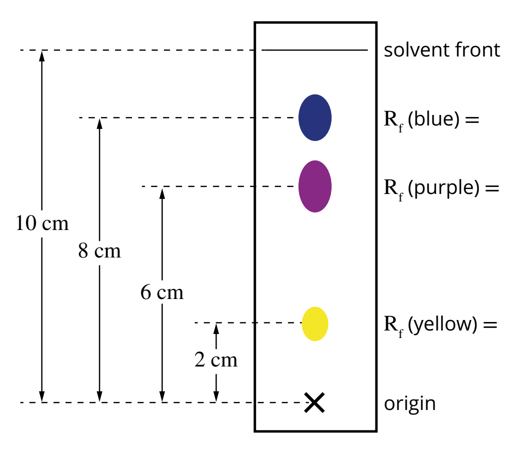
b State the meaning of the terms ‘adsorbed’ and ‘absorbed’.

5 Define each of the following terms.

chromatography
eluent
mobile phase
stationary phase
retention time

6 Describe how an R_f value is calculated.

7 The chromatogram of a dye is shown below. Identify the R_f values for each of the blue, purple and yellow components. Show answers to 1 decimal place.



8 A sample of brown dye from a lolly is placed at the origin on a chromatography plate. The solvent front moves 9.0 cm from the origin. A blue component of the dye moves 7.5 cm and a red component moves 5.2 cm in the same time. Indicate the R_f values of the two components.

9 Describe how the components are separated by GC.

Comprehension

10 Paper chromatography is used to separate the pigments in a plant leaf. A spot of each pigment is placed on a sheet of chromatography paper and the chromatogram is run inside a closed jar that is partly filled with ethanol. Select which one of the following options is the mobile phase in this instance.

A pigments
B ethanol
C paper
D ethanol vapour in the closed jar

11 From the following changes that could be made to an analysis to be performed by HPLC, determine which change is most likely to increase the retention time of a component.

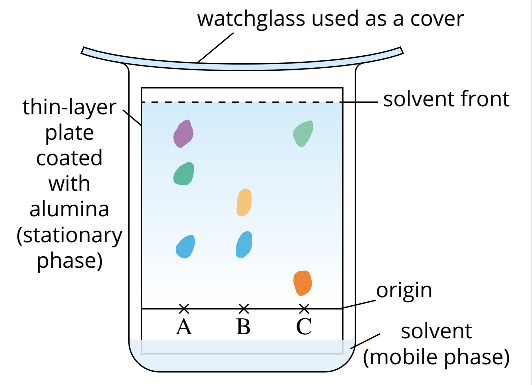
A reducing the length of the column
B increasing the amount of sample injected into the instrument
C using a more tightly packed column
D reducing the pressure of the mobile phase

12 Use the following terms to complete the summary of column chromatography. Not all terms are used.

mobile	adsorb	desorb
longest	shortest	stationary

During column chromatography, the components of the sample ________ onto the stationary phase and ________ into the liquid mobile phase. A component that adsorbs most strongly to the ________ phase and is least soluble in the ________ phase would be expected to take the ________ time to pass through the column.

13 Consider the paper chromatography shown in the figure below.



a Determine how many different components are present in sample B.

b Determine which component of colour A is least strongly adsorbed on the stationary phase.

c Identify which samples are likely to have one component in common.

d The green spot in sample A is 8.0 cm above the origin. The R_f value of this green spot is 0.74.
The orange and green spots of sample C are 3.0 cm and 9.0 cm above the origin respectively. Calculate the R_f values of each component of sample C.
Give your answer to 2 decimal places.

14 Consider the paper chromatogram of three food colours in Question **13**.

a Describe why the level of the solvent must be lower than the origin where spots of the mixture are originally placed.

b Explain why R_f values are always less than one.

c Determine how many different components are shown on the chromatogram.

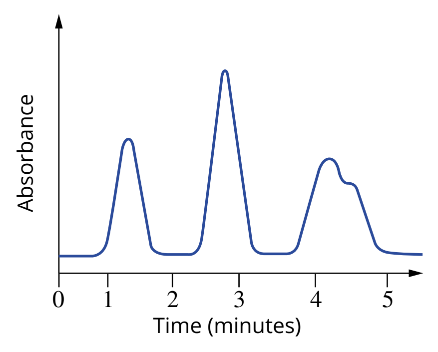
15 The ethanol content of wastewater from a brewery can be analysed by HPLC. The peak areas produced from an HPLC analysis of a sample of wastewater and a number of standard solutions of ethanol are shown in the table below.

	Ethanol (%)	**Relative peak area**
wastewater sample	?	5.6
standard 1	0.2	1.6
standard 2	0.4	3.2
standard 3	0.6	4.8
standard 4	0.8	6.4

a Draw a calibration curve of relative peak area against concentration of ethanol (%).

b Determine the percentage of ethanol in the wastewater sample, correct to 1 decimal place.

16 A herbal tea extract was analysed by HPLC. The chromatogram obtained is shown in the graph below.



Briefly explain how the components are separated by the HPLC technique.

Analysis

17 A dielectric constant is a measure of the polarity of a compound. A high dielectric constant refers to a polar compound. Using the data presented below, identify the position of the following compounds on a paper chromatogram using water as the solvent. Present your answer in the order of highest R_f value to lowest R_f value.

Name	Structure	Dielectric constant
hexane	$CH_3(CH_2)_4CH_3$	2.02
1-butanol	$CH_3CH_2CH_2CH_2–OH$	17.8
1-propanol	$CH_3CH_2CH_2–OH$	20.1
ethanol	$CH_3CH_2–OH$	24.3
methanol	$CH_3–OH$	33.0
acetonitrile	$H_3C–C≡N$	36.6

18 A chemist working for a local council wishes to analyse the composition of a river where children swim. It is thought the water might contain high levels of dioxins.

Determine which one of the following statements concerning the analysis of this water is correct.

A Paper chromatography would allow determination of the concentration of dioxins present but not their identity.

B The results obtained from HPLC would provide identification and accurate concentrations of the dioxins present.

C HPLC can identify the dioxins present but not their concentration.

D Paper chromatography would be suitable for identifying the dioxins and measuring their concentration.

19 A sample containing compounds A, B and C is mixed with ethanol and applied to the top of an HPLC instrument that uses ethanol as the solvent and alumina as the stationary phase. The three compounds have the following properties.

- Solubility in ethanol: C is much more soluble than A and B. A and B are equally soluble.
- Adsorption to alumina: B is more strongly absorbed than A. C is least strongly adsorbed.

Analyse the following lists to identify which one shows the retention times of the components, from lowest to highest.

A A, B, C

B C, B, A

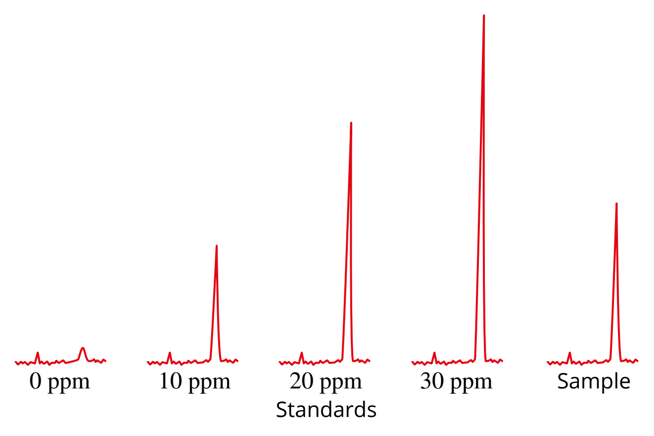
C C, A, B

D B, A, C

20 Compare the rates of adsorption and desorption for a component that appears at the top of a paper chromatogram to the rates for a component that appears at the bottom of the chromatogram.

21 The organophosphorus insecticide parathion has been widely used in mosquito-prone areas. An empty drum of the insecticide was found close to a major reservoir. The EPA was asked to analyse the water to determine whether it was a threat to human health. Levels above 0.01 $mg\,L^{-1}$ in water are a threat to human health. Parathion has an LD_{50} value (lethal dose to 50% of test animals exposed to this concentration) of 8 $mg\,kg^{-1}$.

Parathion standards in water of 0 ppm, 10 ppm, 20 ppm and 30 ppm parathion were prepared and analysed by GC, as shown in the diagram below. An undiluted sample of the reservoir water was run on the column under the same conditions. Note: 1 ppm = 1 µg/mL.



a Complete the table below.

Standard	Peak height (cm)
standard 0 ppm parathion	
standard 10 ppm parathion	
standard 20 ppm parathion	
standard 30 ppm parathion	
reservoir water	

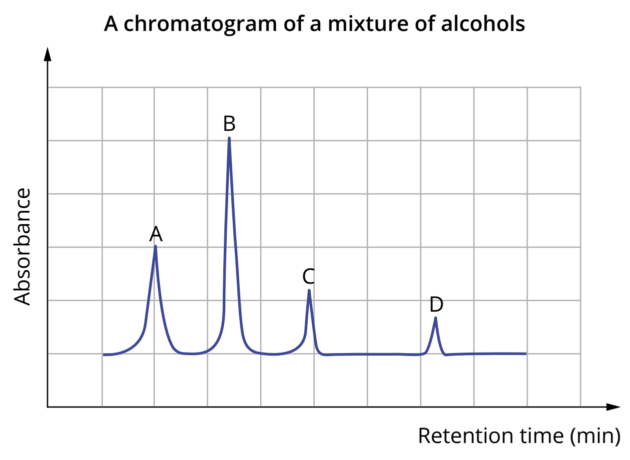
b Draw a calibration curve for the analysis of parathion.

c Determine the concentration of parathion in the water sample. (Assume that peak height is a measure of the concentration of the insecticide.)

d Evaluate if the reservoir water is within the legal limits for safe drinking. Support your answer with a calculation.

e Determine the volume of water a laboratory mouse, mass 150 g, you would need to drink to reach the LD_{50} dose.

22 A mixture was composed of four primary alcohols with formula ROH, where R is a straight-chain alkyl group. The mixture was analysed by GC using a non-polar stationary phase and a blend of water and methanol as the mobile phase. The chromatogram shown below was obtained.



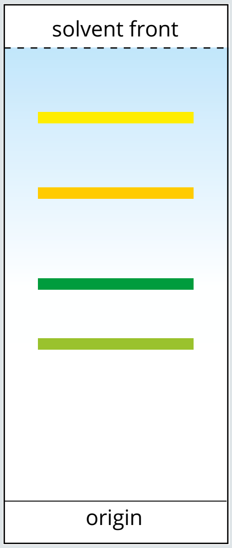
Determine the order from smallest to largest molar mass of the primary alcohols found in peaks A to D.

23 A government department analysed some imported olive oil by HPLC and found that the oil was impure and contained chemicals found only in corn oil. The peak area of one of the corn oil components in the chromatogram of a sample of the impure oil was 15.5 mm^2. The peak areas of four standard solutions of the component were also measured. The results of the olive oil analysis are shown in the table below.

Calculate the concentration of the corn oil component in the sample.

Corn oil component standards (%)	Peak area (mm^2)
1.0	5.8
2.0	11.6
3.0	17.6
4.0	23.0

24 An extract from a plant was analysed by thin-layer chromatography with a non-polar solvent. The chromatogram obtained is shown in the following diagram. The following table gives the R_f values of some chemicals commonly found in plants measured under the same conditions.



Chemicals found in plants	R_f
xanthophyll	0.67
β-carotene	0.82
chlorophyll a	0.48
chlorophyll b	0.35
leutin	0.39
neoxanthin	0.27

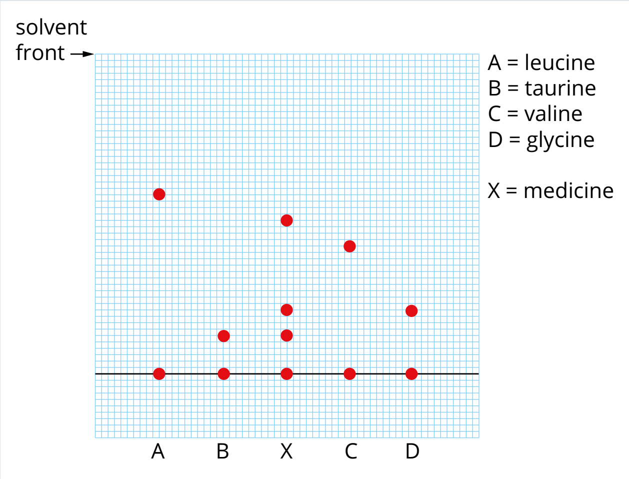
a Measure and record the distance from the origin to the centre of each band, and the distance of the solvent front from the origin.

b Calculate the R_f value of each band.

c Compare R_f values for the bands with the R_f values provided and name the chemicals present in the extract.

d Explain if the chromatogram would have appeared similar if water had been used as the solvent.

25 The diagram below shows a thin-layer chromatogram of amino acids in a medicine.



a Identify which amino acids are in the medicine.

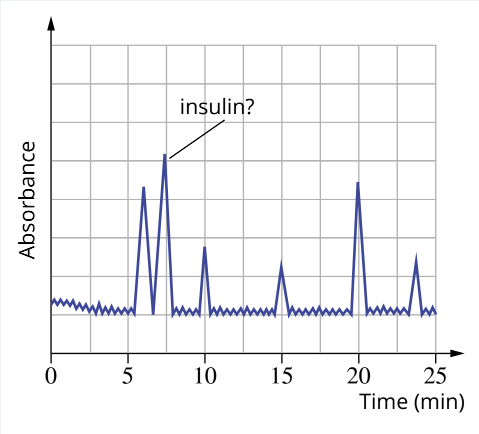
b Amino acids are colourless. Determine what technique could be used to make the amino acid spots visible.

c Analyse which amino acid is bonded least strongly to the stationary phase.

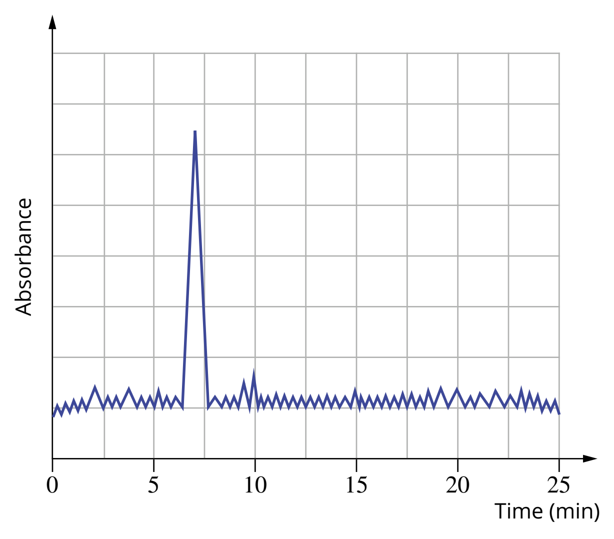
26 Explain the key stages in the thin-layer chromatography procedure used to produce the chromatogram shown in Question **25**.

27 A researcher interested in identifying insulin in body fluids produced the following three HPLC chromatograms. The first is a body fluid sample of interest (X), the second is a reference sample containing insulin (Y) and the third is the same body fluid sample spiked with the insulin reference sample (Z). Consider and explain the strategy used by the researcher to identify insulin in body fluid and conclude whether insulin was found.

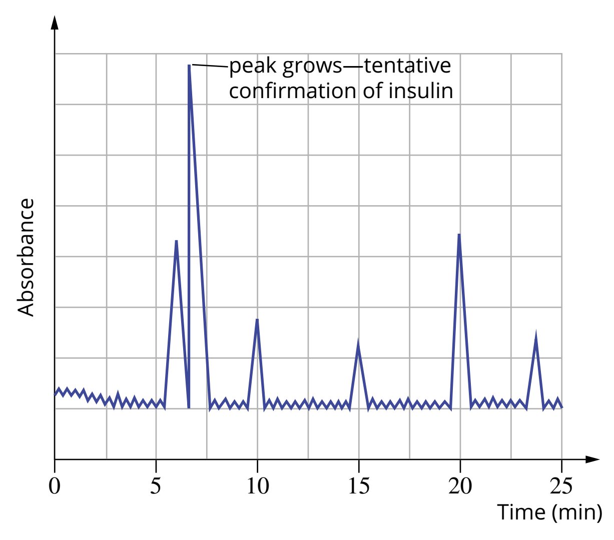
X



Y



Z

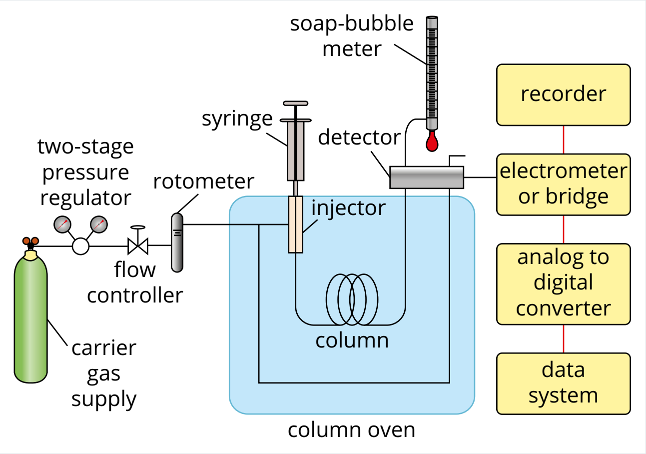


28 The concentration of caffeine in an energy drink was determined by GC. Chromatograms of standards with accurately known concentrations of caffeine were also obtained under the same conditions as the sample.

The peak areas of the sample and the standards are shown in the table. Calculate the concentration of caffeine in the sample.

Caffeine standards ($\mu g\,mL^{-1}$)	**Peak area (mm^2)**
0.010	1.0
0.020	2.4
0.040	4.6
0.060	7.0
sample	3.6

29 Analyse this chromatography equipment. Deduce whether this shows GC or HPLC and explain your reasoning. Identify any components common to both chromatography types.

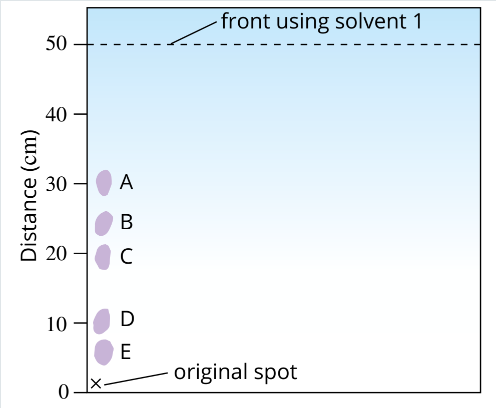


Knowledge utilisation

30 The amino acids present in a sample of fruit juice can be detected by thin-layer chromatography. R_f values of some amino acids in two separate solvents are given in the table below.

Amino acid	Solvent 1 R_f	Solvent 2 R_f
lysine	0.12	0.55
leucine	0.58	0.82
proline	0.39	0.88
valine	0.40	0.74
2-aminobutyric acid	0.28	0.58
threonine	0.21	0.49
hydroxyproline	0.21	0.67
β-phenylamine	0.50	0.86
isoleucine	0.57	0.81
alanine	0.24	0.55
serine	0.19	0.34
glutamic acid	0.25	0.33
glycine	0.20	0.40
arginine	0.13	0.60
taurine	0.12	0.33
tyrosine	0.38	0.62

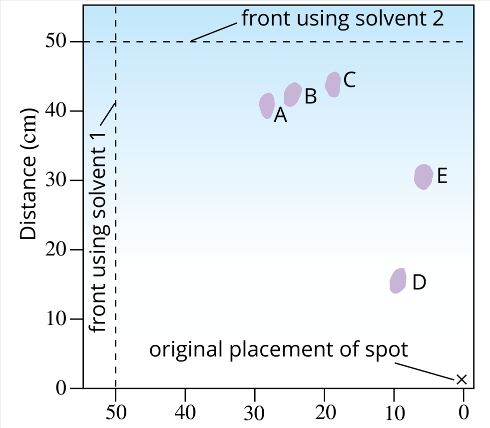
To achieve better separation of the complex mixture of substances present in the juice, a ‘two-way’ chromatogram was prepared. The first step in this procedure was to run a chromatogram using solvent 1. The results of this chromatogram are shown in the following figure.



a Calculate the R_f value of each spot on the chromatogram.

b Determine the amino acids responsible for each spot.

The TLC plate was then turned around anticlockwise so that it lay at a right angle to the original and a second chromatogram was produced using solvent 2. This figure shows the appearance of the TLC plate after some time.

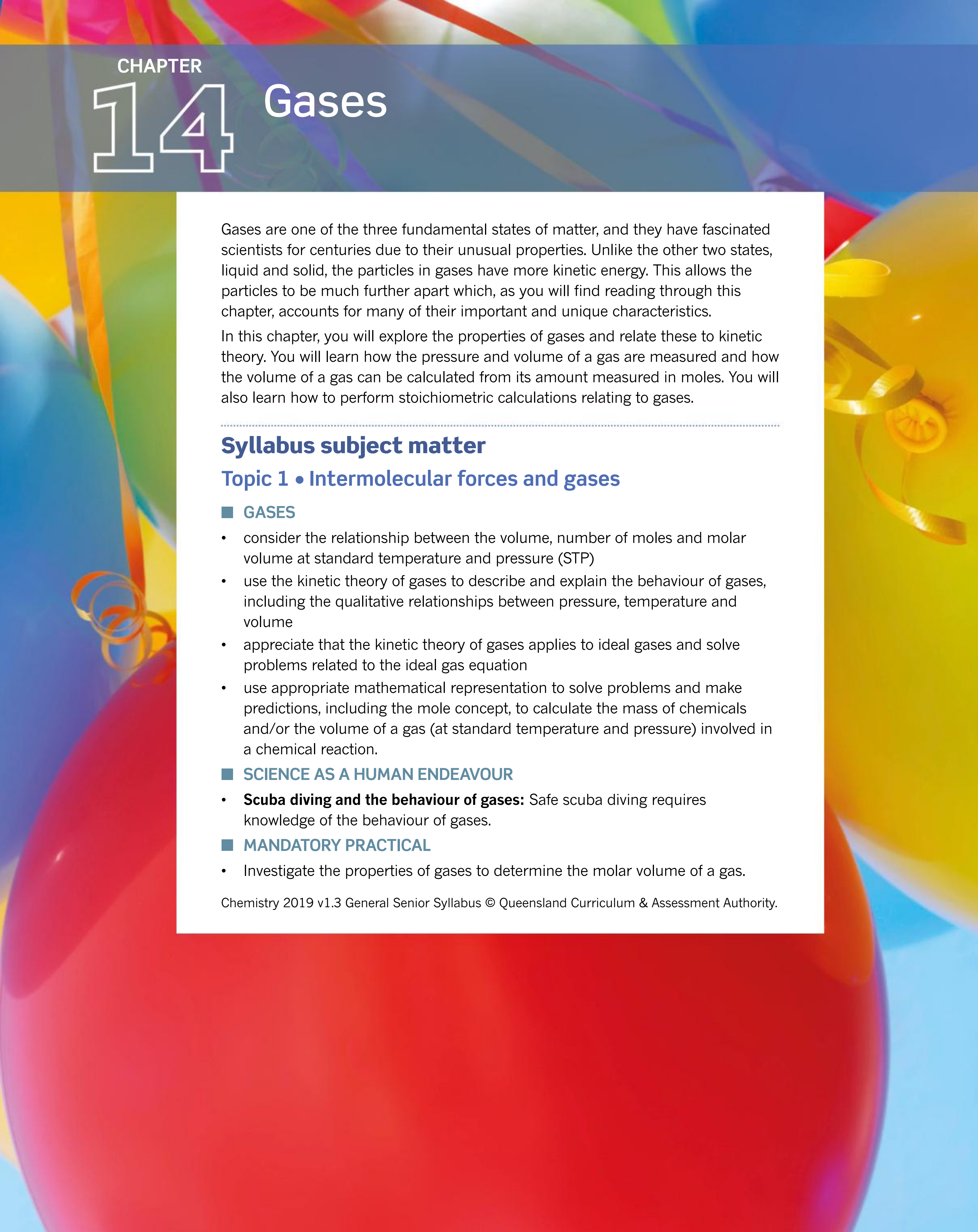


c Use the table to identify each component in the mixture.

d Propose a possible advantage of a two-way chromatogram.

31 The following materials are commonly used in chromatography: water, butanol, paraffin wax, glass sheets, paper strips, powdered alumina plates, nitrogen gas and hexane. Research amino acid identification and choose a suitable stationary phase and mobile phase from the list to use in the analysis of amino acids by thin-layer chromatography.

32 Some polar substances in blood and urine can be detected using a ‘dipstick’ test. The test stick is dipped into the sample and the assay is completed as the water-based fluid is drawn up the stick. Determine how a thin-layer chromatography version could be used to separate the polar compound of interest from unrelated substances in the mixture and then tested for its presence.



CHAPTER

14 Gases

Gases are one of the three fundamental states of matter, and they have fascinated scientists for centuries due to their unusual properties. Unlike the other two states, liquid and solid, the particles in gases have more kinetic energy. This allows the particles to be much further apart which, as you will find reading through this chapter, accounts for many of their important and unique characteristics.

In this chapter, you will explore the properties of gases and relate these to kinetic theory. You will learn how the pressure and volume of a gas are measured and how the volume of a gas can be calculated from its amount measured in moles. You will also learn how to perform stoichiometric calculations relating to gases.

Syllabus subject matter

Topic 1 • Intermolecular forces and gases

GASES

- consider the relationship between the volume, number of moles and molar volume at standard temperature and pressure (STP)
- use the kinetic theory of gases to describe and explain the behaviour of gases, including the qualitative relationships between pressure, temperature and volume
- appreciate that the kinetic theory of gases applies to ideal gases and solve problems related to the ideal gas equation
- use appropriate mathematical representation to solve problems and make predictions, including the mole concept, to calculate the mass of chemicals and/or the volume of a gas (at standard temperature and pressure) involved in a chemical reaction.

SCIENCE AS A HUMAN ENDEAVOUR

- **Scuba diving and the behaviour of gases:** Safe scuba diving requires knowledge of the behaviour of gases.

MANDATORY PRACTICAL

- Investigate the properties of gases to determine the molar volume of a gas.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

14.1 Introducing gases

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- describe the properties of gases
- explain the properties of gases using kinetic theory
- describe the concepts of volume and pressure
- perform calculations to convert units of volume and units of pressure.

Under the right conditions of **pressure** and temperature, most substances can exist as a **gas**. Every day you observe the behaviour of gases (Figure 14.1.1). Air is a mixture of gases and is easily compressed. When the car goes over a bump in the road, the air in the tyres compresses slightly and absorbs the impact of the bump (Figure 14.1.1a). Gases mix readily and, unlike solids and liquids, occupy all available space (Figure 14.1.1b). Volume of a gas increases because of pressure changes as it ascends into the atmosphere (Figure 14.1.1c). Such examples can tell you a great deal about the physical properties of gases—those properties that can be observed and measured without changing the nature of the gas itself.

In this module, you will learn about the properties and behaviour of gases.

FIGURE 14.1.1 (a) Air is used to inflate vehicle tyres. (b) The gases that cause the smell of a freshly brewed cup of coffee rapidly fill an entire room. (c) This weather balloon is only partially inflated when released to collect data.

PROPERTIES OF GASES

The physical properties of gases were introduced in Chapter 10, and each of the examples shown in Figure 14.1.1 can be explained by these properties. Gases assume the volume and shape of their container (Figure 14.1.2), and two gases will mix evenly and completely when placed in the same container. This suggests that the particles move independently of one another. The low **density** of gases, relative to that of liquids and solids, is because the particles in a gas are much more widely spaced apart. This allows gases to be the most compressible form of matter.

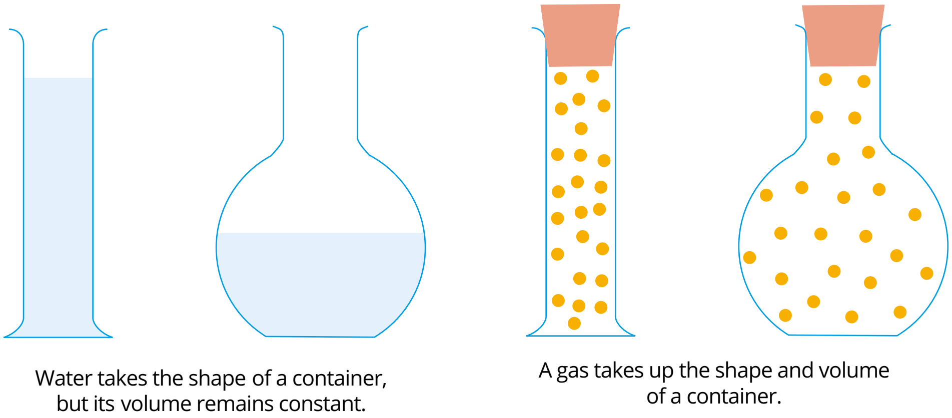


FIGURE 14.1.2 Both liquids and gases take the shape of the container they are in. However, a liquid has a fixed volume, whereas the volume of a gas expands to fill all available space in a container.

Gases assume the volume and shape of their container, are the most compressible form of matter, mix evenly and completely when confined to the same container, and have much lower densities than liquids and solids.

The properties of gases compared to the properties of solids and liquids are summarised in Table 14.1.1. These observations can be used to develop a model of gas behaviour.

TABLE 14.1.1 Some properties of the three states of matter

	Gases	Liquids	Solids
density	low	high	high
volume and shape	fill the space available; particles move independently of one another	fixed volume, adopt shape of container; particles are affected by attractive forces	fixed volume and shape; particles are affected by attractive forces
compressibility	compress easily	almost incompressible	almost incompressible
ability to mix	mix together rapidly	mix together slowly unless stirred	do not mix unless finely divided

KINETIC THEORY

The properties of gases can be explained using **kinetic theory**. Kinetic theory is described in Chapter 10, and states that all matter consists of particles in motion, and as the temperature increases, the movement of the particles increases.

Gases are composed of small particles, either atoms or molecules. These particles are very small compared with the volume of the container, so most of the volume occupied by a gas is actually empty space. This is not the case with liquids and solids. In liquids and solids, the particles are more closely associated with one another and have very little space between them. For this reason, gases have comparatively low density and are much more compressible.

The **intermolecular forces** between gas particles are extremely weak and are easily broken by the kinetic energy of the molecules. As a result, gas particles tend to move in an independent manner. They have a rapid, random and straight-line motion, colliding with one another and the sides of the container. This explains why gases have the property of assuming the shape and volume of their container.

Kinetic energy (KE) is the energy of motion. The kinetic energy of a particle is calculated from the formula:

$$\text{KE} = \frac{1}{2}mv^2$$

where KE is in joules, mass (m) is in grams, and velocity (v) is in metres per second.

In a gas, when particles collide, the kinetic energy of the gas particles can be transferred from one particle to another, but the total kinetic energy remains constant. Therefore, collisions between gas particles are described as **elastic collisions**—collisions where kinetic energy is conserved. According to kinetic theory, as the temperature of the particles increases, so does their motion. This means the particles at a higher temperature have more kinetic energy when they move around and collide with each other and the sides of their container. This influences their volume and pressure.

VOLUME AND PRESSURE OF GASES

There is a relationship between the volume, pressure and temperature of gases. For example, when you blow up a balloon, the balloon expands because you put more air into it and the rubber of the balloon stretches. If you then plunge that balloon into liquid nitrogen, as shown in Figure 14.1.3, the balloon shrinks dramatically. The temperature of the gas decreases, causing the volume to decrease. Pumping more air into a tyre, as shown in Figure 14.1.4, increases the pressure inside the tyre.

FIGURE 14.1.3 A balloon filled with air at room temperature is dipped into liquid nitrogen at −196°C. The volume of the air inside the balloon decreases dramatically.

FIGURE 14.1.4 Pumping more air into a tyre increases the pressure in the tyre because more particles are being pumped into a nearly fixed volume.

Volume

Volume is the quantity used to describe the space that a substance occupies. Because a gas occupies the whole container that it is in, the volume of a gas is equal to the volume of its container.

There are several different units used for volume, some of which are represented in Figure 14.1.5. The common units are litre (L), millilitre (mL), cubic metre (m^3) and cubic centimetre (cm^3). Small volumes of gas are usually measured in millilitres (mL) or litres (L). Very large samples are measured in cubic metres (m^3).

- $1\,mL = 1\,cm \times 1\,cm \times 1\,cm = 1\,cm^3$
- $1\,L = 1 \times 10^3\,mL$ or $1000\,mL$
- $1\,m^3 = 1\,m \times 1\,m \times 1\,m = 100\,cm \times 100\,cm \times 100\,cm = 1 \times 10^6\,cm^3$
- $1\,m^3 = 1 \times 10^6\,cm^3 = 1 \times 10^6\,mL = 1000\,L$

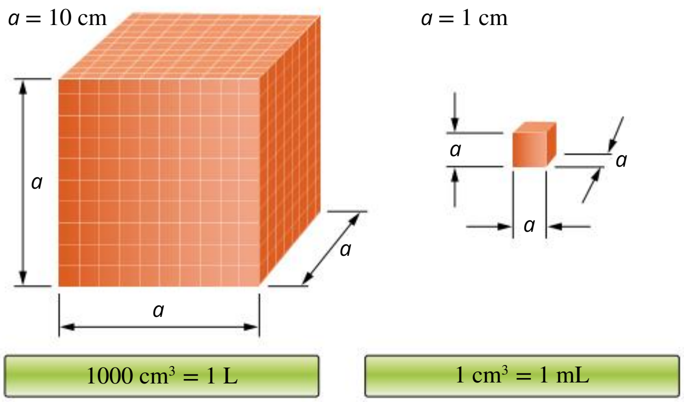


FIGURE 14.1.5 Cubic centimetres, millilitres and litres are the most commonly used units of volume.

SKILLBUILDER 14.1.1

Converting units

In chemistry, sometimes more than one type of unit is used to measure the same thing. For example, pressure is commonly measured in pascals, bars, standard atmospheres or millimetres of mercury. You will find knowing a method to convert units is very helpful and you can use the method presented below to perform the calculations in this chapter.

Step 1

Look up the conversion and write it down.

For example, if you need to convert 2 kilograms (kg) into pounds (lb), you would find that 1 kg is equivalent to 2.2 lb:

$$1.0\,\text{kg} = 2.20\,\text{lb}$$

Step 2

Write the conversion as a fraction, including units. This fraction is called the conversion factor, and should follow the form:

$$\text{conversion factor} = \frac{\text{new unit}}{\text{original unit}}$$

For example:

$$\text{conversion factor} = \frac{2.20\,\text{lb}}{1.0\,\text{kg}}$$

Step 3

Write a multiplication problem with the original unit and the conversion factor. It should have the form:

$$\text{new unit} = \text{original unit} \times \text{conversion factor.}$$

For example:

$$\text{new unit} = 2.0\,\text{kg} \times \frac{2.20\,\text{lb}}{1.0\,\text{kg}}$$

Step 4

Solve the multiplication problem. Keep track of the units and cancel units that are on the top and bottom. If your units do not cancel, you might have put the wrong half of the fraction on top.

For example:

$$\begin{aligned}\text{new unit} &= 2.0\,\text{kg} \times \frac{2.20\,\text{lb}}{1.0\,\text{kg}}\\ &= \frac{2.0\,\text{kg} \times 2.20\,\text{lb}}{1.0\,\text{kg}}\\ &= \frac{2.0 \times 2.20\,\text{lb}}{1.0}\\ &= 4.40\,\text{lb}\end{aligned}$$

Worked example 14.1.1

CONVERTING VOLUME UNITS

A gas has a volume of 255 mL.

Calculate its volume in:

a cubic centimetres (cm^3)

b litres (L)

c cubic metres (m^3)

Thinking	Working
Recall the conversion factors for each of the units of volume. Apply the correct conversion to each situation.	**a** The units of mL and cm^3 are equivalent. $1\text{ mL} = 1\text{ cm}^3$ $255\text{ mL} = 255\text{ cm}^3$ **b** $1000\text{ mL} = 1\text{ L}$ new volume = $255\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}}$ $= 0.255\text{ L}$ **c** $1 \times 10^6\text{ mL} = 1\text{ m}^3$ new volume = $255\text{ mL} \times \frac{1\text{ m}^3}{1 \times 10^6\text{ mL}}$ $= 2.55 \times 10^{-4}\text{ m}^3$

➤ Try yourself 14.1.1

CONVERTING VOLUME UNITS

A gas has a volume of 700 mL.

Calculate its volume in:

a cubic centimetres (cm^3)

b litres (L)

c cubic metres (m^3)

Pressure

Pressure is the force exerted on any surface compared to the area of surface to which the force applies. Gases exert pressure on any surface they come into contact with. This is because the constant motion of the particles results in collisions with that surface.

The more a gas is compressed, the more frequently the gas particles collide with each other and the walls of their container. The increased frequency of collisions with the walls of the container increases the force on the walls of the container. The force per unit area is the pressure.

Air is a mixture of gases including nitrogen, oxygen, carbon dioxide and argon. In air, the nitrogen molecules collide with the walls of a container, exerting a pressure. In a similar way, the oxygen molecules exert a pressure, as do molecules of the other gases present in the mixture. In the gaseous mixture of nitrogen and oxygen shown in Figure 14.1.6, the measured pressure is the sum of the partial pressure of oxygen and the partial pressure of nitrogen.

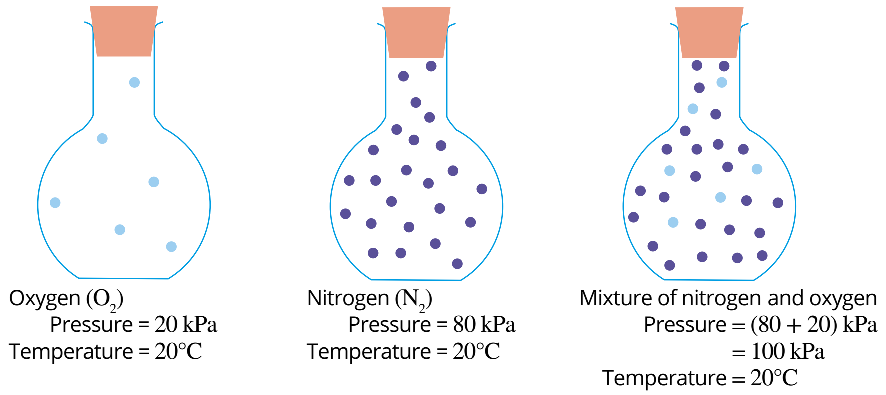


FIGURE 14.1.6 The total pressure of a mixture of gases is the sum of the partial pressures (individual pressures) of each of the gases in the mixture.

Because pressure is the force exerted on a unit area of a surface, the relationship can be written as:

$$\text{pressure} = \frac{\text{force}}{\text{area}} \text{ or } P = \frac{F}{A}$$

The units of pressure depend on the units used to measure force and area. Over the years, scientists in different countries have used different units to measure force and area, so there are a number of different units of pressure.

The SI unit for force is the newton (N) and the SI unit for area is the square metre (m^2). Therefore, the SI units of pressure are newtons per square metre ($N\ m^{-2}$). One newton per square metre is equivalent to a pressure of 1 **pascal** (Pa).

In 1982, the International Union of Pure and Applied Chemistry (IUPAC), the organisation responsible for naming chemicals and setting standards, adopted a standard for pressure equivalent to 100 000 Pa or 100 kPa. This gave rise to a new pressure unit, the **bar**, where 1 bar equals 100 kPa.

The use of mercury barometers resulted in pressure often being measured in millimetres of mercury, or mmHg. Another unit for pressure is the **standard atmosphere** (atm). One standard atmosphere (1 atm) is the pressure required to support 760 millimetres of mercury (760 mmHg) in a mercury barometer at 25°C. This is the average atmospheric pressure at sea level. One atmosphere equals 100 kPa.

The common units of gas pressure are summarised in Table 14.1.2.

TABLE 14.1.2 Common units of gas pressure

Name of unit	Symbol for unit	Conversion to $N\ m^{-2}$
newtons per square metre	$N\ m^{-2}$	
pascal	Pa	$1\ Pa = 1\ N\ m^{-2}$
kilopascal	kPa	$1\ kPa = 1 \times 10^3\ Pa = 1 \times 10^3\ N\ m^{-2}$
atmosphere	atm	$1\ atm = 101.325\ kPa = 1.01325 \times 10^5$
bar	bar	$1\ bar = 100\ kPa = 1.00 \times 10^5\ N\ m^{-2}$
millimetres of mercury	mmHg	$760\ mmHg = 1\ atm = 1.01325 \times 10^5$

These relationships can be used to convert pressure from one unit to another.

Worked example 14.1.2

CONVERTING PRESSURE UNITS

Mount Everest is the highest mountain on Earth.

a The atmospheric pressure at the top of Mount Everest is 0.337 bar. Calculate the pressure in kilopascals (kPa).

b The atmospheric pressure at the top of Mount Everest is 253 mmHg. Calculate the pressure in atmospheres (atm).

c The atmospheric pressure at the top of Mount Everest is 0.333 atm. Calculate the pressure in kilopascals (kPa).

d The atmospheric pressure at the top of Mount Everest is 253 mmHg. Calculate the pressure in bars.

Thinking	**Working**
a To convert bars to kilopascals, use the conversion relationship: 1 bar = 100 kPa To change bar to kPa, multiply the value by 100.	$0.337\,\text{bar} \times \frac{100\,\text{kPa}}{1\,\text{bar}} = 33.7\,\text{kPa}$
b To convert millimetres of mercury to atmospheres, use the relationship: 1 atm = 760 mmHg To change mmHg to atm, divide the value by 760.	$253\,\text{mmHg} \times \frac{1\,\text{atm}}{760\,\text{mmHg}} = 0.333\,\text{atm}$
c To convert atmospheres to kilopascals, use the conversion relationship: 1 atm = 101.3 kPa To change atm to kPa, multiply the value by 101.3.	$0.333\,\text{atm} \times \frac{101.325\ \text{kPa}}{1\,\text{atm}} = 33.7\,\text{kPa}$
d This can be done in two steps. First, convert millimetres of mercury to atmospheres. Use the conversion relationship: 760 mmHg = 1 atm Next, convert atmospheres to bars. Use the conversion relationship: 1 atm = 1.01325 bar	$253\,\text{mmHg} \times \frac{1\,\text{atm}}{760\,\text{mmHg}} \times \frac{1.01325\,\text{bar}}{1\,\text{atm}}$ $= 0.337\,\text{bar}$

➤ Try yourself 14.1.2

CONVERTING PRESSURE UNITS

Cyclone Yasi was one of the biggest cyclones in Australian history.

a The atmospheric pressure in the eye of Cyclone Yasi was measured as 0.902 bar. Calculate the pressure in kilopascals (kPa).

b Calculate the pressure in the eye of Cyclone Yasi in atmospheres (atm) if it is known to be 677 mmHg.

c If the atmospheric pressure in the eye of Cyclone Yasi was 0.891 atm, calculate the pressure in kilopascals (kPa).

d If the atmospheric pressure in the eye of Cyclone Yasi was 677 mmHg, calculate the pressure in bars.

14.1 Review

SUMMARY

- Gases are one of the fundamental states of matter.
- Gases have these physical properties.
  - They have low density.
  - They assume the shape and volume of their container.
  - They compress easily.
  - They mix together rapidly.
- The properties of gases can be understood using kinetic theory.
  - The particles in gases are constantly in motion.
  - The particles move independently, which is why gases assume the shape and volume of their container and mix easily.
  - There is a comparatively large amount of empty space between the particles of gases, which is why they compress easily and are low in density.
- Volume is the quantity used to describe the space that a gas occupies.
- There are several different units used for volume. The common units are litre (L), millilitre (mL), cubic metre (m^3) and cubic centimetre (cm^3).
- To convert between different volume units, use these relationships:

$$1\,L = 1000\,mL = 1000\,cm^3$$

$$1\,m^3 = 1 \times 10^6\,cm^3 = 1 \times 10^6\,mL = 1000\,L$$

- Pressure is the force exerted by a gas on any surface compared to the area of surface to which the force applies. Gases exert pressure on any surface because the constant motion of the particles results in collisions with that surface.
- Because pressure is the force exerted on a unit area of a surface, the relationship can be written as:

$$\text{pressure} = \frac{\text{force}}{\text{area}} \text{ or } P = \frac{F}{A}$$

- The SI unit for pressure is the bar, where 1 bar equals 100 kPa.
- To convert between different pressure units, use these relationships:
  - $1\,bar = 100\,kPa = 1.00 \times 10^5\,Pa$
  - $1\,atm = 760\,mmHg = 101.325 \times 10^5$
    $= 101.325\,kPa = 1.01325\,bar$

KEY QUESTIONS

Retrieval

1 List four physical properties of gases.

2 Define elastic collisions in relation to gases.

3 Define the concept of pressure.

4 Identify if this statement is true or false. The SI unit for pressure is the bar, where 1 bar equals 101.3 kPa.

Comprehension

5 Use the kinetic theory to explain the following observed properties of gases.
  - **a** Gases occupy all the available space in a container.
  - **b** Gases can be easily compressed compared with their corresponding liquid forms.
  - **c** A given volume of a gaseous substance has a mass less than the same volume of the substance in the liquid state.

6 Use the ideas of the kinetic theory of gases to explain the following observations.
  - **a** The pressure in a car's tyres will increase if a long distance is travelled on a hot day.
  - **b** You can smell dinner cooking as you enter your house.
  - **c** A balloon will burst if you blow it up too much.

Analysis

7 Compare the physical properties of gases to those of liquids.

8 Airbags are a safety device fitted to car steering wheels. Determine why a collision with an airbag causes less damage to a driver than a collision with the steering wheel.

9 Pressure is described as the force per unit area of surface. Predict what happens to the pressure in each of the following situations using kinetic theory.
  - **a** The temperature of a filled aerosol can is increased.
  - **b** A gas in a syringe is compressed.

10 Intermolecular forces exist between all gas particles, which should result in the particles moving in a non-random manner—either attracted to or repelled from one another. Determine why it is that gas particles are said to move in an independent manner.

11 Convert the following volumes to the unit specified.
  - **a** 2 L to mL
  - **b** 3.7 m^3 to L
  - **c** 285 mL to m^3

12 Convert each of the following pressures to the units specified.
  - **a** 140 kPa to Pa
  - **b** 4.24 atm to mmHg and Pa
  - **c** 1400 mmHg to atm, Pa and bar

14.2 The ideal gas equation

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain the relationships between the temperature, volume and pressure of gases
- describe the concept of molar volume
- describe standard temperature
- explain the concept of an ideal gas and compare ideal gases to real gases
- state the ideal gas equation and solve problems using this equation.

FIGURE 14.2.1 Robert Boyle (1627–1691) was an Anglo-Irish natural philosopher, chemist, physicist and inventor.

The behaviour of gases has been described qualitatively in Module 14.1. In this module, you will learn about the mathematical relationships that link volume, pressure, temperature and the number of particles of a gas.

These mathematical relationships were developed over a period of several hundred years by different scientists who performed experiments on gases. Robert Boyle, pictured in Figure 14.2.1, showed by experiment that for a given amount of gas at constant temperature, the volume of a gas is inversely proportional to its pressure. The relationship between gas volume and pressure is known as Boyle's law. Figure 14.2.2 shows Jacques Charles, who identified the relationship between the volume and temperature of a gas. In 1787, Charles determined that the volume of a fixed amount of gas is directly proportional to the absolute temperature, provided the pressure remains constant. The relationship between gas volume and temperature is known as Charles' law.

These relationships have become known as the gas laws. The gas laws are used to describe the behaviour of all gases, regardless of their chemical composition.

VOLUME AND PRESSURE

Changing the volume of a fixed amount of gas at constant temperature causes a change in the pressure of the gas. The pressure of the gas in the syringe shown in Figure 14.2.3 increases as the plunger is pushed in and decreases as the plunger is pulled out.

For a given amount of gas at constant temperature, the volume of the gas is inversely proportional to its pressure. This relationship is seen in the changing volume of a weather balloon as it rises to altitudes with much lower pressure than at ground level. A weather balloon filled with helium gas to a volume of 40 L at a pressure of 1 atm increases in volume to 200 L by the time it reaches an altitude with a pressure of 0.2 atm.

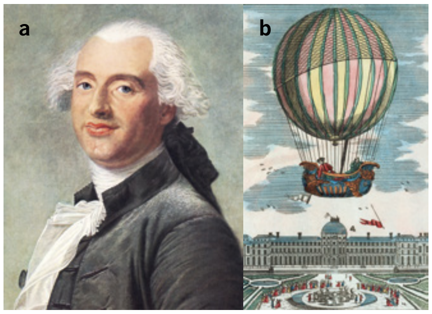


FIGURE 14.2.2 (a) Jacques Charles (1746–1823) was a French chemist, physicist and aeronaut. (b) Charles was famous for making the first flight in a hydrogen balloon on 1 December 1783.

For a fixed amount of gas at a constant temperature, as the pressure increases the volume decreases. The relationship between the volume and pressure of a gas is inversely proportional.

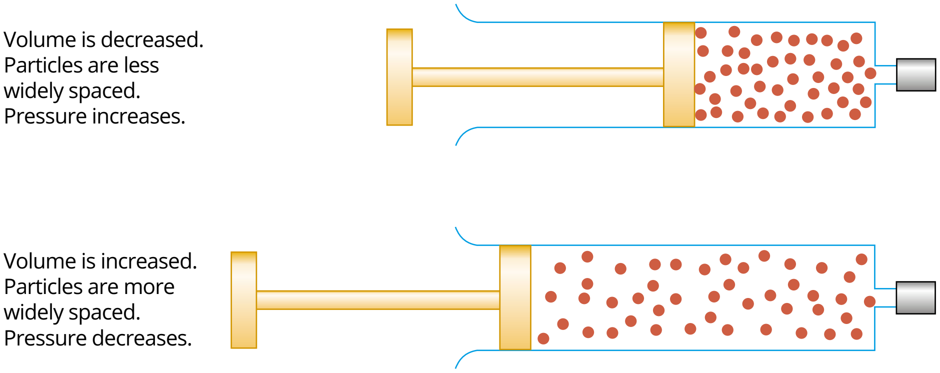


FIGURE 14.2.3 The pressure of the gas in the syringe is affected by a change in volume.

The mathematical relationship between the pressure, P, exerted by a gas and the volume, V, it occupies can be written as:

$$P \propto \frac{1}{V}$$

This relationship can also be expressed as $PV = k$, where k is a constant at a given temperature.

VOLUME AND TEMPERATURE

The kinetic theory states that increasing the temperature of a gas increases the average kinetic energy of the molecules. The molecules move more rapidly and collide with the walls of the container more frequently and with greater force. Increasing the temperature of a gas affects volume and pressure.

- It causes the volume of gas to increase if the pressure on the gas remains the same, such as a gas in a syringe or balloon.
- It causes the pressure to increase if the volume of the gas container is fixed, such as in a sealed flask or a gas cylinder.

Table 14.2.1 shows the results of an experiment that investigates the relationship between temperature and the volume of a given amount of gas. In this experiment, the gas in a syringe is heated slowly in an oven. The pressure on the plunger of the syringe is held constant.

TABLE 14.2.1 Variation of volume with temperature

Temperature (°C)	20	40	60	80	100	120	140
Volume (mL)	60.0	64.1	68.2	72.3	76.4	80.5	84.6

The graph of these results is linear, as shown in Figure 14.2.4. When the graph is extrapolated to a volume of 0 L, it crosses the temperature axis at −273°C. If the origin is reset at −273°C, this graph passes through the origin. This is what led scientists to develop the **Kelvin scale**, described in Chapter 10. On the Kelvin scale, each temperature increment is equal to one temperature increment on the Celsius scale, and 0°C is equal to 273 K.

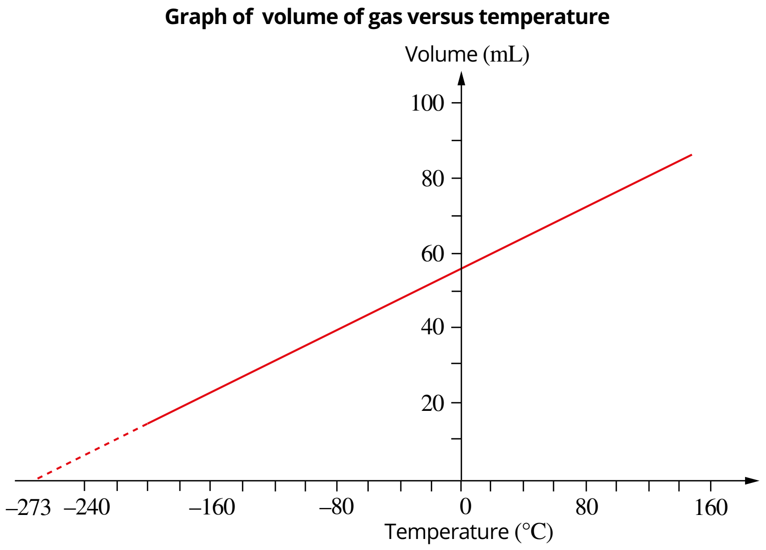


FIGURE 14.2.4 The variation of volume with temperature for a fixed amount of gas at constant pressure

The relationship between temperature on the Celsius scale and temperature on the Kelvin scale is given by the equation:

$$T \text{ (in K)} = T \text{ (in °C)} + 273$$

The temperature 0 K (–273°C) is the lowest temperature theoretically possible. For this reason, 0 K is known as absolute zero. At this temperature, all molecules and atoms have minimum kinetic energy. Therefore, a temperature in degrees Celsius is converted to absolute temperature in kelvin by adding 273. For example, the boiling point of water at sea level is converted to absolute temperature by adding 273, so 100°C is equivalent to 373 K (Figure 14.2.5).

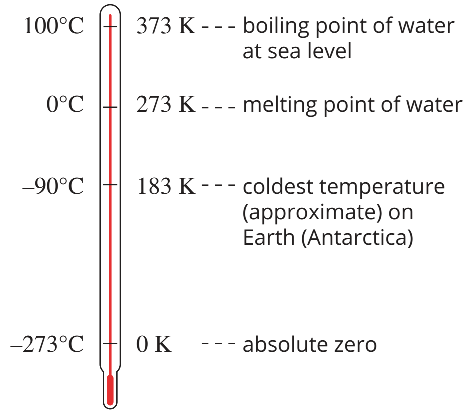


FIGURE 14.2.5 The Celsius and Kelvin temperature scales

FIGURE 14.2.6 Hot air balloonists take advantage of the relationship between the volume of a gas and temperature.

The ascending hot-air balloon in Figure 14.2.6 demonstrates the relationship between volume and temperature. Heating the air inside the balloon causes it to expand, which reduces its density and makes it buoyant. This allows the balloon to rise.

As the temperature, T, of a gas increases, the volume, V, of the gas increases. Therefore, V and T are directly proportional:

$$V \propto T$$

This relationship can also be expressed as $\frac{V}{T} = k$, where k is a constant at a given pressure. Note that the value of T must be expressed in kelvin, not degrees Celsius.

Worked example 14.2.1

CONVERTING TEMPERATURES FROM CELSIUS TO KELVIN

Calculate 300°C on the Kelvin temperature scale.	
Thinking	**Working**
T (in K) = T (in °C) + 273	T (in K) = T (in °C) + 273 = 300 + 273 = 573 K

➤ Try yourself 14.2.1

CONVERTING TEMPERATURES FROM CELSIUS TO KELVIN

Calculate 100°C on the Kelvin temperature scale.

ℹ For a fixed amount of gas at a constant pressure, as the temperature increases the volume increases. The relationship between the volume and temperature of a gas is directly proportional.

VOLUME AND MOLES OF GAS

The volume, V, occupied by a gas also depends directly on the amount of gas, n, in moles. This relationship is shown in Figure 14.2.7. Both syringes show a gas at a constant temperature and pressure. The volume doubles with twice the number of molecules of gas in the syringe.

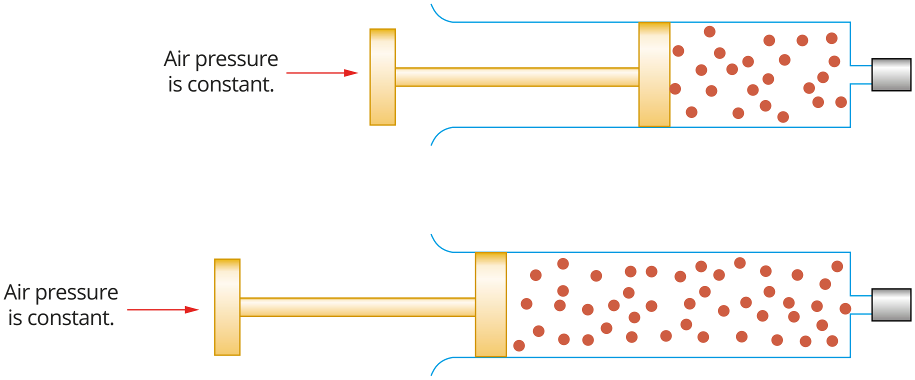


FIGURE 14.2.7 When the amount of gas in the syringe is doubled, the volume doubles, provided the pressure on the plunger and the temperature of the gas remain constant.

The mathematical relationship between the volume, V, occupied by a gas and the amount in moles of gas, n, at constant temperature and pressure can be written as:

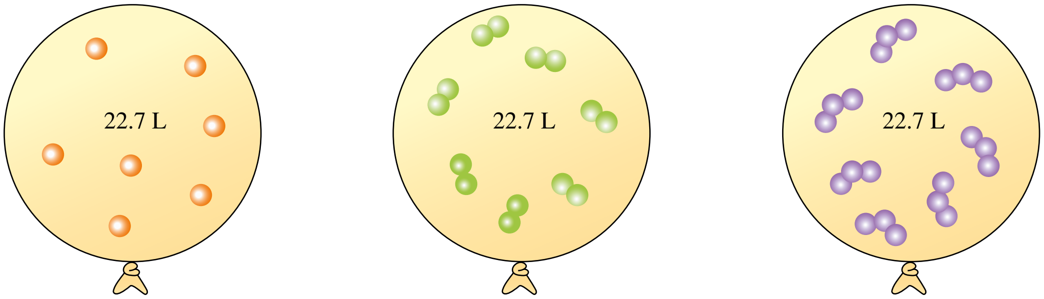
$$V \propto n$$

This relationship can also be expressed as $\frac{V}{n} = k$, where k is a constant at a given temperature and pressure.

If pressure and temperature are constant, as the moles of gas increase the volume increases. The relationship between the number of moles and volume of gas is directly proportional.

Molar volume

If the amount of gas is fixed at 1 mole as shown in Figure 14.2.8, the volume it occupies depends almost entirely on its temperature and pressure. This volume is defined as the **molar volume**, V_m, of a gas. Molar volume is the amount of space, or volume, occupied by 1 mole of any gas at a particular pressure and temperature.



1 mole of neon, Ne
P = 100 kPa
T = 0°C
V = 22.7 L
number of molecules = 6.02×10^{23}
mass = 20.18 g

1 mole of nitrogen, N_2
P = 100 kPa
T = 0°C
V = 22.7 L
number of molecules = 6.02×10^{23}
mass = 28.02 g

1 mole of ozone, O_3
P = 100 kPa
T = 0°C
V = 22.7 L
number of molecules = 6.02×10^{23}
mass = 48.00 g

FIGURE 14.2.8 One mole of any gas occupies the same volume given the same temperature and pressure.

The volume of 1 mole of gas, V_m, is equal to its total volume, V, divided by the number of moles, n, of gas present. This can be represented by the relationship:

$$V_m = \frac{V}{n} \text{ (at a given temperature and pressure)}$$

Or, by rearranging this expression:

$$n = \frac{V}{V_m} \text{ (at a given temperature and pressure)}$$

The molar volume of a gas varies with temperature and pressure. However, molar volume does not vary with the identity of the gas. For example, 1 mole of hydrogen gas occupies the same volume as 1 mole of oxygen gas at the same temperature and pressure.

Standard temperature and pressure

A commonly used set of standard conditions is **standard temperature and pressure (STP)**. STP refers to a temperature of 0°C (273 K) and a pressure of 100 kPa. The molar volume of any gas at STP is 22.7 L mol^{-1}.

$$V_m = 22.7 \text{ L mol}^{-1} \text{ at STP (0°C or 273 K and 100 kPa)}$$

Worked example 14.2.2

CALCULATING THE VOLUME OF A GAS FROM ITS NUMBER OF MOLES

Calculate the volume, in L, occupied by 0.24 mol of nitrogen gas at STP.	
Thinking	**Working**
Rearrange $n = \frac{V}{V_m}$ to make volume the subject.	$n = \frac{V}{V_m}$ $V = n \times V_m$
Substitute in the known values where V_m = 22.7 L mol^{-1} (at STP) and solve.	$V = n \times V_m$ $= 0.24 \times 22.7$ $= 5.448$ L
Consider the units and significant figures. The answer should be given to the smallest number of significant figures in the measurement.	$V = 5.4$ L

➤ Try yourself 14.2.2

CALCULATING THE VOLUME OF A GAS FROM ITS NUMBER OF MOLES

Calculate the volume, in L, occupied by 3.5 mol of oxygen gas at STP.

IDEAL GASES

You have learnt that the properties of gases can be understood using kinetic theory. The gas particles are in constant motion, and as the temperature increases, the movement of the particles increases. Temperature and volume are proportional, while pressure and volume are inversely proportional. All gases respond to changes in temperature and pressure in the same way. These are the assumptions that the gas laws make.

For a gas to behave in a way that is consistent with kinetic theory, the collisions need to be elastic. In elastic collisions no kinetic energy is lost, so this means there cannot be any attraction between the particles or between the particles and the sides of the container. Also, according to kinetic theory, if the temperature is low enough or the pressure high enough, the gas would have no volume. This would mean the particles themselves would also have to have no volume. A gas that obeys kinetic theory and the gas laws is called an **ideal gas**. An ideal gas is one that follows the gas laws at all conditions.

Strictly speaking, an ideal gas does not exist. This is because the collisions of particles can never be truly elastic. There is always some attraction between particles as a result of intermolecular forces. While impossible in theory, many gases behave like ideal gases under most conditions. If you consider that the electrostatic attraction of intermolecular forces gets weaker with increasing distance, and that the particles of gases are further apart than liquids or solids, it means they are less affected by those intermolecular forces. The intermolecular forces are so small, when compared to the kinetic energy of the moving particle, as to be essentially zero.

It is also impossible for gas particles to have no volume. They are made of matter and thus must occupy space. However, when compared to the volume of their container, the volume taken up by gas particles is small enough to be effectively zero. This means under most conditions, such as those found in the laboratory, the environment or at STP, gas particles usually behave as though they have no volume. It is only at very low temperatures, close to absolute zero, or very high pressures, that the particle volume to container volume ratio becomes measurable. Only under these conditions, when gas particles are much closer together than they would normally be, do gases behave in a non-ideal way. Gases that do not obey the gas laws are called **real gases**. Figure 14.2.9 compares the volumes of some example of real gases.

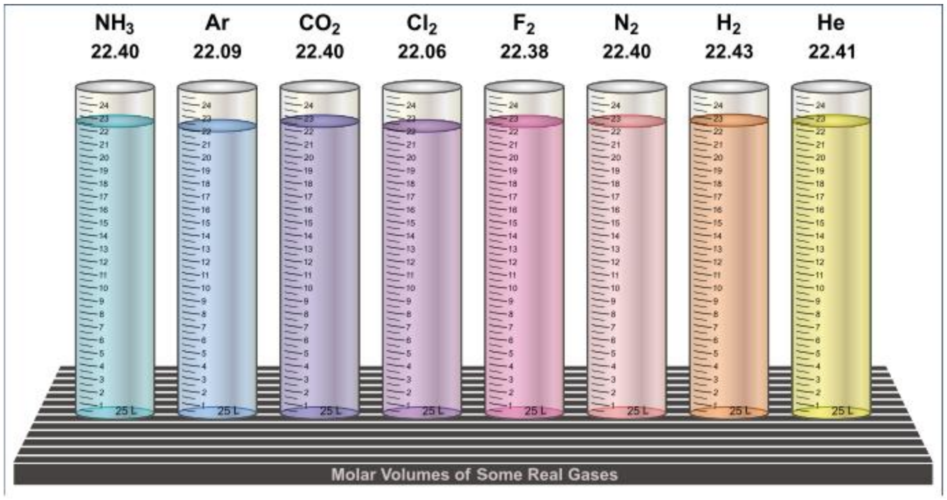


FIGURE 14.2.9 A comparison of the volume of 1 mole of different real gases at the same temperature and pressure (273 K and 100 kPa). An ideal gas would have a volume of 22.7 L at these conditions.

Most gases, particularly at low pressures and high temperatures, behave like ideal gases and obey the gas laws. At very low temperatures and very high pressures, gases are more likely to behave like real gases.

Different gases vary in the range of conditions in which they exhibit ideal behaviour. As you might expect, being the smallest noble gas, Helium comes closest to being an ideal gas at most conditions. Very polar molecules, or larger molecules, can tend to have lower than expected volumes because the intermolecular forces pulling the molecules together are stronger.

Ideal gas equation

The following three laws describe the behaviour of gases under different conditions:

- $V \propto \frac{1}{P}$ (for constant T and n)
- $V \propto T$ (for constant P and n)
- $V \propto n$ (for constant P and T).

All three laws can be combined by writing:

$$V \propto \frac{nT}{P}$$

This relationship can be expressed as $V = \frac{RnT}{P}$, where R is a proportionality constant. The equation is known as the **ideal gas equation** and is more usually written in the form:

$$PV = nRT$$

where R is called the universal gas constant, or simply the **gas constant**.

This constant can be determined experimentally by measuring the volume occupied by a known amount of gas at a known temperature and pressure. The value of R depends on the units of pressure and volume used. It has a value of $8.31\,\mathrm{J\,K^{-1}\,mol^{-1}}$ when:

- P is measured in kilopascals, kPa
- V is measured in litres, L
- n is measured in moles, mol
- T is measured on the Kelvin scale, K.

Worked example 14.2.3

CALCULATING THE VOLUME OF A GAS USING THE UNIVERSAL GAS EQUATION

Calculate the volume, in L, occupied by 2.24 mol of oxygen gas (O_2) if the pressure is 200.0 kPa at 50.0°C.

Thinking	**Working**
Convert units, if necessary. Pressure is in kPa and temperature in K.	$P = 200.0\,\mathrm{kPa}$ (no conversion required) $T = 50.0 + 273$ $= 323\,\mathrm{K}$
Rearrange the universal gas equation so that volume, V, is the subject.	$PV = nRT$ $V = \frac{nRT}{P}$
Substitute values for pressure, amount, temperature and the gas constant, R, then solve for V. Express the answer to the correct number of significant figures.	$V = \frac{2.24 \times 8.31 \times 323}{200.0}$ $= 30.1\,\mathrm{L}$

➤ Try yourself 14.2.3

CALCULATING THE VOLUME OF A GAS USING THE UNIVERSAL GAS EQUATION

PA 2.1.2

Calculate the volume, in L, occupied by 13.0 mol of carbon dioxide gas (CO_2) if the pressure is 250 kPa at 75.0°C.

SCIENCE AS A HUMAN ENDEAVOUR

Decompression chambers

If you swim at the water's surface, your body experiences a pressure, due to the surrounding air, of about 1 atm. Below the surface, your body experiences an additional pressure due to the water. This additional pressure amounts to about 1 atm for every 10 m of depth. Therefore, at 20 m the pressure on your body is about 3 atm.

As the pressure on your body increases, the volume of your body cavities such as your lungs and inner ears decreases. This squeezing effect makes diving well below the water's surface without scuba equipment very uncomfortable. Scuba equipment overcomes this problem by supplying air from tanks to the mouth at the same pressure as that produced by the underwater environment.

As the pressure in a diver's lungs increases during a dive, more gas dissolves in the blood. Nitrogen (N_2) is one of these gases. When a diver ascends, the pressure drops, the nitrogen becomes less soluble in the blood and so comes out of solution. If a diver ascends too quickly, the rapid pressure drop causes the nitrogen to come out of the blood as tiny bubbles (Figure 14.2.10a). This is like the bubbles of carbon dioxide you observe when you open a bottle of soft drink.

These bubbles cause pain in joints and muscles. If they form in the spinal cord, brain or lungs, they can cause paralysis or death. Treatment for divers suffering from this effect (the bends) involves time in a decompression chamber like the one shown in Figure 14.2.10b.

The chamber increases the pressure surrounding the diver's body, forcing any nitrogen bubbles to dissolve in the blood, and then slowly reduces the pressure back to 1 atm. The duration of treatment depends on the severity of the symptoms, the dive history and the patient's response to treatment.

Hyperbaric oxygen therapy (HBOT) is a medical treatment in which patients breathe pure oxygen while inside a decompression chamber at a pressure higher than 1 atm. Its medical uses include improved wound healing by reduction of swelling, infection control and the stimulation of new blood vessel growth. The hyperbaric unit at the Royal Brisbane and Women's Hospital is equipped with hyperbaric chambers used for treating the bends, ulcers, soft tissue infections, carbon monoxide poisoning and wounds that will not heal such as those resulting from diabetes or radiotherapy. HBOT is also used by athletes for faster recovery from soft tissue injuries, to produce sharper performance and to maintain more intense training schedules.

Review

1. Describe HBOT.
2. Explain why it is dangerous for a scuba diver to ascend too quickly.
3. Both temperature and pressure can influence the solubility of gases. Research how the solubility of gases can be affected by temperature and design an experiment that investigates how temperature can change the rate at which soft drink goes flat.

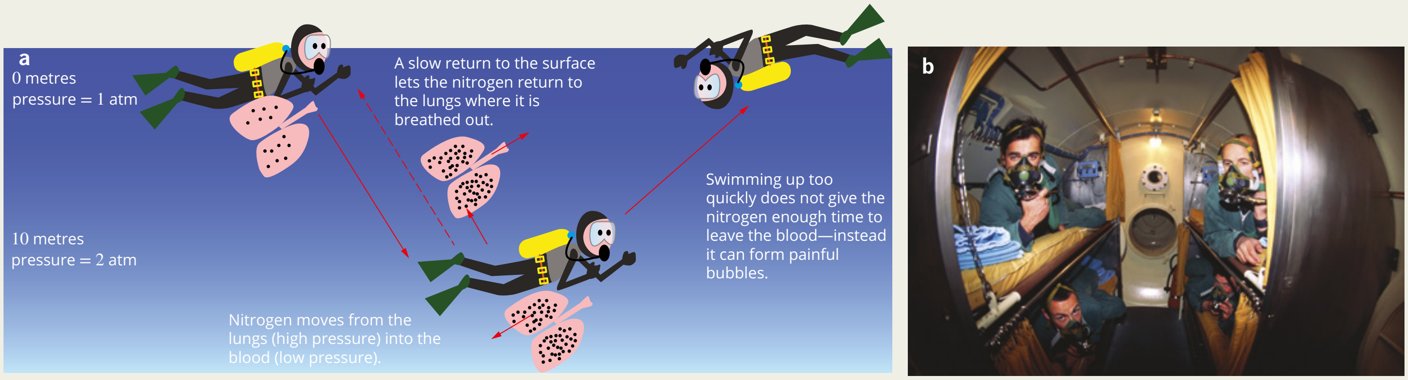


FIGURE 14.2.10 (a) Scuba diving and the bends. (b) A decompression chamber is used to treat divers with the bends.

14.2 Review

SUMMARY

- For a given amount of gas at constant temperature, the volume of the gas is inversely proportional to its pressure.
- The mathematical relationship between the pressure, P, exerted by a gas and the volume, V, it occupies can be written as $PV = k$, where k is a constant at a given temperature.
- For a given amount of gas at constant pressure, the volume of the gas is directly proportional to its temperature.
- The mathematical relationship between temperature, T, and volume, V, can also be expressed as $\frac{V}{T} = k$, where k is a constant at a given pressure. Note that the value of T must be expressed in kelvin, not degrees Celsius.
- For a gas at constant pressure, the volume of the gas is directly proportional to the amount of gas in moles.
- The mathematical relationship between the volume, V, occupied by a gas and the moles of gas, n, at constant temperature and pressure can be written as $\frac{V}{n} = k$, where k is a constant at a given temperature and pressure.
- Molar volume is the amount of space, or volume, occupied by 1 mole of any gas at a particular pressure and temperature.
- An ideal gas is a hypothetical gas whose molecules occupy negligible space and have no interactions, and which consequently obeys the gas laws exactly. Gases that do not obey the gas laws are called real gases.
- Standard temperature and pressure (STP) refers to a temperature of 0°C (273 K) and a pressure of 100 kPa.
- Absolute zero is a temperature of −273°C or 0 K. Molecules and atoms have minimum kinetic energy at this temperature.
- The value of V_m is 22.7 L mol^{-1} at STP.
- $PV = nRT$ is the universal gas equation. It can be used to calculate one variable (P, V, n or T) when the other three variables are known.
- R is the universal gas constant and its value is 8.31 J K^{-1} mol^{-1} when pressure is in kPa, volume is in L, amount is in mol and temperature is in K.

KEY QUESTIONS

Retrieval

1 Describe the relationship between the volume and pressure of a gas, assuming temperature is constant.

2 State the mathematical relationship between temperature and volume.

3 Describe the concept of molar volume.

4 Define the term ‘ideal gas’.

5 Identify if this statement is true or false. In the ideal gas equation, the value of R is independent of the units of pressure and volume used. It always has a value of 8.31 J K^{-1} mol^{-1}.

Comprehension

6 When the gas inside a syringe is compressed, the volume decreases and the pressure increases. Explain why, when the gas inside a syringe is compressed, the pressure of the gas inside increases.

7 Explain, using your understanding of the relationship between the temperature and volume of gases, how a hot air balloon floats.

8 The molar volume of a gas varies with temperature and pressure. Explain why the molar volume of a gas does not vary with the identity of the gas.

Analysis

9 When a weather balloon rises in the atmosphere, its volume increases. However, at higher altitudes there is a drop in temperature. Determine why the balloon does not decrease in volume.

10 Compare the concept of an ideal gas to a real gas.

11 Convert the following Celsius temperatures to absolute temperatures.

- **a** 100°C
- **b** 175°C
- **c** −145°C

12 Calculate the volume of the following gases at STP.

- **a** 1.4 moles of chlorine (Cl_2)
- **b** 1.0×10^{-3} moles of hydrogen (H_2)
- **c** 1.4 g of nitrogen (N_2)

13 Calculate the mass of the following gas samples. All volumes are measured at STP.

- **a** 2.80 L of neon (Ne)
- **b** 50.0 L of oxygen (O_2)
- **c** 140 mL of carbon dioxide (CO_2)

14 A flask of volume 5.0 L at a temperature of 5°C contains 0.25 moles of nitrogen. Determine the pressure in the flask.

15 Determine the volume of gas, in litres, occupied by:

- **a** 0.20 moles of hydrogen at 115 kPa and 40°C
- **b** 12.5 moles of carbon dioxide at 5 atm and 150°C
- **c** 8.50 g of hydrogen sulfide (H_2S) at 100 kPa and 27°C.

16 Calculate the mass of helium in a balloon if the volume is 100 L at a pressure of 95 000 Pa and a temperature of 0°C.

17 At a given temperature, a sample of nitrogen, of mass 11.3 g, exerts a pressure of 102 kPa in a gas cylinder of volume 10.0 L. Calculate the temperature of the gas.

18 Determine which sample of gas contains the greater amount of particles, in moles: 3.2 L of nitrogen at 25°C and a pressure of 1.2 bar or 2.5 L of helium at 23°C and a pressure of 1.2 atm.

14.3 Gas stoichiometry

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain that stoichiometric calculations are based on the law of conservation of mass
- use stoichiometry to perform mass–volume calculations
- use stoichiometry to perform volume–volume calculations
- determine the excess reagent for reactions involving gases.

FIGURE 14.3.1 It is useful to be able to calculate the volume of carbon dioxide emissions produced by the combustion of a particular fuel.

In Chapter 9, you learnt that stoichiometric calculations allow you to use the mole ratio to predict the amount of product that will be formed or how much reactant will be used. Measuring and predicting quantities is a very important part of chemistry. For example, scientists can compare fuels by calculating the quantities of chemicals involved in their reactions. The ethanol produced in the plant shown in Figure 14.3.1 can be compared to petroleum by calculating how much oxygen is required for complete combustion. This allows the amount of atmospheric pollution a fuel will produce to be determined by calculating the mass of carbon dioxide given off.

In this module, you will apply stoichiometry to chemical reactions involving gases.

STOICHIOMETRY

Calculations that involve the use of the mole concept, combined with an understanding of chemical equations, are called stoichiometric calculations. **Stoichiometry** is the study of ratios of the number of moles of substances.

Stoichiometric calculations are based on the law of conservation of mass, according to which the total mass of all products is equal to the total mass of all reactants. Another way of expressing this idea is that, in a chemical reaction, atoms are neither created nor destroyed. Consequently, given the amount of one substance involved in a chemical reaction and a balanced equation for the reaction, you can calculate the amounts of all other substances involved.

The steps and details of formulas used when performing stoichiometric calculations involving gases are summarised in Figure 14.3.2.

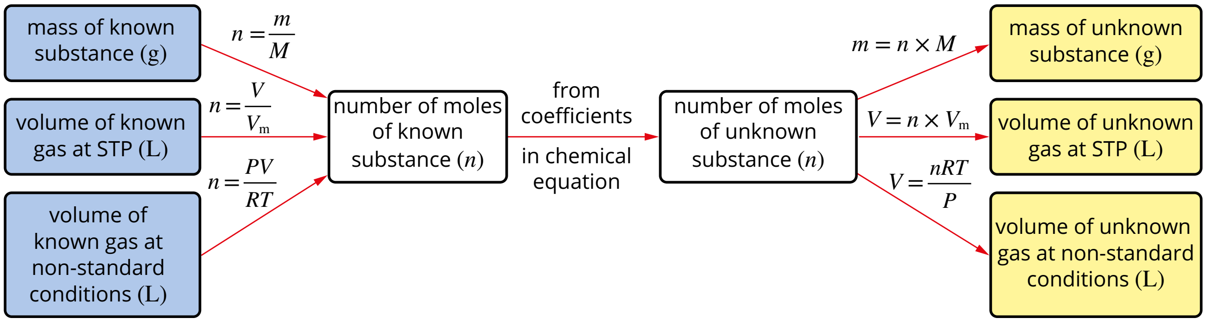


FIGURE 14.3.2 Most stoichiometric calculations follow the steps shown in the flow chart. Calculating the number of moles and using a mole ratio from a balanced chemical equation are always central to any stoichiometric calculation.

Mass–volume stoichiometry

Some stoichiometric calculations require you to determine the volume of a gas produced from a given mass of reactant. For these calculations, you need to determine the number of moles of the reactant using its mass, and use the mole ratio from the balanced chemical equation.

The volume of the gas, in litres, is determined from its number of moles using the formula:

1 At STP (0°C and 100 kPa), the molar volume equation is used:

$$n = \frac{V}{V_m}$$

This formula can be rearranged to make volume the subject:

$$V = n \times V_m$$

Remember that at STP, the accepted volume of 1 mole of any gas is 22.7 L mol^{-1}.

2 At non-standard conditions, the ideal gas equation is used:

$$PV = nRT$$

The ideal gas equation can also be rearranged to make volume the subject:

$$V = \frac{nRT}{P}$$

Remember that for use in the ideal gas equation, pressure must be in kPa, temperature must be in K and the gas constant is equal to 8.31 J K^{-1} mol^{-1}.

Worked example 14.3.1

MASS–VOLUME STOICHIOMETRIC CALCULATIONS AT STANDARD TEMPERATURE AND PRESSURE (STP)

Calculate the volume of carbon dioxide, in L, produced when 2.00 kg of propane is burnt completely in oxygen. The gas volume is measured at STP.

Thinking	**Working**
Write a balanced equation for the reaction.	$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
Molar mass of propane (C_3H_8) = 44.11 g Calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(C_3H_8) = \frac{2000}{44.11}$ $= 45.3$ mol
From the balanced equation determine the mole ratio of carbon dioxide formed from propane burnt: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{3}{1} \times 45.3$ $= 135.9$ mol
Calculate the volume of the unknown substance using: $V = n \times V_m$	$V(CO_2) = 135.9 \times 22.7$ $= 3084.93$ L $= 3.08 \times 10^3$ L

➤ Try yourself 14.3.1

MASS–VOLUME STOICHIOMETRIC CALCULATIONS AT STANDARD TEMPERATURE AND PRESSURE (STP)

Calculate the volume of carbon dioxide, in L, produced when 300 g of butane is burnt completely in oxygen. The gas volume is measured at STP.

Worked example 14.3.2

MASS–VOLUME STOICHIOMETRIC CALCULATIONS AT NON-STANDARD CONDITIONS

Calculate the volume of carbon dioxide, in L, produced when 800 g of propane is burnt completely in oxygen. The gas volume is measured at 60°C and 200 kPa.

Thinking	**Working**
Write a balanced equation for the reaction.	$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
Calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(C_3H_8) = \frac{800}{44.11}$ $= 18.1\text{ mol}$
From the balanced equation determine the mole ratio of carbon dioxide formed from propane burnt: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1}$
Calculate the number of moles of the unknown substance using: n(unknown) = mole ratio × n(known)	$n(CO_2) = \frac{3}{1} \times 18.1$ $= 54.3\text{ mol}$
Express the pressure and temperature in required units.	$P = 200\text{ kPa}$ $T = 60 + 273$ $= 333\text{ K}$
Calculate the volume of the unknown substance using: $V = \frac{nRT}{P}$	$V(CO_2) = \frac{54.3 \times 8.31 \times 333}{200} = 751\text{ L}$

➤ Try yourself 14.3.2

MASS–VOLUME STOICHIOMETRIC CALCULATIONS AT NON-STANDARD CONDITIONS

Calculate the volume of carbon dioxide, in L, produced when 5.00 kg of butane is burnt completely in oxygen. The gas volume is measured at 40°C and 400 kPa.

Volume–volume stoichiometry

For chemical reactions where both the reactants and products are in the gaseous state, it is often convenient to measure volumes, rather than masses. For example, the reaction between propane gas and oxygen can be represented by the equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

This equation tells us that when 1 mole of propane gas reacts with 5 moles of oxygen gas, 3 moles of carbon dioxide gas and 4 moles of water vapour is produced.

You saw in Module 14.2 that if gases are behaving in an ideal way, the molar volume of any gas is the same. Therefore, the mole ratios in the balanced equation become volume ratios at the same temperature and pressure. In the above reaction, this means that when 1 litre of propane gas reacts with 5 litres of oxygen gas, 3 litres of carbon dioxide gas and 4 litres of water vapour is produced.

Worked example 14.3.3

GAS VOLUME–VOLUME CALCULATIONS

Methane gas (CH_4) is burnt in a gas stove according to the following equation: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ If 50 mL of methane is burnt, calculate the volume of O_2 gas required for complete combustion of the methane under constant temperature and pressure conditions.	
Thinking	**Working**
Use the balanced equation to find the mole ratio of the two gases involved.	1 mole of CH_4 gas reacts with 2 moles of O_2 gas.
The temperature and pressure are constant, so volume ratios are the same as mole ratios.	1 volume of CH_4 reacts with 2 volumes of O_2 gas, so 50 mL of CH_4 reacts with 100 mL of O_2.

➤ Try yourself 14.3.3

GAS VOLUME–VOLUME CALCULATIONS

Methane gas (CH_4) is burnt in a gas stove according to the following equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

If 50 mL of methane is burnt in air, calculate the volume of CO_2 gas produced under constant temperature and pressure conditions.

Calculations involving excess reactants

As you learnt in Chapter 9, stoichiometry calculations become more complex if the reactants are not present in their stoichiometric ratio. In these cases, you must determine which reactant is completely consumed in the reaction, the limiting reactant, and which one is present in excess. The amount of limiting reactant determines how much product is formed.

Worked example 14.3.4 introduces a strategy that can be used to determine the limiting reactant in a reaction.

Worked example 14.3.4

EXCESS REACTANT CALCULATIONS

A gaseous mixture of 25.0 g of hydrogen gas and 100.0 g of oxygen gas is ignited. The water produced is collected and weighed. Calculate the theoretical mass of water produced. The equation for the reaction is:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

a Determine which reactant is the limiting reactant.

b Calculate the mass of water vapour formed.

Thinking	**Working**
a Calculate the number of moles of each reactant using $n=\frac{m}{M}$ or $n=\frac{V}{V_m}$ as appropriate.	$n(H_2) = \frac{m}{M}$ $= \frac{25.0}{2.02}$ $= 12.4$ mol $n(O_2) = \frac{m}{M}$ $= \frac{100.0}{32.00}$ $= 3.13$ mol
Use the coefficients of the equation to determine the limiting reactant.	The equation shows 2 moles of H_2 reacts with 1 mole of O_2. So to react all of the O_2 will require: $2 \times n(O_2)$ of H_2 $= 2 \times 3.13$ $= 6.26$ mol As there are 12.4 moles of H_2, the H_2 is in excess. The O_2 is the limiting reactant (it will be completely consumed).
b Determine the mole ratio using: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$ The limiting reactant is the known substance.	$\frac{n(H_2O)}{n(O_2)} = \frac{2}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(H_2O) = 2 \times 3.13$ $= 6.26$ moles
Calculate the required quantity of the unknown using $m = n \times M$ or $n=\frac{V}{V_m}$ or $PV = nRT$ as appropriate.	$m(H_2O) = 6.26 \times 18.02$ $= 113$ g

➤ Try yourself 14.3.4

EXCESS REACTANT CALCULATIONS

WS 2.1.4

Calculate the volume of carbon dioxide, in L, produced when 65.0 g of butane is burnt completely in 200 L of oxygen. The gas volume is measured at STP. The equation for the reaction is:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

a Determine which reactant is the limiting reactant.

b Calculate the volume of carbon dioxide formed.

14.3 Review

SUMMARY

- Stoichiometry is the study of ratios of the amount of moles of substances.
- A balanced equation shows the ratio of the amount, in moles, of reactants and products in the reaction.
- Stoichiometric calculations follow these general steps.
  1. Calculate the number of moles of a known substance from the data given.

     Use $n = \frac{m}{M}$, $n = \frac{V}{V_m}$ or $n = \frac{PV}{RT}$.
  2. Use the mole ratio from a balanced chemical equation to determine the amount, in moles, of the unknown substance.

     $$\frac{n(\text{unknown chemical})}{n(\text{known chemical})} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$
  3. Find the desired quantity of the unknown substance from its amount, in moles, using:

     $m = n \times M$, $V = n \times V_m$ or $PV = nRT$
- To solve mass–volume calculations, follow these steps.
  1. Determine the number of moles of the reactant using its mass.
  2. Use the mole ratio from the balanced chemical equation.
- To solve volume–volume calculations, the mole ratios in the balanced equation become volume ratios at the same temperature and pressure.
- For calculations involving excess reactants, follow these steps.
  1. Calculate the number of moles.
  2. Use coefficients to find the limiting reactant.

KEY QUESTIONS

Retrieval

1 Define stoichiometry.

2 List three ways the number of moles of a known substance can be calculated.

Comprehension

3 Explain how mass volume calculations are solved.

4 Describe why it is possible to use mole ratios in the balanced equation as volume ratios.

Analysis

5 Propane (C_3H_8) burns in oxygen according to the equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

When the following masses of propane react completely with excess oxygen, calculate the volume of:

i oxygen used at STP

ii carbon dioxide produced at STP

a 22 g

b 3.40 kg

6 Octane is one of the main constituents of petrol. It burns according to the equation:

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

Calculate the mass of octane that must have been used if 50.0 L of carbon dioxide, measured at 120°C and 1.10 atm, was produced.

7 Hydrogen peroxide decomposes according to the equation:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

Calculate the volume of oxygen, collected at 30.0°C and 91.0 kPa, that is produced when 10.0 g of hydrogen peroxide decomposes.

8 Determine the volume of NO_2 produced when 0.5 L of nitrogen(II) oxide (NO) reacts with excess oxygen. (All volumes are measured at 25°C and 100 kPa.) The equation for the reaction is:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

9 Calculate the volume of oxygen needed to completely react with 150 mL of carbon monoxide according to the following equation. Assume all volumes are measured at the same temperature and pressure.

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

Determining the molar volume of hydrogen

Research and planning

Aim

To determine the molar volume of hydrogen gas at STP, 0°C and 100 kPa

Rationale (scientific background to the experiment)

Gases are produced when fuels burn. An understanding of the behaviour of gases and gas laws allows us to calculate the volume of gaseous products and compare volumes of greenhouse gases released by different fuels.

In this experiment, you will determine the number of moles of hydrogen gas produced in a reaction. From measurements of the gas volume and pressure, the molar volume of hydrogen at standard temperature and pressure (STP) can be calculated.

Timing

40 minutes

Materials

- 20 mL 2 M HCl
- 4.5 to 5 cm length of magnesium ribbon that has a mass of no more than 0.08 g
- 100 mL gas syringe
- set of apparatus to clamp the syringe to a retort stand
- 250 mL conical flask
- one-hole stopper to fit conical flask
- 4 cm length of glass tubing to fit the one-hole stopper
- approx. 50 cm length of rubber tubing to connect the gas syringe to the glass tubing in the one-hole stopper
- 100 mL measuring cylinder
- electronic balance
- emery paper or steel wool for cleaning the magnesium ribbon
- safety glasses

PRE-LAB SAFETY INFORMATION

Material used	Hazard	Control
2 M HCl	Toxic by all routes of exposure; lung irritation	Wear eye and skin protection

Please indicate that you have understood the information in the safety table.

Name (print): ______________________________

I understand the safety information (signature): ______________

Method

Risk assessment

Consideration of risks includes chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and Assessment book or download it from your eBook.

1. Clamp the stoppered gas syringe to its retort stand and connect the conical flask and syringe using the rubber tubing as shown in Figure 14.4.1. Check that the equipment is secure.

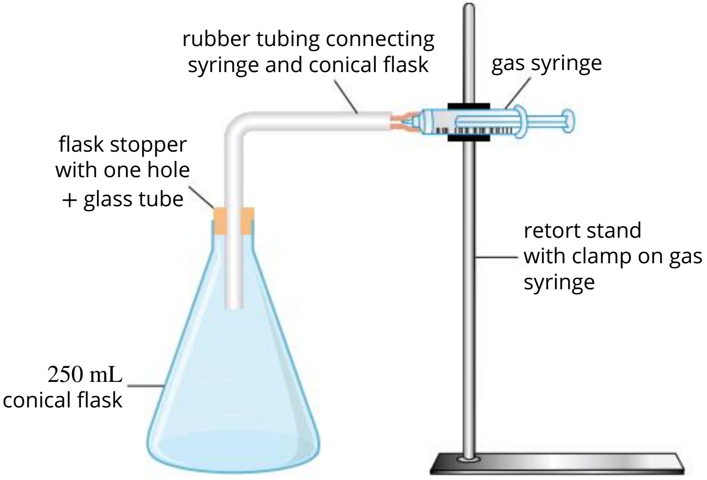


FIGURE 14.4.1 Experimental set-up

2. Remove the stopper from the conical flask and carefully pour about 15 mL of 2 M hydrochloric acid into the flask without touching the sides.
3. Clean and accurately weigh the magnesium ribbon, making sure that it weighs no more than 0.08 g.
4. Tilt the flask and carefully place the magnesium ribbon on a dry side of the flask making sure that the magnesium does not contact the acid. Replace the stopper tightly, still keeping the flask tilted.
5. Carefully withdraw the plunger of the syringe and then release it. If the system has no leaks, the plunger will return to its original position. Once any leaks have been fixed record the initial volume shown on the syringe in results Table 14.4.1.
6. Straighten the conical flask and shake the piece of magnesium into the acid. As gas fills the syringe, rotate the plunger gently to prevent it from sticking.
7. Once the magnesium has been used up, allow the conical flask to cool. In Table 14.4.1, record the final volume of gas in the syringe when the plunger has completely stopped moving. Calculate and record in the table the increase in the volume of gas in the syringe.

8 Record the room temperature and atmospheric pressure.

9 From data supplied by your teacher, record the water vapour pressure at room temperature. To allow for the pressure of water vapour in the syringe, subtract the water vapour pressure from atmospheric pressure.

Variables

i Independent: the mass of magnesium

ii Dependent: the volume of hydrogen

iii Controlled: the temperature, the atmospheric pressure, water vapour pressure at measured temperature

Analysing

Raw data

1 Record your data in the table below.

TABLE 14.4.1 Experimental results

Results	
mass of magnesium (g)	
initial volume of syringe (mL)	
final volume of syringe (mL)	
volume of gas (mL)	
temperature (°C)	
atmospheric pressure, P_{atm} (kPa)	
water vapour pressure at measured temperature, P_{vapour} (kPa)	
pressure of H_2 gas = $P_{atm} - P_{vapour}$	

Processed data

2 Write a balanced equation for the reaction between magnesium and hydrochloric acid.

3 Calculate the amount, in mol, of magnesium reacted.

4 Using the amount of magnesium reacted, calculate the number of mol of H_2 produced.

Reflect and check that your data analysis demonstrates these characteristics

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.

☐ Accurate application of algorithms, visual and graphical representations of data is demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

Analysis

5 The amount of H_2 gas occupied the measured volume at the experimental pressure and temperature. Calculate the volume occupied by one mole (molar volume) at these experimental conditions.

6 Convert your answer to Question **5** to molar volume at standard temperature and pressure conditions (STP), 0°C and 100 kPa.

7 How close is your result for the molar volume of hydrogen to the accepted molar volume of an ideal gas?

8 **a** List each piece of equipment used in this experiment that had error associated with it. State the error value associated with each.

  b Explain how this value may have had an effect on the determination for the molar volume you have calculated.

9 Explain why the molar volume of any gas at specified conditions, such as STP, is the same.

10 Using your measurements for the volume, temperature and pressure of the hydrogen gas sample and the ideal gas equation, calculate the mole of H_2 gas produced. Compare your answer with your answer to Question **4**; how well do the two answers agree?

Reflect and check that your analysis demonstrates these characteristics

☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.

☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.

☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Interpreting and communicating

Conclusion

1 **a** State your conclusion.

  b Describe evidence you have collected that leads you to this conclusion.

Evaluation

2 **a** Considering your analysis and conclusion, summarise whether the experiment provided an effective and efficient method of determining the molar volume of hydrogen.

  b Describe whether the level of uncertainty that was calculated in Question **8** was reasonable. (Provide values that were close to whole number ratios.)

Improvements

3 If you were to repeat the experiment, identify the steps you would do differently. You should address these questions in your answer.

  a How would you change the methodology and how might this improve the results?

  b How well did you perform the tasks and are there skills that you need to improve on in your technique?

  c How could the collection of data be improved or uncertainty reduced?

Reflect and check that your evaluation demonstrates these characteristics

☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis).
☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respect to the error analysis and limitations or sufficiency of the data.
☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment, which are logically derived from the analysis of the evidence.

Chapter review

WS 2.1.5 WS 2.1.6

KEY TERMS

bar
density
elastic collision
gas
gas constant
hyperbaric oxygen therapy (HBOT)
ideal gas
ideal gas equation
intermolecular force
Kelvin scale
kinetic energy
kinetic theory
molar volume
pascal
pressure
real gas
standard atmosphere
standard temperature and pressure (STP)
stoichiometry
volume

KEY QUESTIONS

Retrieval

1 Identify which of the following volumes is equal to 4.5 L.
 A 4.5×10^{-3} mL
 B 4.5×10 mL
 C 4.5×10^{2} mL
 D 4.5×10^{3} mL

2 Identify which one of the following statements best describes the effect of an increase in temperature on gas particles.
 A Both the average kinetic energy and the average speed of the particles increase.
 B Both the average kinetic energy and the average speed of the particles decrease.
 C The average kinetic energy of the particles decreases and the average speed of the particles increases.
 D The average kinetic energy of the particles increases and the average speed of the particles decreases.

3 Select the correct answers from the pairs of words to complete the paragraph about gases.
The volume occupied by the atoms or molecules in a gas is much *smaller/larger* than the total volume occupied by the gas. The particles move in rapid, *straight-line/curved* paths and collide with each other and with the walls of the container. The forces between particles are extremely *weak/strong*. The collisions between particles are *elastic/rigid*. The average kinetic energy of the particles is *directly/inversely* proportional to the temperature of the gas, in units of *K/°C*.

4 State the ideal gas equation.

5 State the value of the gas constant, *R*, assuming pressure is measured in kPa, volume in L, and temperature in K.

6 State the name of the law that stoichiometric calculations are based on.

Comprehension

7 Use the kinetic theory of gases to explain the following.
 a The pressure of a gas increases if its volume is reduced at constant temperature.
 b The pressure of a gas decreases if its temperature is lowered at a constant volume.
 c In a mixture of gases, the total pressure is the sum of the partial pressure of each gas.
 d The pressure of a gas, held at constant volume and temperature, will increase if more gas is added to the container.

8 Explain how the relationship between the temperature and volume of gases led to the development of the Kelvin scale.

9 Most gases, under most conditions, behave in an ideal way. Explain under what circumstances gases tend not to follow the gas laws.

Analysis

10 A weather balloon filled with helium gas has a volume of 20 L at 1 atm. After being released, it rises into the atmosphere where it reaches an altitude with a pressure of 0.2 atm. The balloon increases in volume to 100 L. Determine why this occurs in terms of the relationship between the pressure and volume of gases.

11 The molar volume of a gas varies with temperature and pressure. However, molar volume does not vary with the identity of the gas. Consider and explain why this is the case.

12 Convert the following values to the units specified.
 a 1.7 L to mL
 b 4.35 atm to kPa
 c 502 kPa to bar

13 Calculate the mass of oxygen (O_2) present in a 50.0 L container of oxygen at STP.

14 Use the molar volume of a gas at STP to determine the:
 a volume occupied by 8.0 g of oxygen at STP
 b mass of nitrogen dioxide present in 10.0 L at STP

15 Calculate the volume, in litres, occupied by 10.0 g of carbon dioxide at 25°C and 100 kPa.

16 Determine the mass of oxygen present in a 10.0 L container of oxygen at a pressure of 105 kPa and at 20°C.

17 Calculate the temperature required for 0.2 g of helium to exert a pressure of 80 kPa in a container with a fixed volume of 4.0 L.

18 A sample of gas of mass 10.0 g occupies a volume of 5.4 L at 27°C and 100×10^3 Pa.

a Calculate the number of moles of gas in the sample.

b Determine the molar mass of the gas.

19 Determine which container holds more molecules of oxygen gas: container A of volume 40.0 L at 25°C and 770 mmHg or container B of volume 0.10 L at 45°C and 390 mmHg.

20 A room has a volume of 220 m^3.

a Calculate the number of moles of air particles in the room at 23°C and a pressure of 100 kPa.

b Assume that 20% of the molecules in the air are oxygen molecules and the remaining molecules are nitrogen. Calculate the mass of air in the room.

21 Octane is one of the main constituents of petrol. It burns according to the equation:

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

Calculate the mass of octane used if 50.0 L of carbon dioxide, measured at 120°C and 1.10 atm, was produced.

22 Calculate the volume of oxygen gas (in L) required for the complete combustion of 10.0 L of ethane gas at constant pressure and a temperature of 150°C.

23 Methane burns in excess oxygen according to the equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

This reaction produces 5 L of carbon dioxide at 200°C and 100 kPa. Assuming all volumes are measured at the same temperature and pressure, calculate the:

a volume of methane used

b volume of oxygen used

c mass of water vapour produced

Knowledge utilisation

24 Propane undergoes complete combustion as follows:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

All volumes are measured at 120°C and 100 kPa. An 80 mL volume of propane and 500 mL of oxygen are reacted.

a One of the gases does not react completely. Determine which gas is in excess and calculate the volume that is unreacted.

b Calculate the volumes of carbon dioxide and water that are produced in the reaction.

c Calculate the change in total volume of all the gases that has occurred as a result of the reaction.

25 An indoor gas heater burns propane (C_3H_8) at a rate of 12.7 g per minute. Determine the minimum mass of oxygen per minute, in g, that needs to be available for the complete combustion of propane.

26 The Gladstone Power Station, Queensland's largest, consumes about 4.0 million tonnes (1 tonne = 10^6 g) of coal in one year. The coal used in the power station is composed of approximately 25% carbon. Determine the volume of the greenhouse gas carbon dioxide released each year by the power station at STP.

CHAPTER 15

Water and aqueous solutions

Water has unique properties due to the intermolecular bonds and the intramolecular bonds water is able to form with itself and other compounds. In this chapter, you will investigate how the intermolecular forces increase its melting and boiling points. You will also understand the role water has in biological and environmental systems due to these physical properties.

Water is able to dissolve many different molecular solutes due to its polar nature. It is often referred to as the universal solvent due to its ability to form hydrogen bonds and dipole–dipole interactions when dissolving molecular solutes. Given that ionic solutes also contain charged particles, water is able to dissolve ionic solutes to varying degrees. The concentration of solute dissolved in water will also be discussed when looking at ways of calculating the amount of solute dissolved in water.

WS 2.2.1

Syllabus subject matter

Topic 2 • Aqueous solutions and acidity

AQUEOUS SOLUTIONS AND MOLARITY

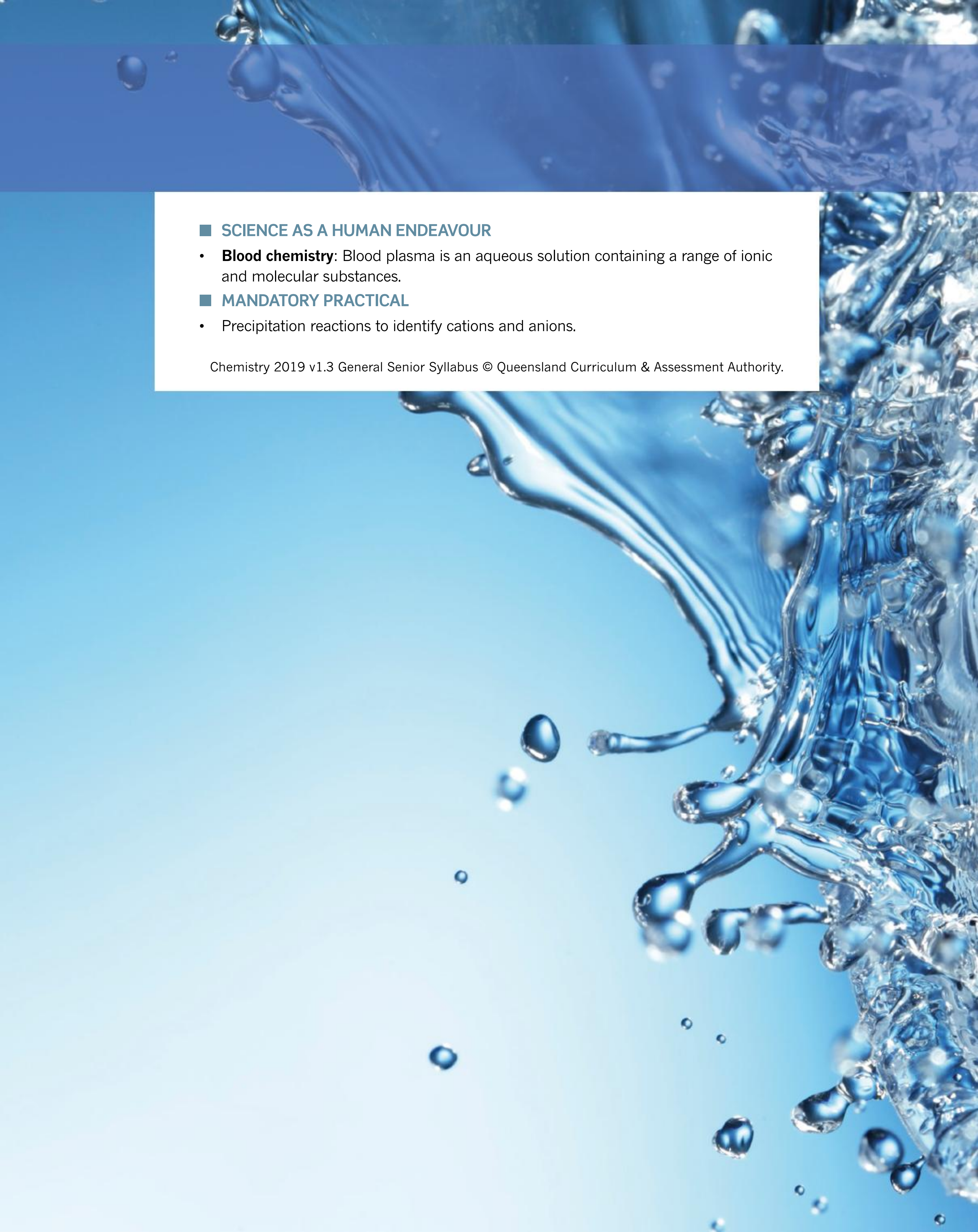
- understand that the unique properties of water, including boiling point, density in solid and liquid phases, surface tension and ability to act as a solvent can be explained by its molecular shape and hydrogen bonding between molecules
- distinguish between the terms ‘solute’, ‘solvent’, ‘solution’, and ‘concentration’
- recall that concentration can be represented in a variety of ways including by the number of moles of the solute per litre of solution ($mol\,L^{-1}$) and the mass of the solute per litre of solution ($g\,L^{-1}$) or parts per million (ppm)
- distinguish between unsaturated, saturated and supersaturated solutions
- use appropriate mathematical representations to solve and make predictions (including using the mole concept and the relationship between the number of moles of solute, concentration and volume of a solution) to calculate unknown values.

IDENTIFYING IONS IN SOLUTION

- apply solubility rules to determine the products of reactions and to predict if a precipitate will form
- determine the presence of specific ions in solutions based on evidence derived from chemical reactions, including precipitation and acid–carbonate reactions
- construct and use appropriate representations, including ionic formulas, chemical formulas, chemical equations and phase descriptions for chemical species to communicate conceptual understanding, solve problems and make predictions.

SOLUBILITY

- explain the relationship between the solubility of substances in water, including ionic and molecular substances, and the intermolecular forces between species in the substances and water molecules
- recognise that changes in temperature can affect solubility and recall that most gases become less soluble as solvent temperature increases while most solutes become more soluble as the solvent temperature increases
- interpret, analyse and evaluate data and solubility curves to communicate conceptual understanding, solve problems and make predictions.



SCIENCE AS A HUMAN ENDEAVOUR

- **Blood chemistry**: Blood plasma is an aqueous solution containing a range of ionic and molecular substances.

MANDATORY PRACTICAL

- Precipitation reactions to identify cations and anions.

Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

15.1 Structure and properties of water

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- discuss the unique properties of water, including its boiling point, surface tension and density
- relate the properties of water to the strength of intermolecular forces between molecules of water
- explain how the shape of water affects its ability to interact and dissolve other compounds.

Water is a special chemical. It moderates our weather, shapes our lands and is essential for the existence of life. Water is the most abundant liquid on Earth, covering more than 70% of our planet. The total water supply is estimated to be more than 1.3 billion cubic kilometres.

This water on Earth constantly cycles, as shown in Figure 15.1.1. It is in continuous movement between land, ocean, rivers and creeks and the atmosphere. Warmth from the Sun causes the water in oceans, lakes and rivers to evaporate and become vapour in the air. As water vapour rises, the air temperature drops and the water vapour condenses to form droplets, which merge to form clouds. When the water droplets are large enough, water falls back to Earth as rain, ice or snow, depending on the air temperature.

FIGURE 15.1.1 The water cycle. The arrows show some of the ways water on Earth is redistributed.

Other processes also add water to the atmosphere. When fossil fuels undergo combustion, water is a product and is released into the air. Water also evaporates from the leaves, stems and flowers of plants.

Water and life are so strongly linked that space scientists search the universe for water in their quest to discover possible life beyond our planet. It is the unique properties of water that allow it to support life.

These properties can be explained by looking at the structure of a water molecule and the **hydrogen bonding** that occurs between the water molecules.

WATER AND HYDROGEN BONDING

Structure of water

Water has the chemical formula H_2O, which means that each water molecule contains two hydrogen atoms and one oxygen atom. As you will recall from Chapter 8, hydrogen and oxygen are both non-metals so they bond with each other by sharing electrons. Each hydrogen atom shares one electron with the oxygen atom and the oxygen atom shares one electron with each hydrogen atom. The sharing of electrons means these **intramolecular bonds** are covalent.

Each oxygen atom also has two **lone pairs** (non-bonding pairs) of electrons that contribute to the molecule's shape. A water molecule is bent as shown in Figure 15.1.2. The lone pairs of electrons on the oxygen atom have greater repulsion than the electrons forming the single bonds between oxygen and hydrogen. This pushes the hydrogen atoms closer together, giving the overall bent shape (the shape of molecules was explained in detail in Chapter 12).

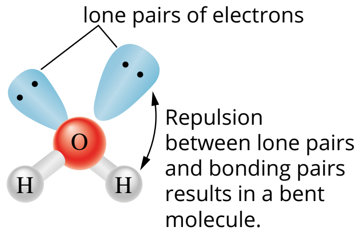


FIGURE 15.1.2 In this diagram of the valence structure of water, the two hydrogen atoms (white) form a bent shape with a central oxygen atom (red).

The covalent bonds in a water molecule are **polar**. The oxygen atom has a higher **electronegativity** than the hydrogen atoms. This means that the electrons in the O–H bonds are more strongly attracted to the oxygen atom. Uneven sharing of the electrons gives the oxygen atom a partial negative charge (δ–) and each hydrogen atom carries a partial positive charge (δ+).

Since the bonding electrons are distributed unevenly across the water molecule, the molecule can be described as polar. Figure 15.1.3 shows how the individual **dipoles** of the O–H bonds in water add together, resulting in an overall molecular dipole.

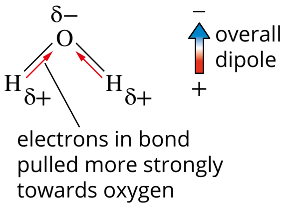


FIGURE 15.1.3 A water molecule has an overall dipole. The oxygen atom has a partial negative charge, while the hydrogen atoms both have partial positive charges.

Polar bonds and polar molecules were explained in detail in Module 12.2.

Water molecules are polar and have a permanent overall dipole.

Hydrogen bonding

The main type of **intermolecular force** between molecules in water is a hydrogen bond. The hydrogen bonds are formed by an **electrostatic attraction** between the partial positive charge on a hydrogen atom on one water molecule and a lone pair of electrons on the oxygen atom of a neighbouring water molecule.

Figure 15.1.4 shows the hydrogen bonding between molecules of water. Each water molecule is colour-coded to show its polarity. The two hydrogen atoms in each water molecule have a slight positive charge (white), while the central oxygen atom has a slight negative charge (red). The black dashed lines represent the hydrogen bonds that form between water molecules. (For more details on hydrogen bonding, see Module 12.3.)

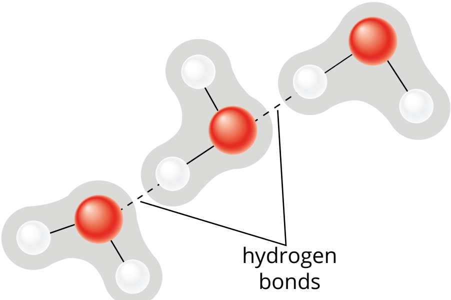


FIGURE 15.1.4 The hydrogen bonding between molecules of water

Because there is a large electronegativity difference between oxygen and hydrogen atoms, the partial charges on a water molecule are relatively large. The electrostatic attraction between these opposite partial charges on neighbouring molecules makes the hydrogen bonds between water molecules quite strong.

As Figure 15.1.5 shows, each water molecule can form hydrogen bonds with up to four other water molecules.

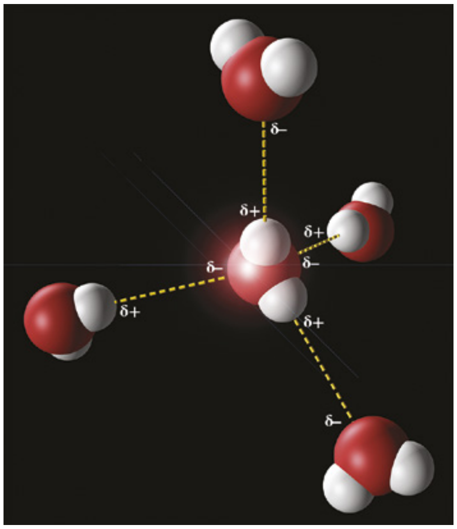


FIGURE 15.1.5 Each water molecule can form hydrogen bonds to four other water molecules.

MELTING AND BOILING POINTS OF WATER

Compared to other molecules of a similar size, water has high boiling and melting points. This can most easily be seen through the observed trends in these properties for the group 16 hydrides.

Trends in group 16 hydrides

The group 16 elements include oxygen (O), sulfur (S), selenium (Se), tellurium (Te) and polonium (Po). Each of these elements can bond with hydrogen to form compounds known as **hydrides**. Water is a group 16 hydride. The group 16 hydrides are listed in Table 15.1.1.

TABLE 15.1.1 Names and formulas of the group 16 hydrides

Element	Name and formula of hydride	Formula
O	water	H_2O
S	hydrogen sulfide	H_2S
Se	hydrogen selenide	H_2Se
Te	hydrogen telluride	H_2Te
Po	hydrogen polonide	H_2Po

The group 16 hydrides are all molecular compounds, so their melting and boiling points reflect the size of the forces between their molecules. The higher the melting and boiling points, the stronger the intermolecular forces must be because more energy is required to overcome the forces and allow the molecules to move apart from each other.

The melting and boiling points of the group 16 hydrides are shown in Table 15.1.2 and Figure 15.1.6.

TABLE 15.1.2 Melting and boiling points of the group 16 hydrides

Hydride	Melting point (°C)	Boiling point (°C)
H_2O	0	100
H_2S	−82	−60.7
H_2Se	−66	−41.5
H_2Te	−49	−2.2
H_2Po	−35	36.1

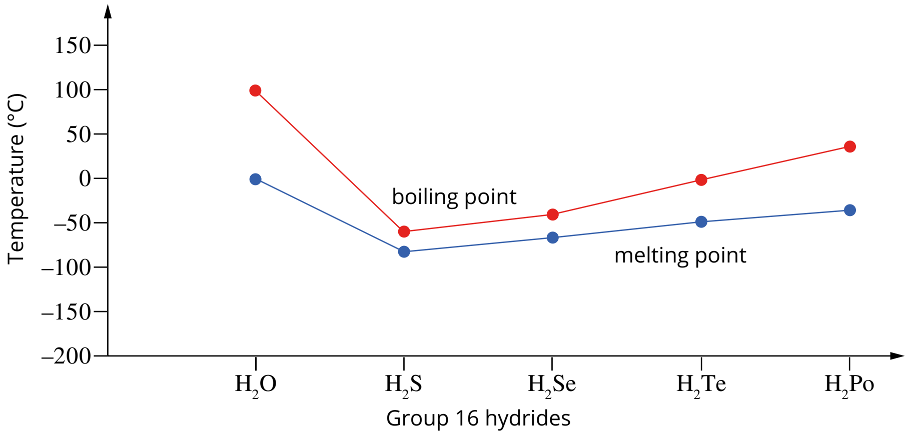


FIGURE 15.1.6 Graph showing the trend in melting and boiling points of water and other group 16 hydrides

WS 2.2.2

You can see from Figure 15.1.6 that, apart from water, there are clear trends in the melting and boiling points of the group 16 hydrides. With the exception of water, melting and boiling points increase going down the group. This increase indicates that the intermolecular forces are also getting stronger down the group. The intermolecular forces responsible for this trend are dispersion forces. Dispersion forces increase in strength with increasing mass. H_2Po has a larger mass than H_2S and H_2Se; hence it has stronger dispersion forces and higher melting and boiling points.

The melting and boiling points of water are exceptional

Water has a melting point of 0°C and a boiling point of 100°C. Both of these values are significantly higher than those of other group 16 hydrides. The values are also significantly higher than those of other molecular substances of a similar mass, as can be seen in Figure 15.1.7.

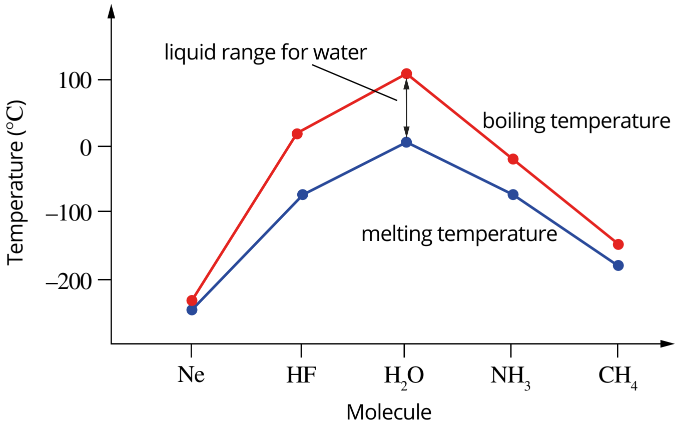


FIGURE 15.1.7 The melting and boiling points of water and other molecules of similar size

It is the relatively strong hydrogen bonds between its molecules that give water its exceptional properties.

A significant amount of energy is needed to disrupt all of the hydrogen bonds between water molecules, resulting in the higher melting points and boiling points observed for water.

ℹ Each water molecule has the potential to form four hydrogen bonds with surrounding water molecules. There are two partially charged hydrogen atoms and two lone pairs of electrons on the oxygen atom in each molecule, so all the hydrogen atoms and lone pairs in the molecule can be involved in hydrogen bonding.

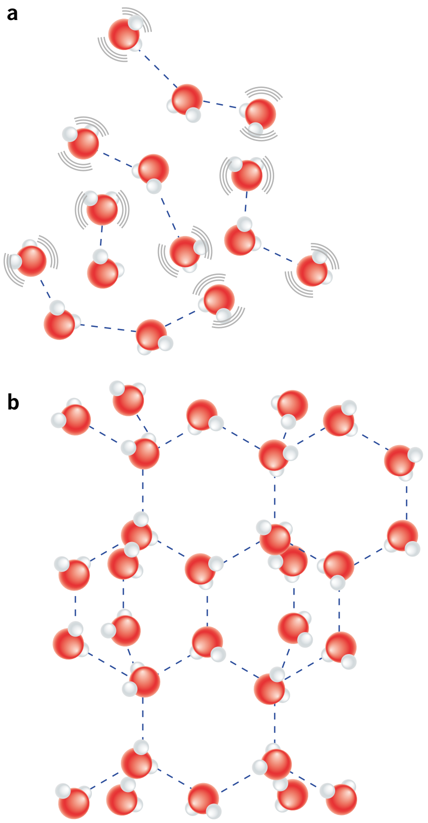


FIGURE 15.1.8 The arrangement of water molecules (a) in liquid water and (b) ice

DENSITY IN THE LIQUID AND SOLID STATES

Density is a characteristic property of a substance and is the measure of the mass of a substance within its volume. It is usually measured in grams per cubic centimetre. The nature of the bonding and intermolecular forces between molecules will affect the density of a substance, and water in the liquid and solid states is quite unique.

Water expands on freezing

When a sample of water freezes, the resulting ice crystal lattice has a greater volume than when the sample of water was a liquid. Water has the unusual property of having a lower density in the solid phase than in the liquid phase.

This property of water is important to life on Earth. When water at the surface of rivers or lakes freezes, the low density of the ice means it remains on the surface. This surface layer of ice insulates the water below from the cold air temperatures, reducing the possibility that the entire river or lake will freeze, and allowing aquatic life to survive.

As liquid water is cooled, the water molecules move more slowly. Upon approaching the freezing temperature of water, the molecules arrange in a way where each water molecule forms four hydrogen bonds to four neighbouring water molecules, as shown in Figure 15.1.8.

This creates a very open arrangement of molecules, meaning that the water molecules in ice are more widely spaced than in liquid water. Therefore, ice is less dense than liquid water. When ice melts, the water molecules move more freely and move closer together.

WATER HAS A HIGH SURFACE TENSION

The water molecules at the surface of a sample of water are not completely surrounded by other water molecules. As a result, they form hydrogen bonds with neighbouring molecules on the sides and below, but not above.

The sideways forces of attraction to neighbouring molecules are equal in all directions, so for molecules in the body of the liquid, there is no net force. However, the attraction of surface water molecules to the molecules below them has no opposite force, so the water molecules at the surface are pulled downwards.

In a beaker of water at room temperature, a water molecule will spend, on average, only 10^{-9} seconds or 1 nanosecond at the surface before being pulled back into the bulk of the liquid (Figure 15.1.9). The surface of the water is in a constant state of tension.

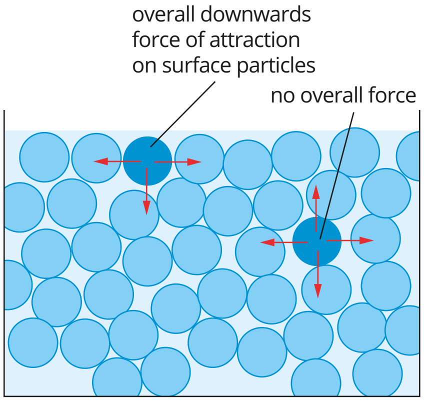


FIGURE 15.1.9 Forces of attraction on particles at the surface and in the body of a liquid. The particles in the body of the liquid experience no net force. The particles at the surface of the liquid have an overall downwards force of attraction.

Water, with relatively strong hydrogen bonds between molecules, has a relatively high **surface tension**. The surface tension of a substance is a measure of the resistance of a liquid to increasing its surface area. The surface tension is why falling water tends to form droplets. It also explains why insects are able to walk across water.

You can see in Figure 15.1.10 that the surface of the water behaves like a stretched skin. The water molecules at the surface are so attracted to each other by strong hydrogen bonds that the paperclip floats, even though the density of the paperclip is higher than the density of water.

FIGURE 15.1.10 The relatively high surface tension of water allows this paperclip to float.

15.1 Review

SUMMARY

- A water molecule is bent and contains polar covalent bonds.
- Water molecules are polar because an overall dipole exists as a result of the unsymmetrical nature of the polar O–H bonds.
- As a result of the O–H bonds in water molecules, the forces that attract one water molecule to another are relatively strong hydrogen bonds.
- Each water molecule can form up to four hydrogen bonds with other water molecules.
- The strength of water's hydrogen bonds means relatively large amounts of energy are required to disrupt the bonds and separate the molecules from each other. This gives water relatively high melting and boiling points.
- Water has significantly higher melting and boiling points than the other group 16 hydrides.
- With the exception of water, the melting and boiling points of the group 16 hydrides increase going down the group. This is due to the increasing strength of dispersion forces.
- Ice is less dense than liquid water because of its unique geometric arrangement of water molecules as a result of hydrogen bonding.
- The attraction of surface water molecules to the molecules below them has no opposite force, so the water molecules at the surface are pulled downwards creating a high surface tension.

KEY QUESTIONS

Retrieval

1 List the physical properties that make water unique.

2 Ice is less dense than water. Describe the bonding formed between water molecules in ice that means solid water is less dense.

3 Recall why the water molecules at the surface of water are pulled downwards generating surface tension, yet molecules in the body of the liquid do not show this.

4 Recall why water can form hydrogen bonds and dipole–dipole interactions with other molecules.

Comprehension

5 Explain why a water molecule can form up to four hydrogen bonds with other water molecules.

6 Describe and compare why water does not follow the trend of the other group 16 hydrides.

7 Explain the significance of polarity and hydrogen bonding in relation to the high melting and boiling points of water, and its density.

8 The overall structure of a water molecule is important in explaining its unique properties. Draw a diagram of the valence structure of water, indicating any partial charges and dipoles.

Analysis

9 Detergents are known to lower the surface tension of water. Determine the implications to aquatic organisms of detergents leeching into natural waterways.

15.2 Water as a solvent

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain how water's ability to act as a solvent depends largely on the shape of the molecule and the intermolecular bonds it can form with different solutes
- discuss the difference between the way ionic compounds, polar and non-polar molecules can interact with water molecules.

Water is an excellent solvent. This is one of its most important properties. Almost all biological processes and many industrial processes occur in water. These systems are known as **aqueous** environments.

When substances are dissolved in water, the particles are free to move throughout the **solution**. When two aqueous reactants are combined in a reaction vessel, the dissolved reactant particles mix freely. This increases the chances of the reactants coming into contact. Because of the increased movement of the reactant particles, interactions between reactants are generally much more effective than if the same reactants were mixed as solids.

In Figure 15.2.1, you can see how the deep purple colour of potassium permanganate spreads through the water as the solid dissolves. Eventually the liquid will appear completely purple as the particles continue to mix and move.

This module looks at the importance of the solvent properties of water in dissolving different materials in blood plasma. The process of dissolving will also be explored in more detail regarding the type of bonds that form when molecular and ionic substances dissolve in water.

FIGURE 15.2.1 When solid potassium permanganate is added to water, it dissolves. The particles disperse into the solution and move around freely.

FORCES BETWEEN SOLUTE AND SOLVENT PARTICLES

An aqueous solution is formed when a solid, liquid or gas is dissolved in water. The substance being dissolved is called the **solute** and the liquid in which the substance is dissolved is the **solvent**. When a solute dissolves, the intermolecular forces holding the molecules of the solute together and the **intermolecular bonds** holding the water molecules together must be broken. New bonds and forces of attraction between the solute and solvent can then form.

Aqueous solutions

The solute of an aqueous solution can change but the solvent will always be water. You can have an aqueous solution of salt (saline) or an aqueous solution of sugar. Table 15.2.1 lists some common aqueous solutions.

TABLE 15.2.1 Some everyday aqueous solutions

Solution	Solute/s	Solvent
saline solution (for use with contact lenses)	sodium chloride	water
soft drink	carbon dioxide, sugar, flavour, colour	water
coffee	coffee, sugar	water

All solutions have the following characteristics.

- The solute and the solvent cannot be distinguished from each other. This means the solution is **homogeneous**.
- The dissolved particles are too small to see.
- The amount of dissolved solute can vary from one solution to another.

ℹ In a homogeneous solution, the solute and solvent particles cannot be distinguished from each other. Every part of the solution is the same as any other part.

Process of dissolving

The process of a substance dissolving in another substance is called **dissolution**. When the two substances are liquids, you can say they are **miscible**.

During dissolution the following processes occur.

- The particles of the solute are separated from one another.
- The particles of the solvent are separated from one another.
- The solute and solvent particles are attracted to each other.

For a solution to form, the solute particles must interact with the solvent molecules. The solute particles are surrounded by solvent molecules and carried throughout the solution.

Forces involved in dissolving

For a substance to dissolve, there must be a change in the way particles in the solute and solvent interact. This means you need to look at the forces of attraction that occur between the particles.

It is useful to think of three different forces of attraction when considering if and how a substance will dissolve in water. These are the forces:

- that hold the particles of the substance (solute) together before it is added to the solvent
- that hold the solvent molecules together. In water, these forces are **hydrogen bonds**
- that can form between the solute particles and solvent molecules if the substance dissolves.

For a substance to dissolve, the attractive forces that form between the solute and solvent particles must be sufficient to overcome the forces between the particles in the solute and the forces between the solvent molecules.

Figure 15.2.2 shows that as a solute dissolves, the solute particles separate and become evenly distributed in the solvent. If the attraction between the solute and solvent particles is not strong enough, then the substance will not readily dissolve.

For a substance to be soluble, the solute–solvent interactions must be stronger than the solute–solute and solvent–solvent interactions.

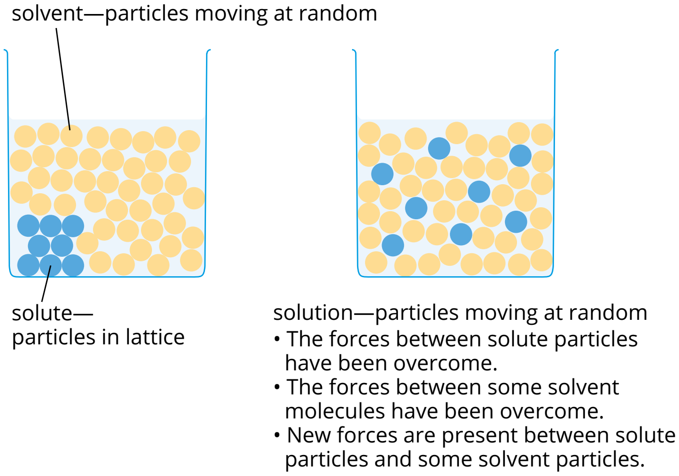


FIGURE 15.2.2 The rearrangement of particles when a solute dissolves in a solvent. The solute particles separate and become evenly distributed in the solvent.

You can use a 'like dissolves like' rule to predict whether a substance is likely to dissolve in another substance. The general statement that a solvent will only dissolve 'like' solutes tells you that:

- polar solvents will generally dissolve substances consisting of polar molecules or charged ions, but will not dissolve solutes made up of **non-polar** molecules
- non-polar solvents can dissolve substances consisting of non-polar molecules, but will not dissolve ones with polar molecules or ions.

Oil and other non-polar molecular substances do not dissolve well in water because the only intermolecular forces in these substances are dispersion forces, whereas the strong intermolecular forces of hydrogen bonding exist between the water molecules. These hydrogen bonds between water molecules are much stronger than the dispersion forces that could occur between molecules of oil and water. As a result, the attraction between the water molecules cannot be overcome and the water molecules do not separate to form a solution with the oil molecules.

You can observe the 'like dissolves like' rule with the miscibility of different liquids. Figure 15.2.3 shows a solution of ethanol and water. The polar nature of the ethanol molecule means it readily dissolves in water, which is also polar. When mixed, a homogeneous solution is formed with no separation between solute and solvent.

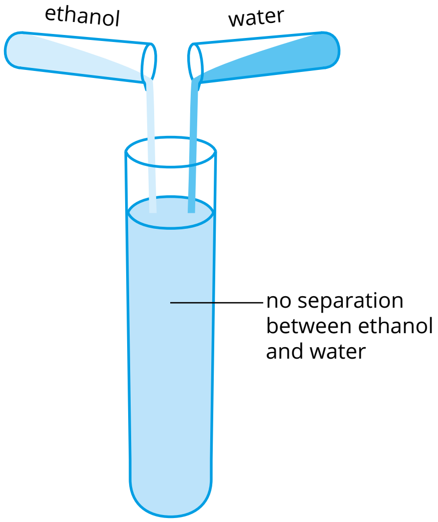


FIGURE 15.2.3 The polar compound ethanol is completely miscible in water. A homogeneous solution is formed on mixing.

In contrast, hexane is immiscible in water. Hexane is composed of non-polar molecules that will not interact with the polar water molecules. Figure 15.2.4 shows that when hexane and water are mixed, two layers form and the less dense hexane sits on top of the water.

Like hexane, olive oil is also composed of non-polar molecules. When olive oil and hexane are mixed with each other, the 'like dissolves like' rule applies and the two non-polar liquids readily mix with each other to form a homogeneous solution (Figure 15.2.5).

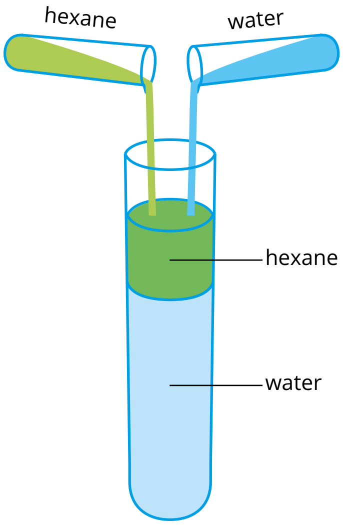


FIGURE 15.2.4 Hexane is a non-polar molecular compound that is not miscible in water. Two layers are formed with the non-polar liquid sitting on top of the water layer.

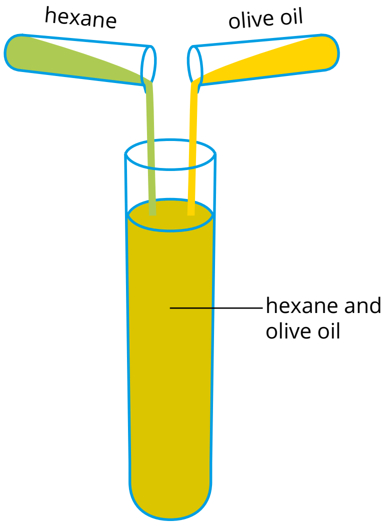


FIGURE 15.2.5 Hexane and olive oil are both non-polar molecular compounds and so are completely miscible.

Use the 'like dissolves like' principle to predict if something will dissolve in a solvent.

MOLECULAR COMPOUNDS THAT FORM HYDROGEN BONDS WITH WATER

One way a molecular compound might dissolve in water is if its molecules form hydrogen bonds with water molecules. An example of such a molecule is ethanol.

Ethanol (C_2H_5OH) is a liquid at room temperature. Its molecules contain the polar –OH group, with lone pairs of electrons on the oxygen atom. The presence of the hydrogen atoms bonded to the more electronegative oxygen atom allows an ethanol molecule to form hydrogen bonds.

Figure 15.2.6 shows how hydrogen bonds form between molecules in pure ethanol.

When ethanol is added to water, it dissolves. The two solutions are miscible. Figure 15.2.7 shows how hydrogen bonds form between the ethanol and surrounding water molecules.

Because the strength of the intermolecular forces between the solute and the solvent are similar, the two substances can readily interact with each other. Therefore, water and ethanol molecules mix freely with each other, held together in solution by hydrogen bonds.

Molecular compounds that can form hydrogen bonds are often soluble in water.

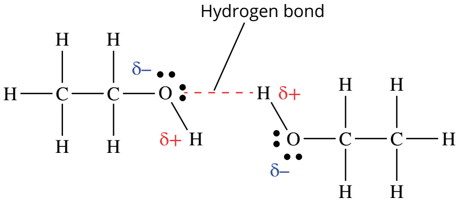


FIGURE 15.2.6 Hydrogen bonding in pure ethanol. The intermolecular hydrogen bond is formed between the lone pair electrons on the oxygen atom of one ethanol molecule and the electron-deficient hydrogen atom of an adjacent ethanol molecule.

In summary, when ethanol dissolves in water, hydrogen bonds between:

- water molecules break
- ethanol molecules break
- ethanol molecules and water molecules form.

An equation for the dissolution of ethanol can be written to represent this process:

$$C_2H_5OH(l) \xrightarrow{H_2O(l)} C_2H_5OH(aq)$$

Note that the formula of water sits above the arrow. This is because there is no direct reaction between the water and the ethanol. The two substances simply mix together. No chemical change occurs; only the state symbol for ethanol is altered from (l) to (aq), indicating it is now dissolved in water.

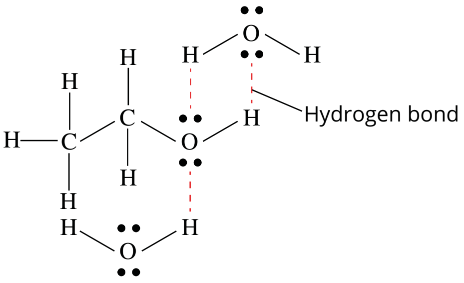


FIGURE 15.2.7 Hydrogen bonding between ethanol and water

When a polar molecular substance dissolves in water by forming hydrogen bonds, the equation for dissolution places water above the arrow.

Other polar compounds that form hydrogen bonds with water

Sugars including glucose, fructose and sucrose can all dissolve in water by forming hydrogen bonds. The structure of each of these sugars is shown in Figure 15.2.8. As with ethanol, their molecules also contain the polar –OH group.

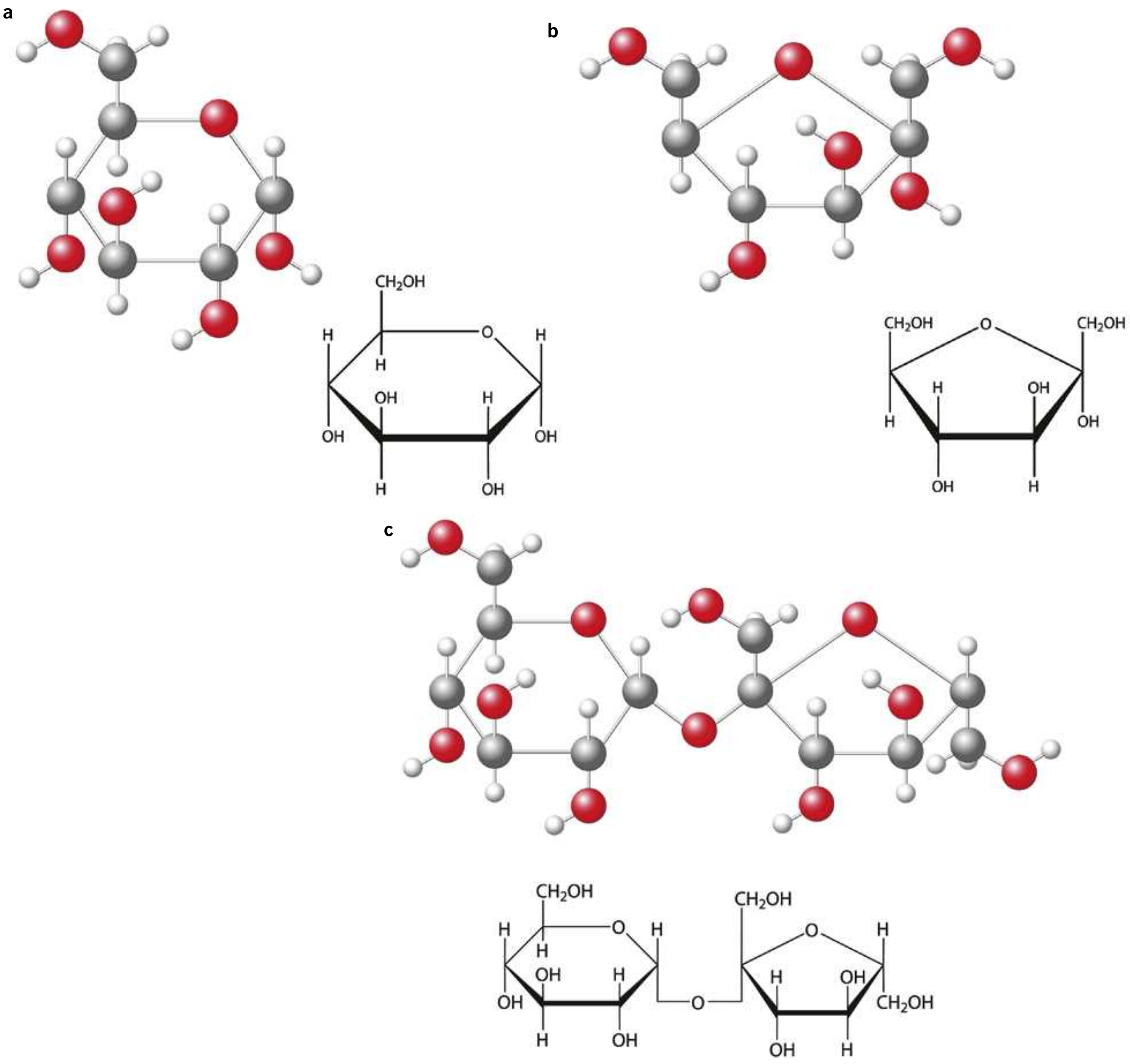


FIGURE 15.2.8 These are the molecular structures of common sugars: (a) glucose (the sugar found in our blood), (b) fructose (fruit sugar) and (c) sucrose (table sugar—the sugar used to sweeten tea and coffee). Each of these sugars can dissolve in water because of the presence of polar –OH groups.

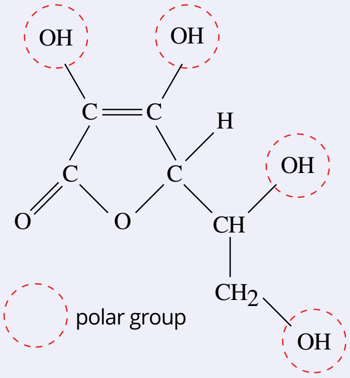
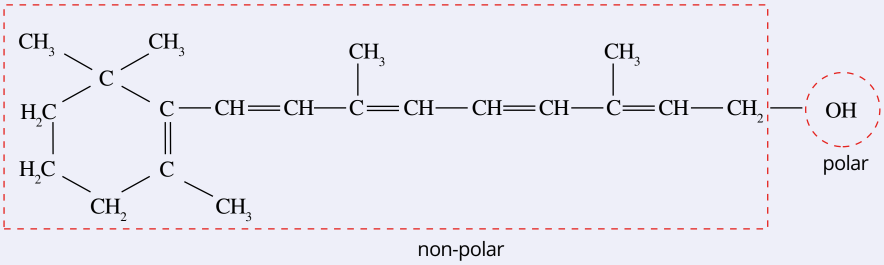
The more polar the molecules of a molecular compound are, the more likely the compound is to dissolve in water. Some molecules have a polar section and a non-polar section. In general, the larger the non-polar section of the molecule, the less soluble it is in water.

Carboxylic acids are a type of organic molecule that have a hydrocarbon chain of varying length and a carboxyl functional group. The carboxyl functional group contains a carbon double-bonded to an oxygen and a bond to a hydroxyl functional group. It is often written as –COOH. The presence of the hydroxyl (–OH) functional group makes this end of the molecule polar. However, the hydrocarbon chain is non-polar due to the symmetrical nature of the hydrocarbon chain. Therefore, the longer the chain, the more of the molecule is non-polar. Therefore, the **solubility** of carboxylic acids and alcohols is largely dependent on the size of the hydrocarbon chain to which it is attached.

Non-polar molecular substances do not have charged ends so there is no significant attraction to water molecules. The only forces of attraction that exist between non-polar substances and water are weak dispersion forces, which are not strong enough to overcome the relatively strong hydrogen bonding between water molecules.

Table 15.2.2 compares the structures and properties of vitamin C and vitamin A. Both molecules contain the polar –OH group, but only vitamin C is soluble in water. The higher proportion of polar –OH groups on the vitamin C molecule allows it to form sufficient hydrogen bonds with water to overcome the strong attraction between water molecules, therefore dissolving to form a solution.

TABLE 15.2.2 Comparison of the solubility and structures of vitamin C and vitamin A

	Vitamin C	Vitamin A
structure	 Vitamin C: • contains four polar O–H groups • molecule quite polar.	 Vitamin A: • contains one polar O–H group • molecule largely non-polar.
solubility	highly soluble in water, insoluble in fats	soluble in fats, insoluble in water
biological significance	The high solubility of vitamin C means it is excreted in urine so it must form a regular part of the diet.	Vitamin A is not excreted in urine but is stored in body fat. The body can tolerate a diet low in vitamin A for a limited period.

MOLECULAR COMPOUNDS THAT IONISE

Some compounds have molecules with one or more covalent bonds that are so polar they break when the compound is placed in water. Hydrogen chloride is such a compound. Hydrogen chloride (HCl) is a gas at room temperature. Chlorine is much more electronegative than hydrogen so the H—Cl covalent bond is highly polar; the molecule forms a dipole.

When hydrogen chloride is added to water, the hydrogen atom in HCl forms such a strong attraction to the oxygen atom in a water molecule that the H—Cl covalent bond breaks. The two electrons that made up the H—Cl covalent bond remain with the more electronegative Cl atom and the newly formed hydrogen ion (H^+) joins the water molecule.

Highly polar molecular compounds will form ions when dissolved in water.

Figure 15.2.9 shows that when the hydrogen ion (H^+) bonds to the water molecule, it forms a new ion known as the **hydronium ion** (H_3O^+).

Since the Cl atom has gained an electron, it has a negative charge, forming a chloride ion (Cl^-).

The HCl molecule is split into smaller particles, so this is a dissociation reaction. The HCl is said to have become **ionised**; that is, it has undergone a process that has produced ions (Figure 15.2.9).

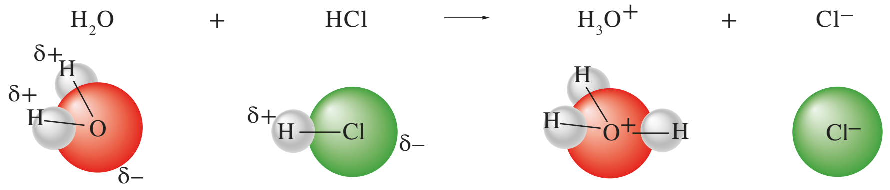


FIGURE 15.2.9 The dipole–dipole attraction between the molecules of water and hydrogen chloride leads to the breaking of the polar covalent bond between the hydrogen and chlorine atoms. New ions, hydronium and chloride, are formed in a process called ionisation.

Figure 15.2.10 shows how the two ions produced in the reaction of HCl with water (Cl^- and H_3O^+) become **hydrated**. The charged ions are surrounded by other polar water molecules. They are held in solution by **ion–dipole attractions**. (Ion–dipole attractions will be discussed later in this module).

In summary, when hydrogen chloride dissolves in water:

- polar covalent bonds within hydrogen chloride molecules break, producing hydrogen ions (H^+) and chloride ions (Cl^-)
- a covalent bond forms between each H^+ and an H_2O molecule, forming H_3O^+ ions
- ion–dipole attractions form between the H_3O^+ and Cl^- ions and polar water molecules.

An equation can be written to represent this process:

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

You should note two important points about this equation.

- Water is a reactant because there has been a rearrangement of atoms to form new substances.
- The aqueous state of the H_3O^+ and Cl^- ions tells you that they are hydrated in solution.

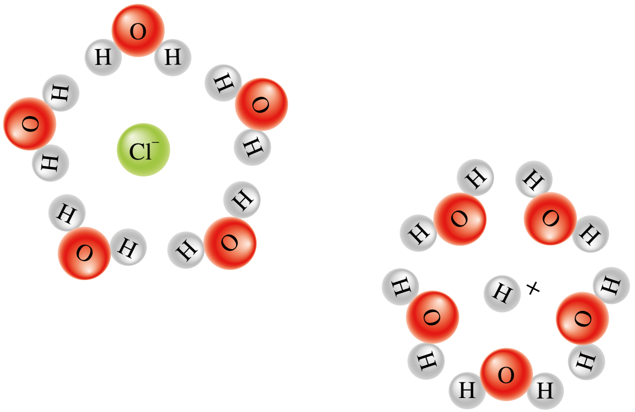


FIGURE 15.2.10 When HCl molecules ionise in water, the newly formed ions are surrounded by adjacent water molecules. The charges of the ions interact strongly with the polar water molecules. For simplicity, the hydronium ion (H_3O^+) is represented by the hydrogen ion (H^+).

Other polar molecular compounds that ionise

Other compounds that dissolve in water by ionising include the common acids nitric acid (HNO_3), sulfuric acid (H_2SO_4) and ethanoic acid (CH_3COOH).

Ammonia (NH_3) is another polar molecule that can dissolve in water. Ammonia can dissolve by either forming hydrogen bonds with water or by reacting to form the ammonium ion (NH_4^+).

WATER AS A SOLVENT OF IONIC COMPOUNDS

Many ionic compounds dissolve readily in water. Sodium chloride is a typical ionic compound that exists as a solid at room temperature. In Figure 15.2.11, you can see the arrangement of sodium cations (Na^+) and chloride anions (Cl^-) in a three-dimensional ionic lattice. The ions are held together by strong electrostatic forces between the positive and negative charges of the ions.

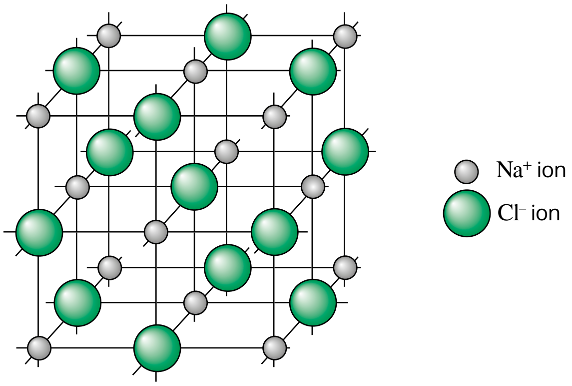


FIGURE 15.2.11 A representation of the crystal lattice of the ionic compound sodium chloride.

When an ionic compound such as sodium chloride is added to water, the positive ends of the water molecules are attracted to the negatively charged chloride ions and the negative ends of the water molecules are attracted to the positively charged sodium ions (Figure 15.2.12). The attraction between an ion and a polar molecule such as water is described as an ion–dipole attraction.

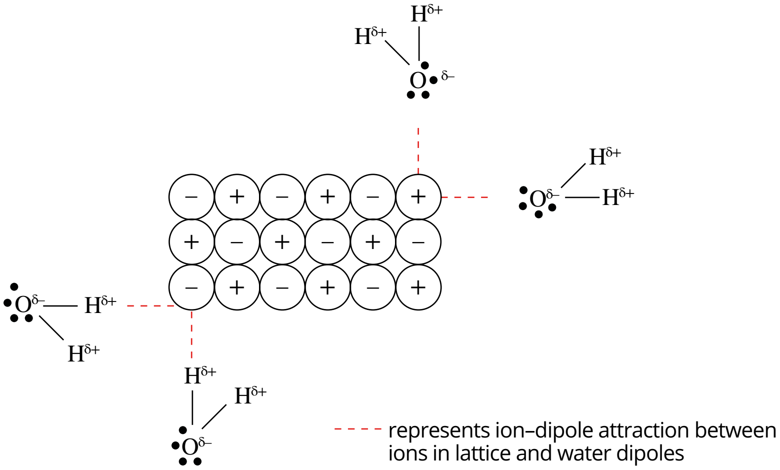


FIGURE 15.2.12 Electrostatic attraction occurs between the negative ions in a NaCl lattice and the hydrogen atoms in polar water molecules. Electrostatic attraction also occurs between the positive ions in the lattice and the oxygen atoms in water molecules.

Water molecules are in a continuous state of random motion. If the ion–dipole attractions between the ions and the water molecules are strong enough, the water molecules can pull the sodium and chloride ions on the outer part of the crystal out of the lattice and into the surrounding solution.

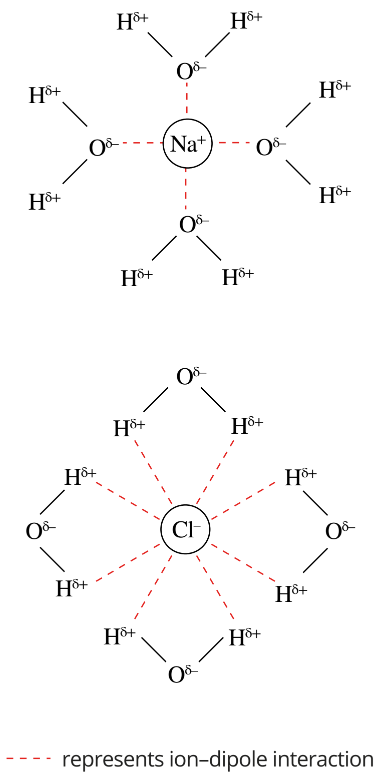


FIGURE 15.2.13 Ion–dipole attraction between the ions and adjacent water molecules form hydrated sodium and chloride ions.

Ionic substances dissolve by dissociation. Ion–dipole bonds are formed between the ions and water molecules.

FIGURE 15.2.14 Limestone caves are formed by limestone (calcium carbonate) dissolving over very long periods of time. This photograph is of a limestone cave at Loch Ard Gorge, Victoria.

Substances are rarely ever completely 'soluble' or 'insoluble'. The solubility of substances can be considered a scale.

Sodium ions and chloride ions pulled out of the lattice become surrounded by water molecules. These ions are said to be hydrated. Water molecules are arranged around the ions as shown in Figure 15.2.13. Note the different arrangements of the water molecules around the positive and negative ions. The hydrogen atoms in the water molecule are more positive so they are orientated towards the negative chloride ion. The positive sodium ion is surrounded by the more negative oxygen atom of the water molecules.

The process of separating positive and negative ions from a solid ionic compound to form hydrated ions when an ionic compound dissolves in water is called **dissociation**.

Although the ionic bonds within the lattice are strong, the ions can be pulled away from the lattice by the interactions of many water molecules.

In summary, when sodium chloride dissolves in water:

- ionic bonds within the sodium chloride lattice are broken
- hydrogen bonds between water molecules are broken
- ion–dipole attractions form between ions and polar water molecules.

An equation can be written to represent the dissociation process:

$$NaCl(s) \xrightarrow{H_2O(l)} Na^+(aq) + Cl^-(aq)$$

Note that the formula of water sits above the arrow. This is because there is no direct reaction between the water and the sodium chloride. No chemical change occurs; only the state symbol for sodium chloride is altered from (s) to (aq), indicating it is now dissolved in water. (You may omit the H_2O from this equation if you wish.)

It is important to note that dissociation of ionic compounds is simply freeing ions from the lattice so that they can move freely throughout the solution. This is different from ionisation of molecular compounds where new ions are formed by the reaction of the molecule with water.

INSOLUBLE IONIC COMPOUNDS

Not all ionic compounds are soluble in water. For example, **limestone** ($CaCO_3$) is almost completely insoluble in water. Limestone caves such as those seen in Figure 15.2.14 are formed over a long period of time as $CaCO_3$ is dissolved and redeposited. The ionic compound $Ca_3(PO_4)_2.Ca(OH)_2$, which gives strength to bones and teeth, is also (fortunately) insoluble in water.

Insoluble ionic compounds do not dissolve in water because the energy required to separate the ions from the lattice is greater than the energy released when the ions are hydrated.

Although substances are usually described as 'soluble' or 'insoluble', this is a generalisation. Substances that are described as 'insoluble' tend to dissolve very slightly. Those that are described as 'soluble' dissolve to varying extents.

The SNAPE rule

A handy way to remember many of the soluble salts is with the initials SNAPE. Salts that contain one or more of the following ions are soluble:

- **S**odium (Na^+)
- **N**itrate (NO_3^-)
- **A**mmonium (NH_4^+)
- **P**otassium (K^+)
- **E**thanoate (CH_3COO^-).

Solubility tables

A **solubility table** can be used to determine whether common ionic compounds are soluble in water. Anions and cations are listed with indications of their solubility in different compounds. Table 15.2.3 on page 463 is a sample solubility table listing ionic compounds that are soluble in water. Note that for some ions there are exceptions. Insoluble compounds are given in Table 15.2.4 on page 463.

TABLE 15.2.3 Relative solubilities of soluble ionic compounds

Soluble in water (>0.1 mol dissolves per L at 25°C)	Exceptions: insoluble (<0.1 mol dissolves per L at 25°C)	Exceptions: slightly soluble (0.01–0.1 mol dissolves per L at 25°C)
most chlorides (Cl^-), bromides (Br^-) and iodides (I^-)	AgCl, AgBr, AgI, PbI_2	$PbCl_2$, $PbBr_2$
all nitrates (NO_3^-)	no exceptions	no exceptions
all ammonium (NH_4^+) salts	no exceptions	no exceptions
all sodium (Na^+) and potassium (K^+) salts	no exceptions	no exceptions
all ethanoates (CH_3COO^-)	no exceptions	no exceptions
most sulfates (SO_4^{2-})	$SrSO_4$, $BaSO_4$, $PbSO_4$	$CaSO_4$, Ag_2SO_4

TABLE 15.2.4 Relative solubilities of insoluble ionic compounds

Insoluble in water	Exceptions: soluble	Exceptions: slightly soluble
most hydroxides (OH^-)	NaOH, KOH, $Ba(OH)_2$, NH_4OH*, AgOH†, $Ca(OH)_2$	$Sr(OH)_2$
most carbonates (CO_3^{2-})	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$	no exceptions
most phosphates (PO_4^{3-})	Na_3PO_4, K_3PO_4, $(NH_4)_3PO_4$	no exceptions
most sulfides (S^{2-})	Na_2S, K_2S, $(NH_4)_2S$	no exceptions

* NH_4OH does not exist in significant amounts in an ammonia solution. Ammonium and hydroxide ions readily combine to form ammonia and water.

† AgOH readily decomposes to form a precipitate of silver oxide and water.

Using a solubility table to predict solubility

The solubility tables provided as Tables 15.2.3 and 15.2.4 can be used to predict whether a particular ionic substance will be soluble, slightly soluble or insoluble in water.

Worked example 15.2.1 shows you a systematic way of using a solubility table to determine whether an ionic compound will be soluble or insoluble in water.

Worked example 15.2.1

DETERMINING IF IONIC COMPOUNDS ARE SOLUBLE OR INSOLUBLE IN WATER

Determine if barium sulfide (BaS) is soluble or insoluble in water. Refer to the solubility data in Tables 15.2.3 and 15.2.4 to complete this question.

Thinking	Working
Identify the ions that are present in the ionic compound.	Barium (Ba^{2+}) and sulfide (S^{2-})
Check the solubility tables to see if compounds containing the cation are usually soluble or insoluble in water.	Barium ions do not appear in Tables 15.2.3 or 15.2.4. This is no help, so look for the other ion.
Check the solubility tables to see if compounds containing the anion are usually soluble or insoluble in water.	Compounds containing sulfide ions are usually insoluble in water. So, barium sulfide will be insoluble in water.

➤ Try yourself 15.2.1

DETERMINING IF IONIC COMPOUNDS ARE SOLUBLE OR INSOLUBLE IN WATER

Determine if ammonium phosphate ($(NH_4)_3PO_4$) is soluble or insoluble in water. Refer to the solubility data in Tables 15.2.3 and 15.2.4 to complete this question.

SCIENCE AS A HUMAN ENDEAVOUR

Blood chemistry

Water's unique properties make it essential to life, particularly its ability to dissolve an enormous range of different solutes. The human body is 66% water by weight. For biological organisms, water provides a system to transport nutrients and soluble wastes. It also transports essential reagents for chemical reactions to cells and tissues.

The circulatory system of the body forms a transportation network (Figure 15.2.15). Blood plasma is an aqueous solution containing a range of ionic and molecular substances, which enables essential nutrients, ions and molecules dissolved in water to be carried to your organs, brain and tissues. The flow of blood through the circulatory system also carries away waste products.

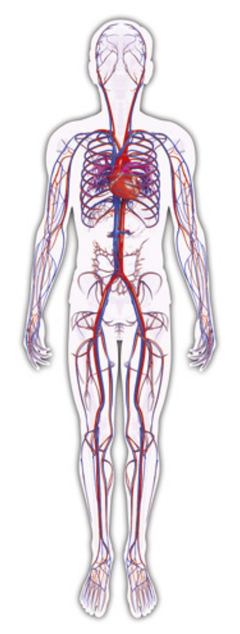


FIGURE 15.2.15 The circulatory system of the body transports many ions and molecules dissolved in water. The solute particles include nutrients and waste products from the body's metabolic processes.

Blood is a suspension of red blood cells, white blood cells and plasma (Figure 15.2.16). Plasma makes up over 90% of the blood and contains mostly water. This water acts as a solvent for carrying other substances throughout the circulatory system.

Glucose is a monosaccharide used to produce energy in all cells via cellular respiration. It is therefore essential that glucose can be transported around the body in the bloodstream. Glucose has many hydroxyl functional groups that form hydrogen bonds with water (Figure 15.2.8 on page 458).

FIGURE 15.2.16 Centrifuging blood separates it into clear layers, demonstrating that blood is actually a heterogeneous suspension. Here you can see the plasma on top and the red blood cells on the bottom.

Proteins are polar molecules made up of amino acids. Amino acids have polar regions, therefore allowing proteins to dissolve readily in the water in plasma. These proteins may be hormones like insulin that is produced in the pancreas. Insulin triggers cells to uptake more glucose from the bloodstream to maintain blood sugar levels.

Sodium and potassium are important salts necessary for the transmission of nerve impulses and also need to be transported around the body. The ions dissolve readily by forming **ion–dipole interactions** with the water molecules in plasma.

Review

1 Antibodies are important biological proteins that fight infections against invading pathogens such as varicella (which causes chicken pox) in the human body. Calcium is an important salt needed for muscle contraction. Distinguish between the ways in which these substances dissolve in blood plasma.

2 Some hormones, such as oestrogen and testosterone, are composed of lipids which are non-polar biomacromolecules. Discuss how these hormones interact with the watery component of blood plasma and how they travel around the body.

15.2 Review

SUMMARY

- A solution is a homogeneous mixture in which molecules or ions are evenly dispersed throughout a solvent.
- Solutions in which water is the solvent are called aqueous solutions.
- A solution forms when the bonds between the solute and solvent particles are sufficiently strong to compete with attractive forces between the solute particles and between the solvent particles.
- 'Like dissolves like' is a rule that predicts that polar solutes will dissolve in polar solvents and non-polar solutes will dissolve in non-polar solvents.
- Water's solvent properties make it essential to life. It provides transport of nutrients in living organisms and provides an aqueous environment for chemical reactions.
- Water is a good solvent for some polar molecular compounds.
- Some polar molecular compounds dissolve by forming hydrogen bonds with water.
- Some covalent molecular compounds are so polar that they dissolve in water by ionising to form hydrated ions.
- Table 15.2.5 below summarises the two ways that molecular compounds dissolve in water.

TABLE 15.2.5 The way molecular compounds dissolve in water

Type of solute	Example of solute	Bond broken in the solute	Bonds formed with water	Equation
polar covalent molecule that can hydrogen bond	ethanol, C_2H_5OH	hydrogen bonds between ethanol molecules	hydrogen bonds between ethanol and water molecules	$C_2H_2OH(l) \xrightarrow{H_2O(l)} C_2H_5OH(aq)$
polar molecules that ionise	hydrogen chloride, HCl	covalent bond between hydrogen and chlorine atoms in the HCl molecule	covalent bond between H^+ from HCl and oxygen atom in the water molecule, forming H_3O^+ ions. H_3O^+ and Cl^- ions form ion–dipole bonds with water molecules	$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

- Soluble ionic compounds dissociate in water to form hydrated ions.
- In a hydrated ion, water molecules are attracted to the central ion by ion–dipole attractions.
- Table 15.2.6 summarises the way soluble ionic compounds dissolve in water.
- The SNAPE rule can be used to predict the solubility of an ionic compound.
- A solubility table can also be used to determine whether a particular ionic compound is likely to be soluble or insoluble in water (Table 15.2.6).

TABLE 15.2.6 Example of a solubility table

Type of solute	Example	Bond broken in the solute	Bonds formed with water	Equation
soluble ionic compounds	sodium chloride (NaCl)	ionic bonds between Na^+ and Cl^- ions	ion–dipole attractions between dissociated ions and polar water molecules	$NaCl(s) \xrightarrow{H_2O(l)} Na^+(aq) + Cl^-(aq)$

continued over page

15.2 Review *continued*

KEY QUESTIONS

Retrieval

1 Identify each of the following components of a glass of wine as solute, solvent or solution.
- **a** water
- **b** wine
- **c** ethanol
- **d** sugar

2 Describe how molecular substances with hydroxyl functional groups are able to dissolve in water.

3 State, in terms of its structure, why vitamin C must be consumed regularly in the diet.

Comprehension

4 Determine if the following substances are likely to be soluble or insoluble in water.
- **a** ammonia (NH_3)
- **b** oxygen gas (O_2)
- **c** hydrogen chloride (HCl)
- **d** methanol (CH_3OH)
- **e** methane (CH_4)
- **f** hydrogen fluoride (HF)
- **g** carbon dioxide (CO_2)

5 Methanol (CH_3OH) and glucose ($C_6H_{12}O_6$) form hydrogen bonds with water. They will dissolve in water without ionising. Write chemical equations to represent the dissolving process for each of these compounds.

6 With reference to the bonds that are broken and new bonds that are formed, describe how ethanol dissolves in water.

7 Hydrogen iodide (HI) will ionise when it dissolves in water. The ionisation reaction is similar to that of HCl. Show a chemical equation to demonstrate the dissolving process for this compound.

8 Sodium nitrate ($NaNO_3$) and calcium hydroxide ($Ca(OH)_2$) will both dissociate when they dissolve in water. Write chemical equations to depict the dissolving process for each of these compounds.

9 Redraw the representations of Na^+ and Cl^- ions below, then show the correct orientation of water molecules around hydrated sodium and chloride ions. Use a bent shape to represent the shape of water molecules.

Analysis

10 Identify reasons for each of the following.
- **a** Concentrated deposits of nitrate compounds are only found in desert regions.
- **b** The sea is a rich source of sodium, chloride and sulfate ions.

11 Predict which of the following substances you would expect to be soluble in water. Refer to Tables 15.2.3 and 15.2.4 on page 463 to predict your answer.
- **A** sodium carbonate
- **B** lead(II) nitrate
- **C** magnesium carbonate
- **D** ammonium sulfate
- **E** iron(II) sulfate
- **F** magnesium phosphate
- **G** zinc carbonate
- **H** sodium sulfide
- **I** silver chloride
- **J** barium sulfate

12 Sodium chloride is very soluble in water. Identify what can be concluded about the forces present between the solute and solvent particles.

13 Using Tables 15.2.3 and 15.2.4 on page 463, identify:
- **a** the negative ion that dissolves most readily in water
- **b** the positive ion that is least soluble in water

14 Fatty acids are long hydrocarbon chains with a double-bonded oxygen and hydroxyl group attached to the last carbon atom. Predict what would happen to the solubility of a fatty acid as the hydrocarbon chain length increases.

15 Compare the arrangement of water molecules around dissolved magnesium ions in water with the arrangement of chloride ions in water.

15.3 Solubility

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recognise that temperature affects the solubility of ionic salts and gases in water
- interpret and analyse solubility curves to determine whether a solution is unsaturated, saturated or supersaturated
- make predictions of a compound's ability to dissolve in water at different temperatures.

The solubility of a substance is a measure of how much of the substance will dissolve in a given amount of a solvent. For example, glucose is very soluble in water and is used by your body as a readily available energy source. Most rocks are made of minerals that are insoluble in water, but, as you saw in Module 15.2, limestone (which contains calcium carbonate) is slightly soluble and so caves form in areas of limestone rock over long periods.

In this module, you will learn how to measure the solubility of some compounds in water and how to predict whether a compound will be soluble or not.

DEFINITION OF SOLUBILITY

In chemistry, the term solubility has a specific meaning. It refers to the maximum amount of a solute (substance that is dissolved) that can be dissolved in a given quantity of a solvent (the common quantity often used is 100 g) at a certain temperature. Table 15.3.1 gives the solubility of some common substances in 100 g of water at 18°C.

TABLE 15.3.1 Solubility of solutes at 18°C

Solute	Solubility (g per 100 g of water)
sugar (sucrose)	200
salt (sodium chloride)	35
limestone (calcium carbonate)	0.0013

Different kinds of solutions

There are three terms that describe the different solutions that result from dissolving solutes in solvents.

- A **saturated solution** is one in which no more solute can be dissolved at a particular temperature.
- An **unsaturated solution** contains less solute than is needed to make the solution saturated. Unsaturated solutions can dissolve more solute.
- A **supersaturated solution** is an unstable solution that contains more dissolved solute than a saturated solution. If this type of solution is disturbed, some of the solute will separate from the solvent as a solid.

Figure 15.3.1 shows a supersaturated solution of sodium ethanoate. Supersaturated sodium ethanoate is prepared by cooling a saturated solution very carefully so that solid crystals do not form. Adding a small **seed crystal** to the supersaturated solution causes the solute to **crystallise** (form solid crystals) so that a saturated solution remains.

FIGURE 15.3.1 Crystals of sodium ethanoate form after a seed crystal is added to a supersaturated solution of the compound.

SOLUBILITY CURVES

The solubility of many substances changes significantly as the temperature changes. For example, you can dissolve more chocolate powder in hot milk than in cold milk. For most solids, increasing the temperature increases the solubility in a liquid. This is because at higher temperatures, both the solute and solvent have more energy to overcome the forces of attraction that hold the particles together in the solid.

The relationship between solubility and temperature can be represented by a **solubility curve** as shown in Figure 15.3.2.

Solubility curves show the solubility of a substance as a function of temperature. For the solubility curves featured in Figure 15.3.2, note the following points.

- Each point on a curve represents the maximum amount of the solute that can be dissolved in 100 g of water at a particular temperature. Therefore, each point on the curve represents a saturated solution.
- Any point below a curve represents an unsaturated solution.
- Any point above a curve represents a supersaturated solution.
- For most solids, as temperature increases, the solubility increases.

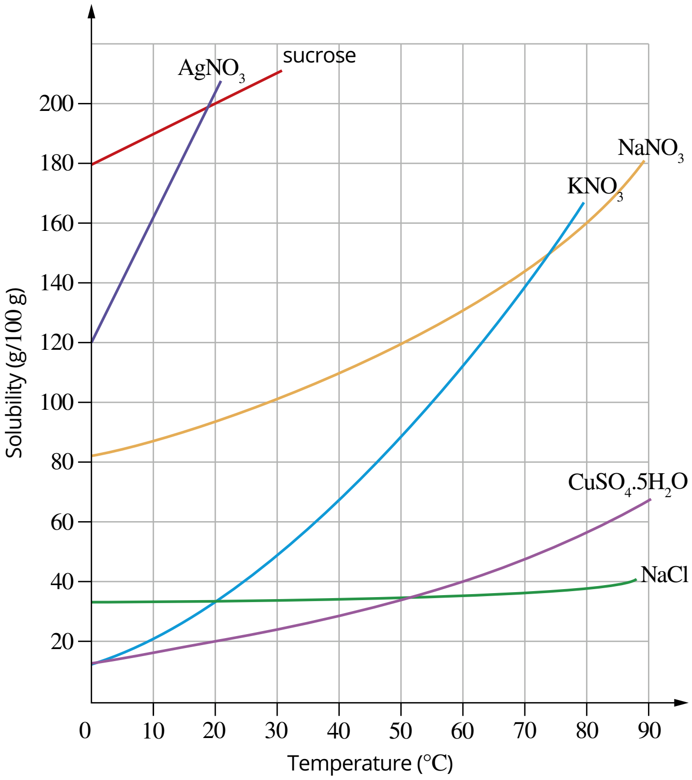
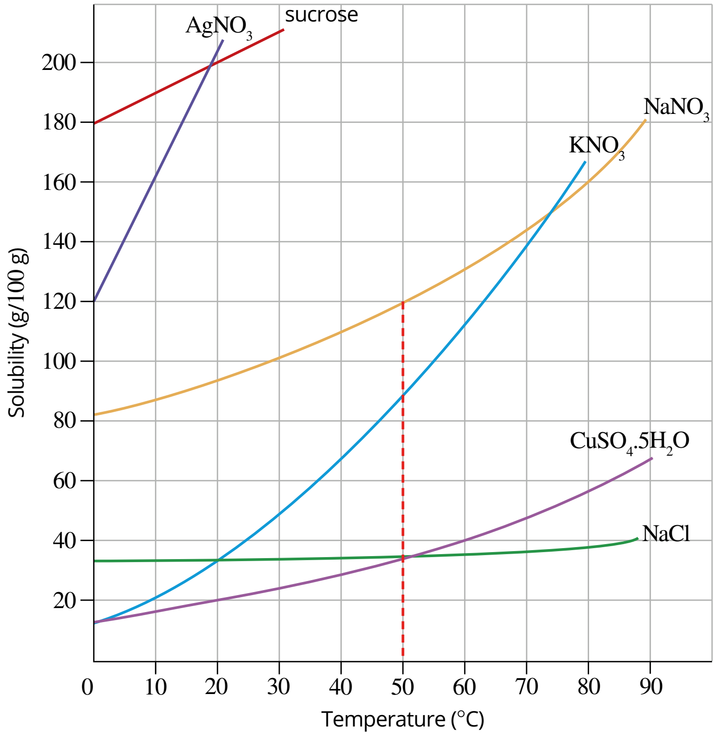
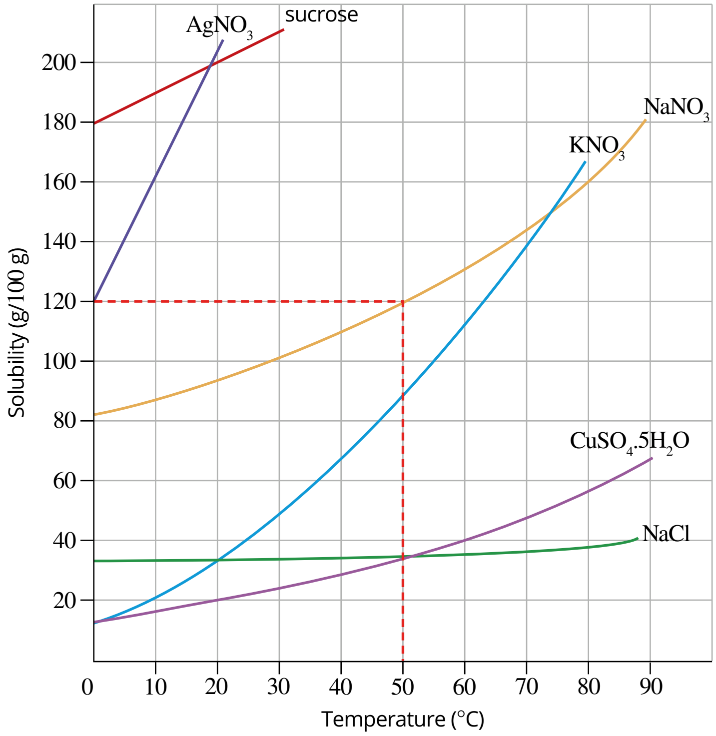


FIGURE 15.3.2 Solubility curves for some common solutes

Worked example 15.3.1

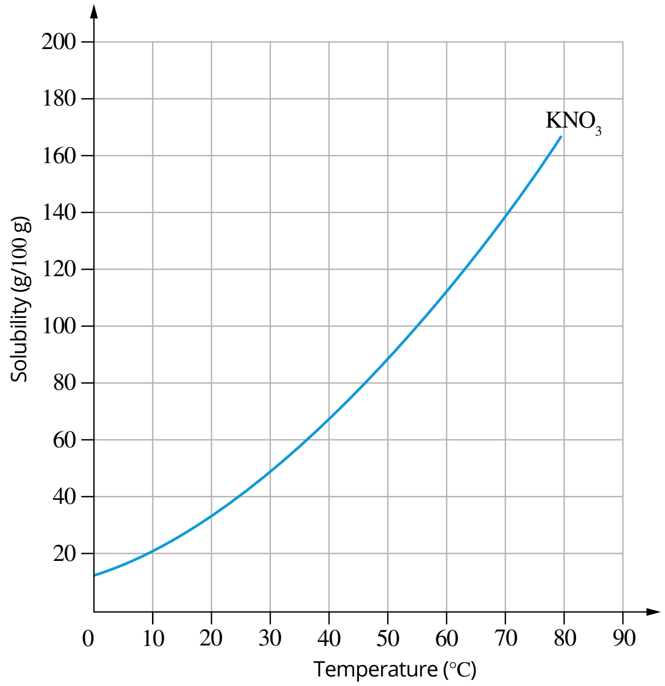
SOLUBILITY CURVE CALCULATIONS

<table>
<tr><td colspan="2">Use Figure 15.3.2 to determine how many grams of sodium nitrate ($NaNO_3$) will dissolve in 100 g of water at 50°C.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Points on a solubility curve show the maximum amount of the solute that will dissolve in 100 g water at a given temperature.
Draw a vertical line from the required temperature on the horizontal axis to intersect with the appropriate solubility curve.</td><td>Draw a vertical line from 50°C on the horizontal axis to intersect with the solubility curve for $NaNO_3$.
 </td></tr>
<tr><td>Draw a horizontal line from the intersection point of the vertical line drawn in the previous step. The point where this horizontal line intersects with the vertical axis will give the mass of dissolved substance.</td><td> </td></tr>
<tr><td>Read the solubility from the graph.</td><td>The horizontal line intersects with the vertical axis at 120 g. Therefore, the mass of $NaNO_3$ that will dissolve in 100 g of water at 50°C is 120 g.</td></tr>
</table>

➤ Try yourself 15.3.1

SOLUBILITY CURVE CALCULATIONS

Use the solubility curve below to determine how many grams of potassium nitrate (KNO_3) will dissolve in 100 g of water at 70°C.



PA
2.2.1

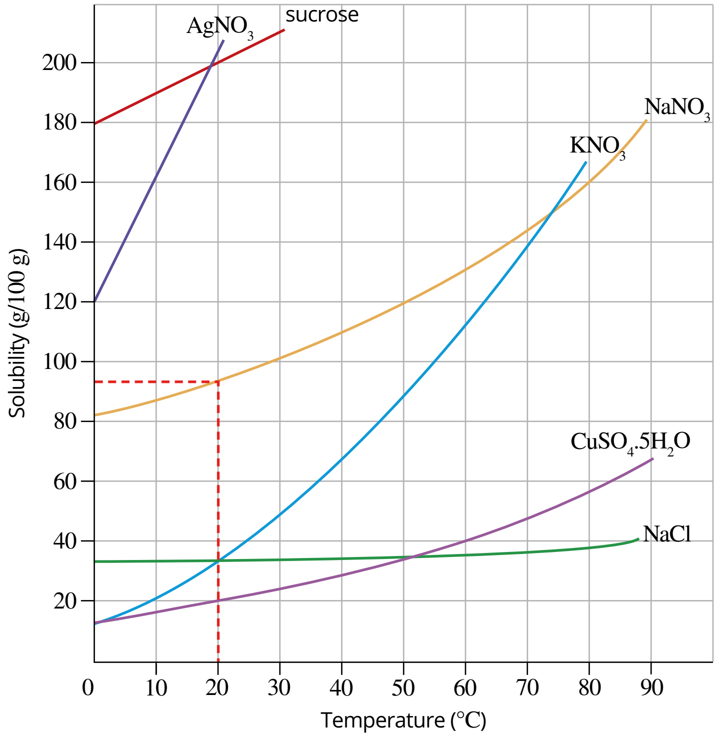
More calculations using solubility curves

Solubility curves allow you to directly calculate the mass of a compound that will dissolve in 100 g water at different temperatures. Worked example 15.3.2 shows how it is possible to use the solubility curves to calculate solubilities of compounds in quantities of water other than 100 g.

Worked example 15.3.2

SOLUBILITY CURVES

An 80 g sample of sodium nitrate ($NaNO_3$) is added to 200 mL of H_2O at 20°C.
Use Figure 15.3.2 to calculate how much more $NaNO_3$ must be added to make the solution saturated with $NaNO_3$ at 20°C.

Thinking	**Working**
Use the solubility curve to find the mass of solute in a saturated solution of 100 g of H_2O at the required temperature.	Draw a line from 20°C on the horizontal axis to the solubility curve for $NaNO_3$ and find the corresponding value on the vertical axis. The value is 92 g. 
Use the amount of solute that will dissolve in 100 g of H_2O to find the mass of solute to make a saturated solution in the mass of H_2O for this question.	The density of water is 1.0 g mL^{-1}, so 200 mL of water will weigh 200 g. So twice the mass of solute can dissolve in 200 g of water as in 100 g. $m(NaNO_3) = 2 \times 92$ $= 184$ g
To find out how much extra solute you need to add, find the difference between the mass of solute needed to make a saturated solution and how much has already been added.	80 g of $NaNO_3$ has already been added to 200 g H_2O. So the extra mass of $NaNO_3$ needed: $= 184 - 80$ $= 104$ g

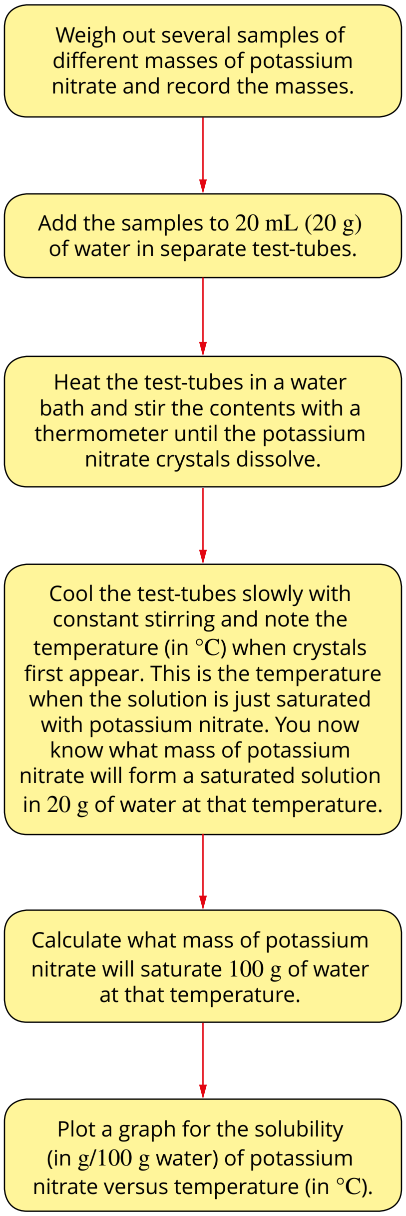
➤ Try yourself 15.3.2

SOLUBILITY CURVES

A 120 g sample of sodium nitrate ($NaNO_3$) is added to 300 mL of H_2O at 40°C.
Use Figure 15.3.2 to calculate how much more $NaNO_3$ must be added to make the solution saturated with $NaNO_3$ at 40°C.

Making your own solubility curve

Solubility curves can be developed relatively easily in a school laboratory. For example, a solubility curve for potassium nitrate in water can be derived in a group activity by carrying out the following steps shown in Figure 15.3.3.



Weigh out several samples of different masses of potassium nitrate and record the masses.

Add the samples to 20 mL (20 g) of water in separate test-tubes.

Heat the test-tubes in a water bath and stir the contents with a thermometer until the potassium nitrate crystals dissolve.

Cool the test-tubes slowly with constant stirring and note the temperature (in °C) when crystals first appear. This is the temperature when the solution is just saturated with potassium nitrate. You now know what mass of potassium nitrate will form a saturated solution in 20 g of water at that temperature.

Calculate what mass of potassium nitrate will saturate 100 g of water at that temperature.

Plot a graph for the solubility (in g/100 g water) of potassium nitrate versus temperature (in °C).

FIGURE 15.3.3 Steps used to derive a solubility curve

CRYSTALLISATION

You have probably seen rocky outcrops near the sea where salt crystals have formed from sea water in shallow pools, like those in Figure 15.3.4. The sea water may have been left there after a storm or high tide. As the water evaporates in the sun and wind, the concentration of the salt in the water increases until the solution becomes saturated. Further evaporation results in the salt coming out of solution and forming solid deposits. This process is called **crystallisation**.

FIGURE 15.3.4 White crystals of salt are left behind at the edge of a rock pool where sea water has evaporated.

Crystallisation from solutions occurs when an unsaturated solution becomes saturated and crystals form. There are two main ways in which this can happen.

Cooling a solution may reduce the solubility of a dissolved solute to the point where not all of the substance present is soluble. An example of this occurs when copper(II) sulfate crystals form after a hot solution of the compound is cooled. At lower temperatures, less of the compound can dissolve in the water and a point will be reached when the solution becomes saturated. Further cooling will result in crystals being formed.

Alternatively, crystallisation can occur as a result of the evaporation of solvent from a solution. For example, salt forms as water evaporates from sea water. This process produced the crystals of salt seen in Figure 15.3.4.

Factors affecting crystal growth

Several factors affect the formation of crystals from solutions. Some of these factors can affect the size of crystals formed.

Rate of cooling of a solution or molten compound

The rate of temperature change during cooling can affect the size of the crystals that are formed.

- Slow cooling generally produces large crystals.
- Rapid cooling produces smaller crystals.

This is true whether crystals are formed from a solution or as a molten compound cools below its melting point.

You can investigate this phenomenon in a laboratory. For example, a saturated solution of copper(II) sulfate can be cooled at different rates. Some of the solution could be cooled quickly by placing it in a refrigerator. Another sample could be cooled slowly by allowing it to sit in an insulated dish. The slower cooling results in larger crystals, as shown in Figure 15.3.5.

FIGURE 15.3.5 Large crystals of copper(II) sulfate formed by slowly cooling a saturated solution of the compound

Igneous rocks illustrate the effects of cooling rates on crystal size. The molten magma from the Earth's interior that forms these rocks can cool down quickly on the Earth's surface during volcanic activity, or very slowly underground. Figure 15.3.6 shows a section of granite, a rock that forms when magma cools below the Earth's surface. As it cools below the surface, the rate of temperature change is slow, so granite contains large crystals.

On the other hand, basalt is a rock that forms when magma flows out from a vent in the Earth and over surrounding land. This results in rapid cooling of the magma and formation of small crystals.

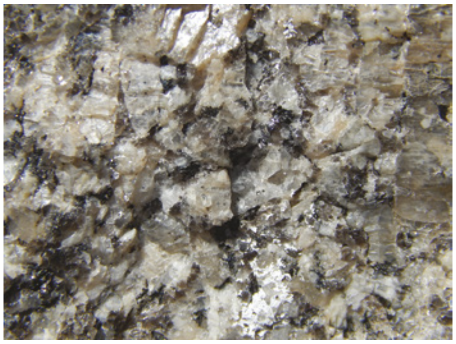


FIGURE 15.3.6 Large crystals (or grains) in granite

Rate of evaporation of solvent

If crystals are formed by evaporation, then the rate of evaporation of the solvent will affect the size of the crystals. Generally, a faster rate of evaporation results in smaller crystals. This can be investigated in the laboratory by placing solutions of a chemical in different situations where the rate of evaporation of the solvent varies; for example, in a cupboard and on a windowsill.

Nucleation

Crystallisation occurs more rapidly if a 'nucleus' is present in the solution on which the new crystals can form. A nucleus can be provided by adding a single seed crystal of the compound itself. Even a small solid impurity in the solution, such as a speck of dust or a scratch on the container wall, can act as a nucleus.

Nature of the compound

Different compounds produce different-shaped crystals. Figure 15.3.7 shows two different-shaped crystals. Sodium chloride crystals are cubic in shape; the crystals of alum (potassium aluminium sulfate, $KAl(SO_4)_2$) are octahedral.

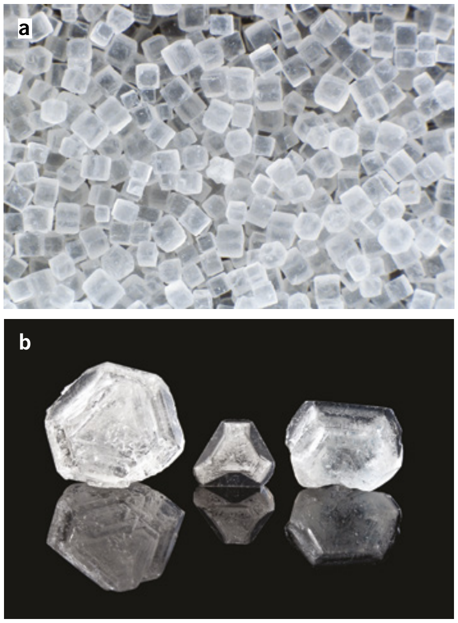


FIGURE 15.3.7 (a) Cubic crystals of table salt (sodium chloride) and (b) octahedral crystals of alum, $KAl(SO_4)_2$

Calculations involving crystallisation

You can predict how much of a compound will crystallise out of a solution as temperature decreases using a solubility curve. Worked example 15.3.3 shows how this can be done.

Worked example 15.3.3

CRYSTALLISATION

A 50 g sample of potassium nitrate (KNO_3) is dissolved in 100 g water at 40°C. Determine the mass of KNO_3 crystals that will form if the temperature is reduced to 20°C. You will need to refer to Figure 15.3.2 to complete this question.

Thinking	Working
Identify the mass of solute dissolved in the original solution.	Mass of KNO_3 in solution = 50 g
Find the maximum mass of solute that will remain dissolved in 100 g water at the final temperature.	The solubility curve shows that the maximum mass that will dissolve at 20°C is 32 g.
Calculate the mass of solute crystals that will form in the solution at the final temperature.	Mass of crystals formed = original mass − remaining mass = 50 – 32 = 18 g

➤ Try yourself 15.3.3

CRYSTALLISATION

A 200 g sample of sucrose is dissolved in 100 g water at 20°C. Determine the mass of sucrose crystals that will form if the temperature is reduced to 10°C. You will need to refer to Figure 15.3.2 to complete this question.

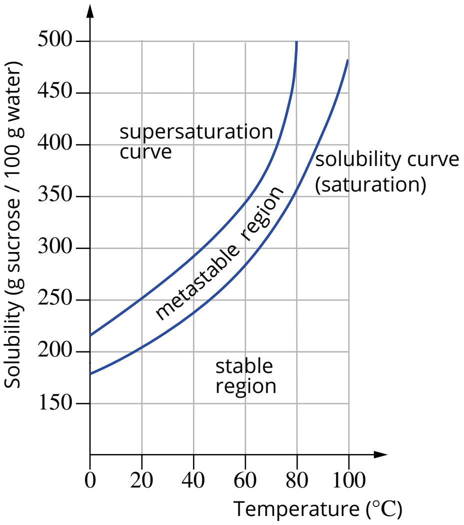


FIGURE 15.3.8 A solubility curve for sucrose used to predict the amount of sugar that will crystallise out under different conditions

Uses of solubility curves in the workplace

Sometimes it is important to know how much of a solute will precipitate out of a solution as it cools to a particular temperature, or if all the solute will stay dissolved. Two examples, one from industry and one from biological research, illustrate how solubility curves can be used.

Crystallising sugar from sugar cane syrup

In Australia, table sugar (sucrose) is prepared in large quantities from sugar cane. During processing, the cane is crushed to release juice. Impurities are removed and water is evaporated from the juice until the solution becomes saturated and the sugar begins to crystallise out. Chemists use solubility curves like the one in Figure 15.3.8 to predict the amount of sucrose that will crystallise at different temperatures and solubilities.

In this solubility curve, the area marked 'supersaturation curve' indicates the conditions under which crystals will form spontaneously. In the 'metastable region', seed crystals must be added to induce crystallisation. In areas below the saturation curve, the solution is unsaturated and crystals will not form.

Research on therapeutic proteins

Therapeutic proteins are compounds that are manufactured in the laboratory for use as pharmaceuticals. Synthetic insulin is an example of a therapeutic protein. Therapeutic proteins need to remain dissolved when injected into the body.

Most biological reactions and many industrial processes take place in solution. Laboratory workers in medical, pharmaceutical and industrial fields need to know the temperature ranges at which compounds remain dissolved. If the temperature of the solutions they are using do not stay within a certain range, then reactants may crystallise out and interfere with the desired reactions.

Solubility curves allow researchers to make accurate predictions about the solubility at different temperatures of the compounds under investigation. They can then ensure that work will not be disrupted by unexpected and unwanted crystallisation.

SOLUBILITY OF LIQUIDS AND GASES IN WATER

It is generally true that as temperature increases, the solubility of solids in solution increases. However, the solubility of liquids in other liquids does not show clear trends in solubility as the temperature changes.

Gases generally become less soluble as the temperature increases. The graph in Figure 15.3.9 shows the solubility of different gases in water at varying temperatures.

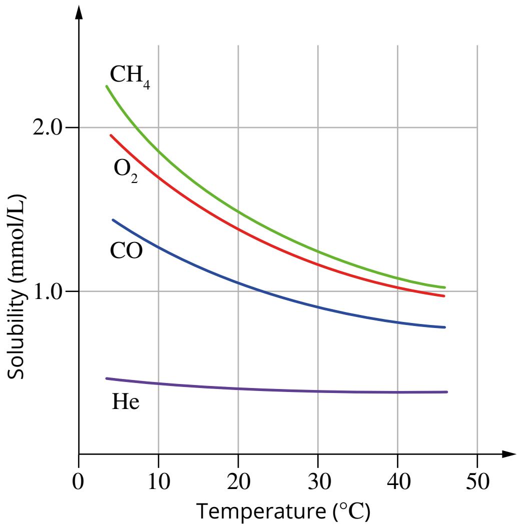


FIGURE 15.3.9 Solubility curves for some gases

Table 15.3.2 shows the solubility of some common gases in water at different temperatures. For each gas, it can be seen that as temperature increases, less gas is able to dissolve in water.

TABLE 15.3.2 Solubility of some gases in water at different temperatures

Gas	Solubility (g of gas per kg of water) at 0°C	Solubility (g of gas per kg of water) at 20°C	Solubility (g of gas per kg of water) at 60°C
oxygen	0.069	0.043	0.023
carbon dioxide	3.4	1.7	0.58
nitrogen	0.029	0.019	0.011
methane	0.040	0.023	0.011
ammonia	897	529	168

FIGURE 15.3.10 Bubbles of air appear as water is heated. This is because solubility of gases decreases at higher temperatures.

In Figure 15.3.10, the bubbles that appear when water is first heated are not steam (water vapour), but air bubbles. Some of the air that was dissolved in the water comes out of solution as the temperature increases and the gas becomes less soluble.

You may have noticed that soft drinks will go 'flat' more quickly if they are left standing in the sun on a hot day than if they are in a cold refrigerator (Figure 15.3.11). The dissolved carbon dioxide in the drink comes out of solution as the drink heats up.

FIGURE 15.3.11 Soft drinks left standing in the sun go 'flat' quickly.

The effect of temperature on the solubility of a gas can have environmental implications. If the temperature of the water in rivers, lakes and oceans increases even slightly, it will contain less dissolved oxygen. This can have serious consequences for oxygen-breathing aquatic organisms, such as the Australian Murray River cod in Figure 15.3.12.

Therefore, hot water from power stations and other industries must be cooled before it can be discharged into waterways. Even a small increase in water temperature can cause the oxygen concentration in the water to drop below levels that are necessary for aquatic life to survive.

FIGURE 15.3.12 Fish, such as this native Australian Murray River cod, are susceptible to lowered levels of oxygen in natural waters.

15.3 Review

SUMMARY

- Solubility is a measure of how much solute will dissolve in a given amount of solvent at a specified temperature.
- Solubility curves show how the solubility of a compound changes with temperature.
- Solubility curves can be used to calculate the amount of a substance that will dissolve in a given amount of solvent at a specified temperature.
- The solubility of most solids in water increases as the temperature increases.
- Crystallisation occurs when a solution is cooled to the point where it becomes saturated with solute and the solute begins to form solid crystals.
- As the temperature of a solution increases, the solubility of solids generally increases and the solubility of gases decreases. Liquids show no overall trend in solubility with temperature.
- Crystallisation can occur from a solution when:
  - the temperature of the solution decreases
  - solvent evaporates.
- Factors that can affect crystallisation, including crystal size, are:
  - the rate of cooling of a solution
  - the rate of evaporation of a solvent
  - nucleation
  - the nature of the compound.
- Solubility curves can be used to predict the amount of a compound that will precipitate from solution at a particular temperature.
- Solubility curves can be used as predictive tools in different fields of industry and research.

KEY QUESTIONS

Retrieval

1 Determine whether the following solutions are unsaturated, saturated or supersaturated. Use Table 15.3.1 on page 467 to help you.

a a solution containing 35 g of sodium chloride dissolved in 100 g of water at 18°C

b a solution containing 0.00045 g of calcium carbonate dissolved in 100 g of water at 18°C

c a solution containing 110 g of sugar dissolved in 50 g of water at 18°C

2 Identify the name given to a solution in which no more solute can be dissolved.

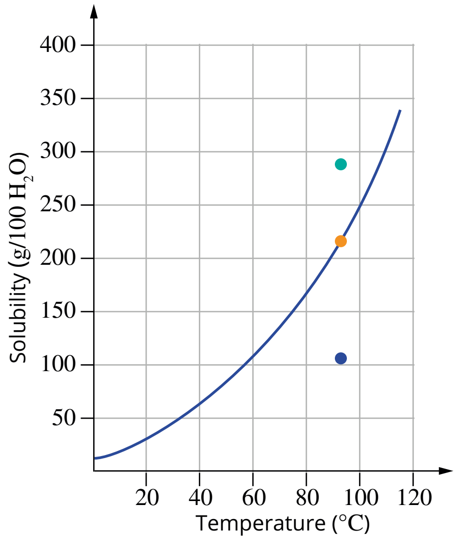
3 Describe the effect of slow cooling on the formation of crystals in solution.

4 List the information that can be determined from a solubility curve.

Comprehension

5 Explain the general trend seen for solubility of solutes in solution with changes in temperature.

6 Look at the graph below and determine which coloured points correspond to saturated, unsaturated and supersaturated solutions of potassium nitrate.



7 Summarise the steps necessary to create a solubility curve for a solute in water.

Analysis

Refer to Figure 15.3.2 on page 468 to answer Questions **8** to **11**.

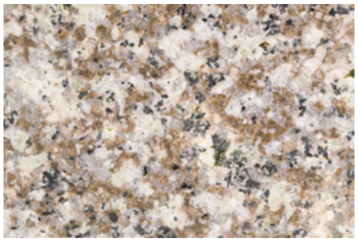
8 A 140 g sample of potassium nitrate (KNO_3) is dissolved in 100 g of water at 70°C. The temperature of the solution is then cooled to 55°C.
Calculate the mass of KNO_3 that will crystallise out of solution.

9 Calculate the mass of crystals that will form when a solution containing 60 g of $CuSO_4.5H_2O$ in 100 g of water at 90°C is cooled to 60°C.

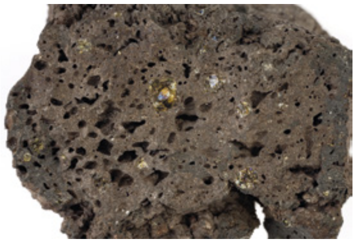
10 Calculate the mass of solute that could dissolve in the following situations.
  a $CuSO_4.5H_2O$ in 100 g of water at 60°C
  b $AgNO_3$ in 100 g of water at 10°C
  c KNO_3 in 50 g of water at 24°C

11 Determine what mass of the following compounds will dissolve in 200 g of water at 30°C to form a saturated solution.
  a $CuSO_4.5H_2O$
  b $NaNO_3$
  c NaCl

12 Referring to the images below of the crystals seen in granite and basalt rocks, determine why granite crystals are a different size from basalt crystals.



Crystals in granite rocks



Crystals in basalt rocks

13 Predict what would happen to a supersaturated solution if poured into a flask containing a chip or crack in the base of the flask.

14 Global climate scientists are concerned about the effect of global warming on ocean temperatures. Predict the effect of increasing the temperature of sea water on dissolved respiratory gases inhaled and exhaled by marine fauna.

15 Identify the two least soluble gases at 20°C.

16 Use Figure 15.3.9 on page 474 to determine the solubility of helium in mmol L^{-1} at 30°C.

17 A 100 g sample of sodium nitrate is dissolved in 100 g of water at 40°C. Calculate how many more grams of sodium nitrate would need to be added to form a saturated solution.

15.4 Precipitation reactions

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- apply the solubility laws to determine the products of reactions
- determine the products produced in precipitation reactions
- determine ions dissolved in solution to write full chemical and ionic equations for a variety of precipitation reactions.

PRECIPITATION REACTIONS

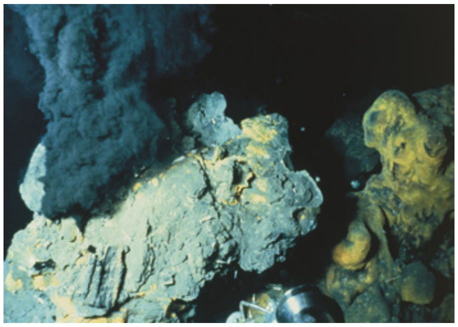


FIGURE 15.4.1 Undersea hydrothermal vents release superheated water containing sulfides, which form precipitates with metal ions.

In the previous module you were introduced to solubility. You also looked at the solubilities of different ions in aqueous solutions. A **precipitation reaction** occurs if ions in solution combine to form a new compound that is insoluble in water. The insoluble compound formed in such a reaction is called a **precipitate**.

Precipitation reactions occur naturally in undersea hydrothermal vents. The vents release superheated solutions containing sulfides, which then combine with metal ions to form precipitates of mineral sulfides, creating the chimney-like structures seen in Figure 15.4.1. The areas around these chimneys are biologically rich, often hosting complex communities fuelled by the chemicals dissolved in the vent fluids.

Precipitation reactions are used to remove minerals from drinking water, to remove heavy metals from wastewater and in the purification plants of reservoirs.

In this module you will examine the process that takes place during precipitation reactions.

Deducing the identity of a precipitate

FIGURE 15.4.2 Mixing aqueous solutions of sodium chloride and silver nitrate produces a solid, called a precipitate.

When a colourless solution of silver nitrate is mixed with a colourless solution of sodium chloride, a white solid is formed, as shown in Figure 15.4.2. The white solid is the precipitate.

To understand what happens in the reaction between silver nitrate and sodium chloride, you need to identify the ions present in the reactant solutions and how they interact with each other.

- In the silver nitrate solution, there are silver ions (Ag^+) and nitrate ions (NO_3^-).
- In the sodium chloride solution, there are sodium ions (Na^+) and chloride ions (Cl^-).
- When one solution is added to the other, the mixture formed will contain all of the ions.

In both of these solutions, all the ions are moving around independently. As the ions move in the solution, they will collide with one another. If positive and negative ions collide, they may join together to form a new insoluble precipitate.

Two new combinations of positive and negative ions are possible:

- sodium and nitrate ions (to form sodium nitrate)
- silver and chloride ions (to form silver chloride).

Tables 15.2.3 and 15.2.4 on page 463 indicate that sodium nitrate is soluble in water, but that silver chloride is insoluble. Therefore, the precipitate must be silver chloride.

The process for the precipitation reaction between sodium chloride and silver nitrate is shown in Figure 15.4.3. When the hydrated Ag^+ and Cl^- ions come into contact, the attraction between the ions is greater than that of the ions for the water molecules. An ionic lattice of AgCl is formed.

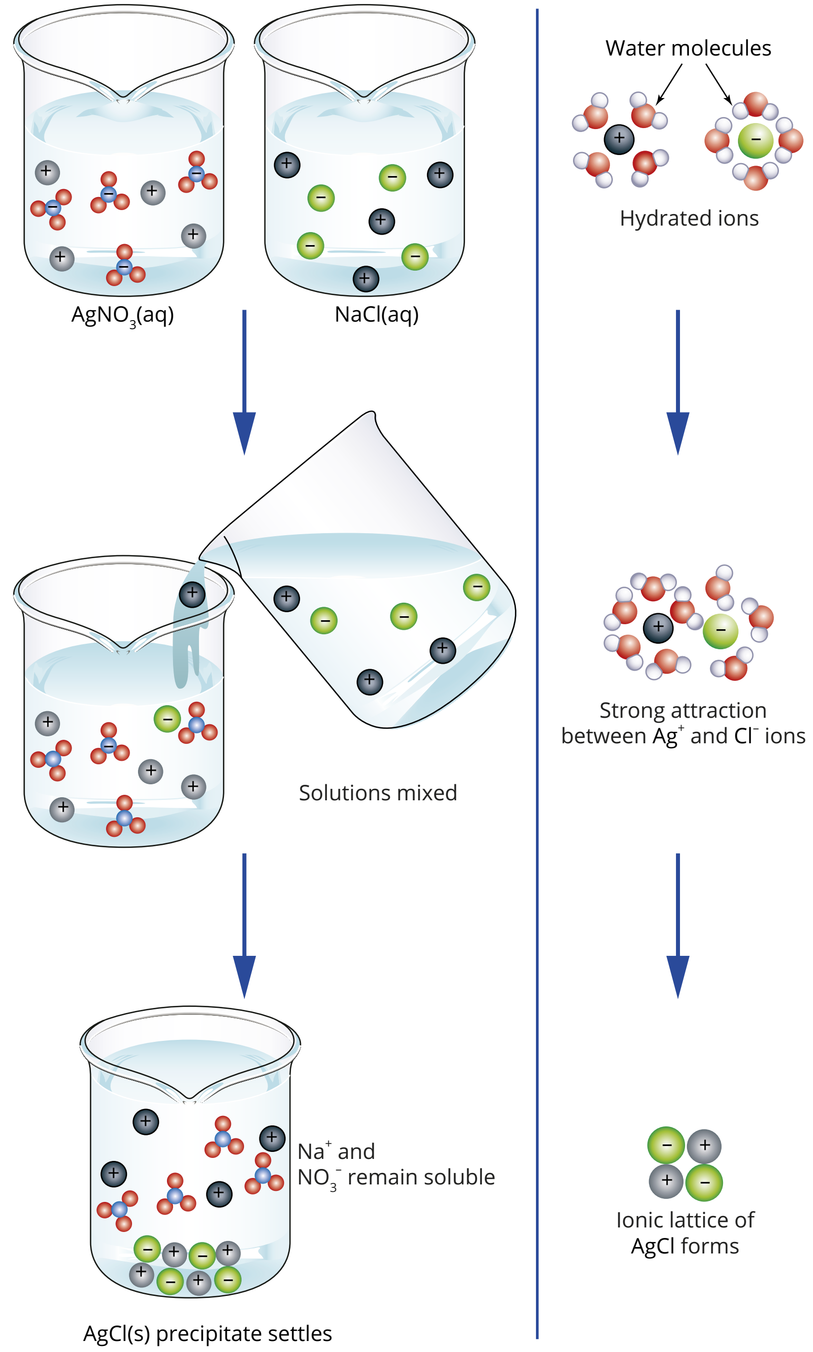


FIGURE 15.4.3 Pictorial representation of mixing aqueous solutions of sodium chloride and silver nitrate to produce a precipitate of silver chloride

There is a simple way that allows you to work out which compound will form the precipitate in a reaction between two ionic solutions.

- Write down the formula for the positive ion of one of the compounds, followed by its negative ion. Repeat the process for the second compound. For example, Ag^+, NO_3^-, Na^+, Cl^-.
- Then draw two lines. The first line joins the positive ion of the first solution to the negative ion of the second. The second line joins the negative ion of the first solution to the positive ion of the second (Figure 15.4.4).

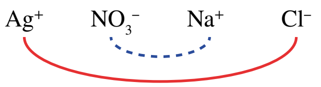


FIGURE 15.4.4 Working out which compound is the precipitate in a reaction

- Finally, use solubility tables to work out which of the two combinations of ions will result in an insoluble compound. This will be the precipitate. The other ions will remain in solution.

It is important for chemists to be able to predict whether a precipitate will form in a reaction and what this precipitate will be. Worked example 15.4.1 takes you through the process of predicting the products of a precipitation reaction.

Worked example 15.4.1

PREDICTING THE PRODUCTS OF A PRECIPITATION REACTION

Determine what precipitate, if any, will be produced when solutions of potassium hydroxide and lead(II) nitrate are added together. You will need to refer to the solubility tables (Tables 15.2.3 and 15.2.4) to complete this question.

Thinking	Working
Identify which ions are produced by each of the ionic compounds in the mixture.	K^+(aq), OH^-(aq), Pb^{2+}(aq) and NO_3^-(aq)
Identify which two new combinations of positive and negative ions are possible in the mixture of the solutions.	K^+(aq) and NO_3^-(aq) Pb^{2+}(aq) and OH^-(aq)
Use the solubility tables to check which, if any, of these combinations will produce an insoluble compound.	Compounds containing potassium ions are usually soluble, so potassium nitrate will not form a precipitate. Compounds containing hydroxide ions are usually insoluble, so lead(II) hydroxide will form a precipitate.

➤ Try yourself 15.4.1

PREDICTING THE PRODUCTS OF A PRECIPITATION REACTION

Determine what precipitate, if any, will be produced when solutions of sodium sulfide (Na_2S) and copper(II) nitrate ($Cu(NO_3)_2$) are added together. You will need to refer to the solubility tables (Tables 15.2.3 and 15.2.4) to complete this question.

Writing equations for precipitation reactions

Now that you can identify the precipitate that forms in a reaction, the next step is to show the complete reaction by writing a chemical equation.

The reaction between silver nitrate and sodium chloride solutions can be summarised in words as:

silver nitrate solution + sodium chloride solution → silver chloride solid + sodium nitrate solution

An alternative representation is an equation that uses formulas. This type of equation is called a **full equation**. The complete formulas of all the reagents and products are shown in the reaction. The state symbols for each of the species in the chemical reaction must be shown:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Although 'sodium nitrate' or '$NaNO_3$(aq)' is written as a product in these equations, the sodium ions and nitrate ions are not combined with each other. You will remember from Module 15.2 that soluble ionic compounds form dissociated ions in solution and are free to move.

These ions are present at the start of the reaction and they are still there, as separate ions, at the end of the reaction. Because the ions have not been involved in forming a precipitate, they are said to be **spectator ions**.

Spectator ions do not undergo a chemical change in the reaction. In a precipitation reaction, they will always start as aqueous (aq) ions and will remain as aqueous ions after the reaction is complete.

Worked example 15.4.2 looks at the process of writing full equations for precipitation reactions and identifying spectator ions.

Worked example 15.4.2

WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Write a balanced equation for the reaction between iron(III) nitrate and sodium sulfide, in which the precipitate is iron(III) sulfide. Identify the spectator ions in this reaction.

Thinking	Working
Write an incomplete, unbalanced equation showing the reactants and the precipitate product. Include symbols of state.	$Fe(NO_3)_3(aq) + Na_2S(aq) \rightarrow Fe_2S_3(s)$
Add to the equation above the formula of the other compound formed in the reaction.	$Fe(NO_3)_3(aq) + Na_2S(aq) \rightarrow Fe_2S_3(s) + NaNO_3(aq)$
Balance the equation.	$2Fe(NO_3)_3(aq) + 3Na_2S(aq) \rightarrow Fe_2S_3(s) + 6NaNO_3(aq)$
Write the formulas of the ions that do not form a precipitate in the reaction. These are the spectator ions.	$Na^+(aq)$ and $NO_3^-(aq)$ are spectator ions.

➤ Try yourself 15.4.2

WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Write a balanced equation for the reaction between copper(II) sulfate and sodium hydroxide, in which the precipitate is copper(II) hydroxide. Identify the spectator ions in this reaction.

Writing ionic equations for precipitation reactions

WS 2.2.4

The essential feature of the reaction between silver nitrate and sodium chloride is the combination of silver ions and chloride ions to form a precipitate. This reaction can be summarised in an **ionic equation**:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Note that spectator ions are not included in an ionic equation. Only the species that change are included.

A simple way to write an ionic equation for a precipitation reaction is shown in the following example of the reactions between solutions of aluminium nitrate and sodium sulfide. The full equation for the precipitation reaction is:

$$2Al(NO_3)_3(aq) + 3Na_2S(aq) \rightarrow Al_2S_3(s) + 6NaNO_3(aq)$$

To write an ionic equation:

- Write down the formula of the precipitate on the right-hand side of the page. Include a symbol of state. Place an arrow to the left of it.

$$\rightarrow Al_2S_3(s)$$

- To the left of this formula, add the formulas of the ions that form the precipitate, in the ratio shown by the formula of the precipitate. Include state symbols.

$$2Al^{3+}(aq) + 3S^{2-}(aq) \rightarrow Al_2S_3(s)$$

- Check that the equation is balanced.

Worked example 15.4.3 looks at the process of writing ionic equations for precipitation reactions.

Worked example 15.4.3

WRITING IONIC EQUATIONS FOR PRECIPITATION REACTIONS

Write a balanced ionic equation for the reaction between silver nitrate and potassium carbonate.	
Thinking	**Working**
Write an incomplete, unbalanced equation showing the reactants and the precipitate product. Include symbols of state.	$AgNO_3(aq) + K_2CO_3(aq) \rightarrow Ag_2CO_3(s) + KNO_3(aq)$
Write down the formula for the precipitate on the right-hand side of the page and place an arrow to the left of it.	$\rightarrow Ag_2CO_3(s)$
Add the formulas of the ions that form the precipitate, in the ratio shown by the formula of the precipitate.	$2Ag^+ + CO_3^{2-} \rightarrow Ag_2CO_3(s)$
Add states to the reactant side of the equation.	$2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$

➤ Try yourself 15.4.3

WRITING IONIC EQUATIONS FOR PRECIPITATION REACTIONS

Write a balanced ionic equation for the reaction between magnesium chloride and sodium sulfide.

15.4 Review

SUMMARY

- A precipitation reaction occurs when two solutions of compounds are mixed and a solid product is formed. The solid product is called a precipitate.
- Solubility tables can be used to predict which compound, if any, will precipitate in a reaction.
- Ions that are not directly involved in the formation of the precipitate are called spectator ions.
- Full and ionic equations can be written for precipitation reactions.
- Ionic equations do not include spectator ions.

KEY QUESTIONS

Retrieval

1 Use the solubility tables on page 463 to identify the precipitate formed, if any, when the following solutions are mixed.
  - **a** silver nitrate and potassium carbonate
  - **b** potassium hydroxide and lead(II) nitrate
  - **c** magnesium chloride and sodium sulfide
  - **d** sodium chloride and iron(II) sulfate

2 **a** Name the precipitate formed when aqueous solutions of the following compounds are mixed.
  - **i** K_2S and $MgCl_2$
  - **ii** $CuCl_2$ and $AgNO_3$
  - **iii** KOH and $AlCl_3$
  - **iv** $MgSO_4$ and NaOH

b Write a fully balanced chemical equation for each reaction.

Comprehension

3 Represent the process of precipitation with a diagram similar to that shown in Figure 15.4.4 on page 479 for the reaction between lead(II) nitrate and sodium hydroxide.

4 Show, using a full chemical equation, how two solutions can produce a solid precipitate using barium chloride and sodium sulfate.

Analysis

5 Distinguish between a complete chemical equation and an ionic equation using the precipitation reaction between ammonium sulfate and barium chloride as an example.

6 Categorise which of the following ions when combined with Fe^{2+} would form a precipitate.
  - **a** SO_4^{2-}
  - **b** PO_4^{3-}
  - **c** Br^-
  - **d** S^{2-}

7 Using your knowledge of solubility, infer which positive ions are most likely to act as spectator ions in a precipitation reaction.

15.5 Concentration of solutions

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- distinguish between the terms solute, solvent, solution and concentration
- calculate the concentration of different solutions using multiple units
- conduct a number of calculations to convert between the different units of concentration.

The concentration of a solution describes the relative amount of solute and solvent present in the solution. A solution in which the ratio of solute to solvent is high is said to be concentrated. Cordial that has not had any water added to it is an example of a **concentrated solution**. A solution in which the ratio of solute to solvent is low is said to be a **dilute solution**. A quarter of a teaspoon of sugar dissolved in a litre of water will produce a dilute solution that tastes slightly sweet.

'Concentrated' and 'dilute' are general terms. However, sometimes you need to know the actual concentration of a solution—the exact ratio of solute to solvent. The use of exact solution concentrations is important in the prescription of medicines, chemical manufacturing and chemical analysis.

UNITS OF CONCENTRATION

Chemists use different measures of concentration depending on the situation. The most common measures describe the amount of solute in a given amount of solution.

Here are some examples of units of concentration.

- mass of solute per litre of solution (grams per litre, $g\,L^{-1}$)
- moles of solute per litre of solution (moles per litre, $mol\,L^{-1}$)
- parts per million (ppm)
- percentage by mass.

Units of concentration measured in this way have two parts.

- The first part gives information about how much solute there is.
- The second part gives information about how much solution there is.

For example, if a solution contains sodium chloride (NaCl) with a concentration of $17\,g\,L^{-1}$, then (first part) 17 g of NaCl is dissolved in every (second part) 1 L of the solution.

You will now look at how to perform calculations of concentration using different units.

Concentration in grams per litre ($g\,L^{-1}$)

The concentration of a solution in grams per litre ($g\,L^{-1}$) indicates the mass, in grams, of solute dissolved in 1 litre of the solution. For example, if the concentration of sodium chloride in sea water is $20\,g\,L^{-1}$, this means that in 1 L of sea water there is 20 g of sodium chloride. A formula used to calculate the concentration in $g\,L^{-1}$ is:

$$\text{concentration } (g\,L^{-1}) = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in L)}}$$

Worked example 15.5.1

CALCULATING CONCENTRATION IN $g\,L^{-1}$ (GRAMS PER LITRE)

Calculate the concentration, in $g\,L^{-1}$, of a solution containing 8.00 g of sodium chloride in 500 mL of solution.	
Thinking	**Working**
Change the volume of solution so it is expressed in litres.	$500\,mL = \frac{500\,mL}{1000\,mL} = 0.500\,L$
Calculate the concentration in $g\,L^{-1}$.	$c = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in L)}}$ $= \frac{8.00\,g}{0.500\,L}$ $= 16.0\ g\,L^{-1}$

➤ Try yourself 15.5.1

CALCULATING CONCENTRATION IN $g\,L^{-1}$ (GRAMS PER LITRE)

Calculate the concentration, in $g\,L^{-1}$, of a solution containing 5.00 g of glucose in 250 mL of solution.

Concentration in milligrams per litre ($mg\,L^{-1}$)

If, for example, the concentration of cadmium in a sample of wastewater was $5\,mg\,L^{-1}$, this means that in 1 litre of the wastewater there will be 5 mg of cadmium. Concentrations in $mg\,L^{-1}$ may be calculated using the formula:

$$\text{concentration } (mg\ L^{-1}) = \frac{\text{mass of solute (in mg)}}{\text{volume of solution (in L)}}$$

A concentration of solution measured in units of ppm has the same value as concentration measured in $mg\,kg^{-1}$. Since 1 litre of solution has a mass of almost exactly 1 kilogram, then a concentration measured in $mg\,L^{-1}$ will also have the same value in units of ppm, i.e.

$$1\text{ ppm} = 1\,mg\,kg^{-1} = 1\,mg\,L^{-1}$$

For example, the concentration, in $mg\,L^{-1}$ and ppm, of 2.0 L of a sugar solution containing 0.050 g of sugar can be calculated as follows:

$$0.050\,g \text{ of sugar} = 50\,mg$$

$$\begin{aligned}\text{concentration } (mg\ L^{-1}) &= \frac{50}{2.0}\\ &= 25\,mg\,L^{-1}\\ &= 25\,ppm\end{aligned}$$

Concentrations in $mg\,L^{-1}$ and in ppm are equivalent.

Concentration in parts per million (ppm)

When very small quantities of solute are dissolved to form a solution, the concentration is often measured in **parts per million (ppm)**. For example, the concentration of mercury in fish that is suitable for consumption is usually expressed in parts per million. The maximum concentration allowed for sale in Australia is 1 ppm (Figure 15.5.1).

In simple terms, the concentration in parts per million can be thought of as the mass in grams of solute dissolved in 1 000 000 g of solution. This can also be expressed as the mass in milligrams of solute dissolved in 1 kg of solution, because there are 1 million milligrams in a kilogram.

FIGURE 15.5.1 Fish sold in Australia must have no more than 1 ppm of mercury (Australia New Zealand Food Standards Code).

For example, a solution of sodium chloride that has a concentration of 154 ppm contains 154 mg of sodium chloride dissolved in 1 kg of solution. A formula used to calculate the concentration of a solution in ppm is:

$$\text{concentration (ppm)} = \frac{\text{mass of solute (in mg)}}{\text{mass of solution (in kg)}}$$

Worked example 15.5.2

CALCULATING CONCENTRATION IN PARTS PER MILLION (PPM)

A saturated solution of calcium carbonate was found to contain 0.0198 g of calcium carbonate dissolved in 2000 g of solution. Calculate the concentration, in ppm, of calcium carbonate in the saturated solution.

Thinking	**Working**
Calculate the mass of solute in mg. Remember: mass (in mg) = mass (in g) × 1000	Mass of solute (calcium carbonate) in mg: 0.0198×1000 $= 19.8\,\text{mg}$
Calculate the mass of solution in kg. Remember: $\text{mass (in kg)} = \frac{\text{mass (in g)}}{1000}$	$\text{Mass of solute (in kg)} = \frac{2000}{1000}$ $= 2.000\,\text{kg}$
Calculate the concentration of the solution in mg kg^{-1}. This is the same as concentration in ppm.	Concentration of calcium carbonate in ppm $= \frac{\text{mass of solute (in mg)}}{\text{mass of solution (in kg)}}$ $= \frac{19.8}{2.000}$ $= 9.90\,\text{mg kg}^{-1}$ $= 9.90\,\text{ppm}$

Try yourself 15.5.2

CALCULATING CONCENTRATION IN PARTS PER MILLION (PPM)

A sample of tap water was found to contain 0.0537 g of NaCl per 250.0 g of solution. Calculate the concentration of NaCl in parts per million (ppm).

Concentration in parts per billion (ppb)

For concentrations that are very small, the unit of **parts per billion (ppb)** can be used. If the concentration of a solution of a pesticide is 5 ppb, this means that in 1 billion grams of the solution, there will be 5 g of pesticide. Concentrations in units of ppb have the same value as concentrations in units of μg (microgram) per kg. A formula for calculating concentration in ppb is:

$$\text{concentration (ppb)} = \frac{\text{mass of solute (in μg)}}{\text{mass of solution (in kg)}}$$

Other units of concentration

You might have noticed symbols such as m/m, v/v or m/v on the labels of some foods, drinks and pharmaceuticals. These symbols represent other concentration units based on masses and volumes of solutes and solutions. These are useful in practical situations because people are familiar with these quantities (Figure 15.5.2).

FIGURE 15.5.2 Consumer products from hardware stores and supermarkets show a wide range of concentration units on their labels.

Percentage by mass (m/m)

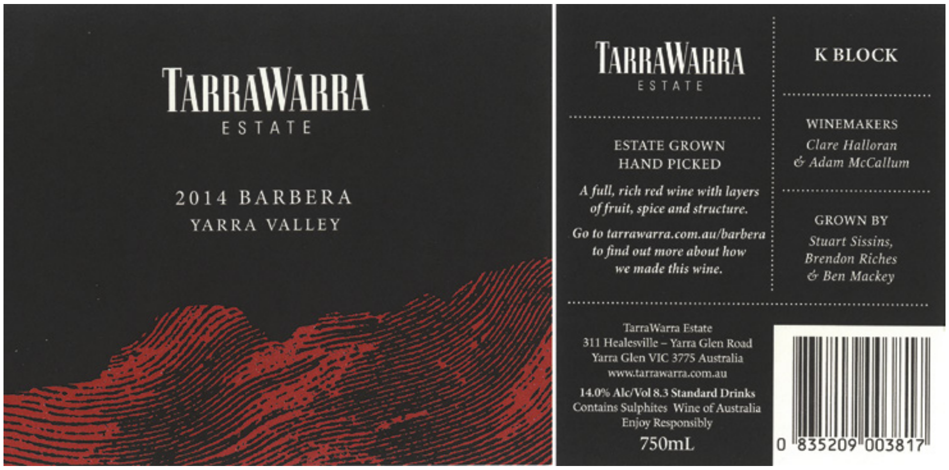
Percentage by mass describes the mass of solute, measured in grams, present in 100 g of the solution.

Normal saline solution for washing contact lenses has a concentration of 0.9%(w/w). This can also be written as 0.9%(m/m). The abbreviation w/w indicates that the percentage is based on the weights or, more correctly, masses of both solute and solution. A concentration of 0.9% indicates that there is 0.9 g of sodium chloride dissolved in 100 g of solution.

Percentage by volume (v/v)

The abbreviation v/v indicates that the percentage is based on volumes of both solute and solution. The same units must be used to record both volumes. Percentage by volume is a more convenient unit to use than w/w when the solute is a liquid.

Just like percentage by mass, percentage by volume is frequently expressed as volume per 100 mL of solution. For example, the wine label in Figure 15.5.3 shows 14.0% alc./vol. This means the wine contains 14.0% alcohol (ethanol) by volume (14.0%(v/v)). There will be 14.0 mL of alcohol in 100 mL of the wine.



TARRAWARRA
ESTATE

2014 BARBERA
YARRA VALLEY

TARRAWARRA
ESTATE

ESTATE GROWN
HAND PICKED

A full, rich red wine with layers of fruit, spice and structure.

Go to tarrawarra.com.au/barbera to find out more about how we made this wine.

K BLOCK

WINEMAKERS
Clare Halloran
& Adam McCallum

GROWN BY
Stuart Sissins,
Brendon Riches
& Ben Mackey

TarraWarra Estate
311 Healesville – Yarra Glen Road
Yarra Glen VIC 3775 Australia
www.tarrawarra.com.au

14.0% Alc/Vol 8.3 Standard Drinks
Contains Sulphites Wine of Australia
Enjoy Responsibly
750mL

0 835209 003817

FIGURE 15.5.3 The label from a bottle of Yarra Valley wine shows the alcohol concentration in units of %(v/v).

For example, a 200 mL glass of champagne contains 28 mL of alcohol. The concentration as %(v/v) of alcohol in this solution can be calculated as follows:

$$
\begin{aligned}
\text{concentration } (\%\ \text{v/v}) &= \frac{\text{volume of solute (in mL)}}{\text{volume of solution (in mL)}} \times 100\% \\
&= \frac{28}{200} \times 100\% \\
&= 14\%(\text{v/v})
\end{aligned}
$$

Percentage mass/volume (m/v)

Percentage mass/volume describes the mass of solute, measured in grams, present in 100 mL of the solution.

For example, if a solution of plant food contains a particular potassium compound at a concentration of 3%(m/v), this indicates that there is 3 g of potassium in 100 mL of solution.

FIGURE 15.5.4 These bottles contain solutions with concentrations shown in units of M (mol L^{-1}).

Throughout this section, $c = \frac{n}{V}$ is used to calculate molarity, as molarity is a unit of concentration.

However, using $M = \frac{n}{V}$ would also be correct. Both will produce the same answer, but what is important is that you ensure your answer uses the unit M.

MOLAR CONCENTRATION

Another unit of concentration that is commonly used by chemists is mol L^{-1}. Concentrations expressed in moles per litre (mol L^{-1}) allow chemists to compare the number of moles of atoms, molecules or ions present in a given volume of solution.

For example, a 0.1 mol L^{-1} solution of hydrochloric acid would contain 0.1 moles of HCl in 1 L of the solution. (A bottle of HCl of this concentration is shown in Figure 15.5.4.)

The concentration, in mol L^{-1}, of a solution can be calculated as follows:

$$\text{concentration (mol L}^{-1}) = \frac{\text{number of moles in the solute}}{\text{volume of solution (in L)}}$$

$$c = \frac{n}{V}$$

where c is the concentration (mol L^{-1}), n is the amount (mol) and V is the volume (L).

The concentration of a solution in moles per litre is often referred to as the **molarity**, or molar concentration, of the solution, and is given the unit M. As a result, sometimes the formula is written like this:

$$\text{molarity} = \frac{\text{moles of solute } (n)}{\text{volume of solution } (V)}$$

$$M = \frac{n}{V}$$

A solution containing 1 mole of solute dissolved in 1 litre of solution can therefore be described in several different ways.

- It has a concentration of 1 mole per litre.
- It has a concentration of 1 mol L^{-1}.
- It is 1 molar.
- It is 1 M.
- It has a molarity of 1 M.

Worked example 15.5.3

CALCULATING MOLAR CONCENTRATIONS (MOLARITY)

Calculate the molar concentration of a solution that contains 0.105 mol of potassium nitrate dissolved in 200 mL of solution.

Thinking	**Working**
Convert the given volume to litres.	$V(KNO_3) = \frac{200}{1000}$ $= 0.200$ L
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(KNO_3) = \frac{n}{V}$ $= \frac{0.105}{0.200}$ $= 0.525$ mol L^{-1} or 0.525 M

➤ Try yourself 15.5.3

CALCULATING MOLAR CONCENTRATIONS (MOLARITY)

Calculate the molar concentration of a solution that contains 0.24 mol of glucose dissolved in 500 mL of solution.

Calculating molarity given the mass of solute

Sometimes you will know the mass of a solute and volume of solution and want to calculate the molarity. Two main calculations are involved.

1 Calculate the number of moles of solute from its mass.
2 Calculate the concentration using the number of moles and the volume (in litres).

Worked example 15.5.4

CALCULATING MOLARITY GIVEN THE MASS OF SOLUTE

Calculate the concentration, in mol L^{-1}, of a solution that contains 16.8 mg of silver nitrate ($AgNO_3$) dissolved in 150 mL of solution.

Thinking	**Working**
Convert the volume to litres.	$V(AgNO_3) = \frac{150}{1000}$ $= 0.150\ L$
Convert the mass to grams.	$m(AgNO_3) = \frac{16.8}{1000}$ $= 0.0168\ g$
Calculate the molar mass of the solute. To do this, add up the atomic masses of all the atoms in the compound.	$M(AgNO_3) = 107.87 + 14.01 + (3 \times 16.00)$ $= 169.88\ g\ mol^{-1}$
Calculate the number of moles of solute using the formula: $n = \frac{m}{M}$	$n(AgNO_3) = \frac{m}{M}$ $= \frac{0.0168}{169.9}$ $= 169.89 \times 10^{-5}\ mol$
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(AgNO_3) = \frac{n}{V}$ $= \frac{9.89 \times 10^{-5}}{0.150}$ $= 6.59 \times 10^{-4}\ mol L^{-1}$ or $6.59 \times 10^{-4}\ M$

The symbol *M* was used earlier to represent molar mass. Take care not to confuse the two different uses; *M* is a quantity symbol, M is a unit symbol.

➤ Try yourself 15.5.4

CALCULATING MOLARITY GIVEN THE MASS OF SOLUTE

Calculate the concentration, in mol L^{-1}, of a solution that contains 4000 mg of ethanoic acid (CH_3COOH) dissolved in 100 mL of solution.

Calculating the number of moles of solute in a solution

The formula used to calculate the molarity of a solution is $c = \frac{n}{V}$. If you rearrange the formula, you can use it to calculate the number of moles of solute in a solution of given concentration and volume:

$$n = c \times V$$

where n is the amount (mol), c is the concentration (mol L^{-1}) and V is the volume (L).

Worked example 15.5.5

CALCULATING THE NUMBER OF MOLES OF SOLUTE IN A SOLUTION

Calculate the number of moles of ammonia (NH_3) in 25.0 mL of a 0.3277 M ammonia solution.	
Thinking	**Working**
Convert the given volume to litres.	$V(NH_3) = \frac{25.0}{1000}$ $= 0.0250\,L$
Calculate the number of moles of the compound, using the formula: $n = c \times V$	$n(NH_3) = c \times V$ $= 0.3277 \times 0.0250$ $= 8.19 \times 10^{-3}\,mol$

➤ Try yourself 15.5.5

CALCULATING THE NUMBER OF MOLES OF SOLUTE IN A SOLUTION

Calculate the number of moles of potassium permanganate ($KMnO_4$) in 100 mL of a 0.0250 M solution of the compound.

DILUTION

Many commercially available domestic and industrial products come in the form of concentrated solutions. Examples are pesticides (Figure 15.5.5), fertilisers, detergents, fruit juices, acids and other chemicals. A major reason for using concentrates is to save on transportation costs. Diluted solutions contain a lot of water and that extra mass has to be transported, which increases costs. It is also more convenient to buy concentrated products and dilute with water at home or in the workplace.

FIGURE 15.5.5 Pesticide solutions used in aerial spraying are prepared by diluting a concentrated solution of the pesticide compound.

Everyday examples of dilution are:

- adding water to cordial
- a laboratory technician making a 1 M solution of hydrochloric acid from a bottle of concentrated hydrochloric acid
- a home gardener diluting fertiliser concentrate to spray on the lawn
- a farmer diluting weedkiller concentrate to spray on a wheat crop
- an assistant in a commercial kitchen diluting concentrated detergent solution before using it to wash dishes.

In this module, you will learn how to calculate the concentrations of solutions before and after they have been diluted.

Calculating concentrations when solutions are diluted

The process of adding more solvent to a solution is known as **dilution**. When a solution is diluted, its concentration is decreased.

WS 2.2.3

For example, if 50 mL of water is added to 50 mL of 0.10 mol L^{-1} sugar solution, the amount of sugar remains unchanged but the volume of the solution in which it is dissolved doubles. As Figure 15.5.6 shows, this means the sugar molecules are spread further apart during the dilution process, and so the concentration of the sugar solution is decreased (it will become 0.050 mol L^{-1} in this instance).

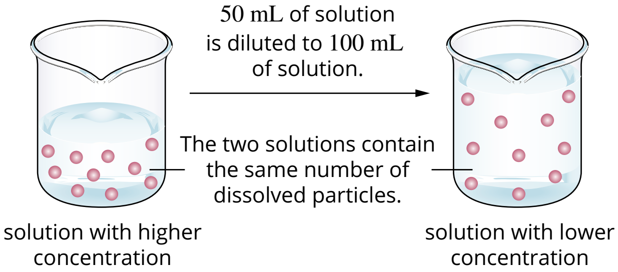


FIGURE 15.5.6 Dilution does not change the number of solute particles but the concentration of the solute decreases.

It is important to recognise that diluting a solution (by adding more solvent) does not change the number of moles of the solute present.

Suppose you had V_1 litres of a solution and the concentration was c_1 mol L^{-1}. The number of moles of the solute is given by:

$$n_1 = c_1 \times V_1$$

Suppose water was added to make a new volume, V_2, which changes the concentration to c_2. The number of moles of the solute, n_2, in this diluted solution is given by:

$$n_2 = c_2 \times V_2$$

Since the number of moles of solute has not changed, $n_1 = n_2$; therefore:

$$c_1V_1 = c_2V_2$$

This formula is useful when solving problems involving diluted solutions. (Note that this formula can also be used with different concentration and volume units as long as they are the same on both sides.)

Worked example 15.5.6

QUESTIONS INVOLVING DILUTION

Calculate the concentration of the solution formed when 10.0 mL of water is added to 5.00 mL of 1.2 M HCl.	
Thinking	**Working**
Write down the value of c_1 and V_1. Note: c_1 and V_1 refer to the original solution, before water was added.	$c_1 = 1.2$ M $V_1 = 5.00$ mL
Write down the value of V_2. Note: V_2 is the total volume of the original solution plus the added water.	$V_2 = 10.0 + 5.00$ $= 15.0$ mL
Transpose the equation $c_1V_1 = c_2V_2$ to allow calculation of the concentration, c_2, of the new solution.	$c_1V_1 = c_2V_2$ $c_2 = \frac{c_1V_1}{V_2}$
Calculate the concentration of the diluted solution.	$c_2 = \frac{1.2 \times 5.00}{15.0}$ $= 0.40$ M

➤ Try yourself 15.5.6

QUESTIONS INVOLVING DILUTION

Calculate the concentration of the solution formed when 95.0 mL of water is added to 5.00 mL of 0.500 M HCl.

Changing the units of concentration

At times it is useful to change, or 'convert', concentration values from one unit to another unit. One way to do this is in a series of simple steps. Worked example 15.5.7 shows how this can be done.

Worked example 15.5.7

CONCENTRATION UNIT CONVERSIONS

Determine the concentration, in ppm, of a 0.00200 M solution of NaCl. Remember that concentration in ppm is the same as $mg\,L^{-1}$.	
Thinking	**Working**
Calculate the number of moles of solute in 1.00 L of the solution.	$n(\text{NaCl}) = c \times V$ $= 0.00200 \times 1.00$ $= 0.00200\,\text{mol}$
Calculate the mass, in grams, of solute in 1.00 L of the solution.	$M(\text{NaCl}) = 35.45 + 22.99$ $= 58.44\,\text{g mol}^{-1}$ $m(\text{NaCl}) = n \times M$ $= 0.00200 \times 58.44$ $= 0.117\,\text{g}$
Calculate the mass, in mg, of solute in 1.00 L of the solution.	$m(\text{NaCl}) = 0.117 \times 1000$ $= 117\,\text{mg}$
Express the concentration of the solute in ppm.	$c(\text{NaCl}) = 117\,\text{ppm}$

➤ Try yourself 15.5.7

CONCENTRATION UNIT CONVERSIONS

Calculate the concentration, in ppm, of a 0.0100 M solution of NaOH. Remember that concentration in ppm is the same as $mg\,L^{-1}$.

15.5 Review

SUMMARY

- The concentration of a solution is defined as the quantity of solute dissolved in a quantity of solution.
- Concentration units have two parts. The first part provides information about the quantity of solute. The second part provides information about the quantity of solution.
- The concentrations of solutions can be expressed in different units including $g\,L^{-1}$, $mg\,L^{-1}$, %(m/m), %(m/v), %(v/v), ppm and ppb.
- Molarity is defined as the number of moles of solute per litre of solution.
- Molarity is calculated using the formula $c = \frac{n}{V}$, where c is the concentration in $mol\,L^{-1}$, n is the number of moles and V is the volume in litres.
- The formula $n = c \times V$ can be used to calculate the number of moles of solute in a given volume of solution.
- Numerical problems involving dilution can be solved using the formula:

$$c_1V_1 = c_2V_2$$

where c is the concentration and V is the volume. When using this formula, c_1 and c_2 must be in the same units of concentration and V_1 and V_2 must be in the same units of volume.
- One type of concentration unit can be converted to another unit in simple steps. For example, units of %(m/m) can be converted to units of $mol\,L^{-1}$.

KEY QUESTIONS

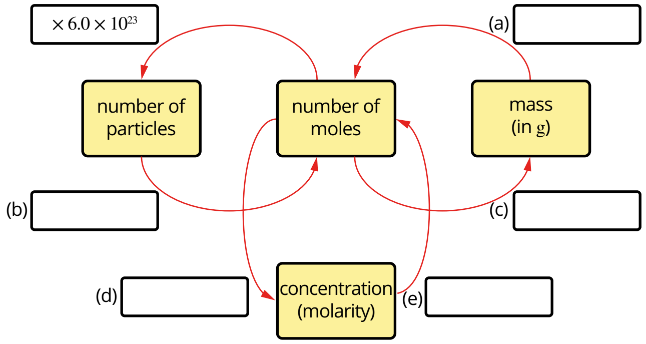
Retrieval

1 The nutrition panel on a flavoured milk drink states that it contains 35.0 g of sugar and 7.5 g of fat in every 250 mL serving. Determine the concentrations of sugar and fat expressed as %(m/v).

2 Describe what measures of concentration are used for household products like cleaning agents and alcohol.

Comprehension

3 Use a flow chart to show how to convert between the quantities listed in in the figure below.



4 Explain why ppm is a more commonly used measure for the concentration of mercury in fish for human consumption rather than $mol\,L^{-1}$ or %m/m.

Analysis

5 Calculate the final concentration of each of the following diluted solutions.

- **a** A 10.0 mL sample of water is added to 5.0 mL of 1.2 M HCl.
- **b** A 1.0 L sample of water is added to 3.0 L of 0.10 M HCl.
- **c** A 5.0 mL sample of 0.50 M HCl is added to 95.0 mL of water.

6 Calculate the concentration of the following solutions in ppm.

- **a** A 25 mg sample of $CaCl_2$ dissolved in 5.0 kg of solution.
- **b** A 1.25 g sample of lead nitrate dissolved in 2000 g of solution.
- **c** A 4.0×10^{-3} g sample of $MgSO_4$ dissolved in 150 g of solution.

7 Calculate which of the following is the most concentrated solution: 5.00 g of NaCl in 100 mL of water or 2.00 M NaCl.

8 Calculate the concentration of a 200 mL solution that contains 4.56 g of barium sulfate.

9 Ciguatoxin is a dangerous toxin that is found in some fish in Australia's tropical waters. The amount of this chemical that can kill a 70 kg human is approximately 15 mg. Calculate the lethal concentration of the toxin in a human body in parts per billion (ppb).

10 The concentration of a solution of ammonia (NH_3) is 1.5%(m/v). Determine the molar concentration of a solution produced by diluting 25.0 mL of this solution with 250 mL of water.

11 The concentration of a solution of NaOH is 17.0%(m/v). Calculate the concentration of this solution in $mol\,L^{-1}$.

12 Referring to Figure 15.5.3 on page 487, determine the concentration of alcohol. (Assume ethanol CH_3COOH is in $mol\,L^{-1}$; density of ethanol is $0.7893\ g\,cm^{-3}$.)

MANDATORY PRACTICAL 5

Precipitation reactions

Research and planning

Aim

To use a precipitation table to identify the positive and negative ions present in a series of unknown samples

Rationale (scientific background to the experiment)

An important aspect of chemistry involves qualitative analysis—the process of finding out what substances are present in an unknown sample. It is known as 'qualitative' because it is based on what substances are present, rather than on quantities that can be measured. This experiment uses the formation of a precipitate to identify unknown cations and anions.

For example, silver nitrate forms a white precipitate when added to any solution containing chloride ions, and therefore it is commonly used as a test for the presence of chloride ions.

Precipitation reactions are those that produce a solid when two solutions are mixed. The original reactants are soluble, but one of the products is insoluble, and precipitates out of solution. A precipitate generally forms very quickly, and may, in some cases, consist of fine particles. This gives the appearance of cloudy water or thick paint, depending on the nature and amount of precipitate formed. Some precipitates will slowly settle to the bottom of the test-tube.

Timing

80 minutes

Materials

- dropping bottles containing the following test solutions:

Cations	Anions
Ba^{2+}	SO_4^{2-}
Al^{3+}	Br^-
Fe^{3+}	NO_3^-
Fe^{2+}	NH_3
Cu^{2+}	OH^-
Ca^{2+}	CO_3^{2-}
Pb^{2+}	I^-
Ag^+	Cl^-
Zn^{2+}	

- solutions of five unknown compounds, A to E
- semi-micro test-tubes
- semi-micro test-tube rack
- safety glasses, lab coats

Risk assessment

As directed by your teacher, complete the risk assessment and management table by referring to the hazard labels on the reagent bottles or safety data sheets (SDS) or your teacher's risk assessment for the activity.

PRE-LAB SAFETY INFORMATION		
Material used	**Hazard**	**Control**
dissolved compounds of lead and barium	highly toxic by all routes of exposure; lung irritation	Wear eye and skin protection.
silver nitrate	can burn and stain skin and clothes	Wear eye and skin protection.
compounds containing Fe^{3+}, Cu^{2+} and I^-	can cause skin and eye irritation	Avoid contact, and wear eye and skin protection.
calcium hydroxide	slightly toxic if ingested	Wear eye and skin protection.
Please indicate that you have understood the information in the safety table. Name (print): ______________________ I understand the safety information (signature): __________		

Method

Risk assessment

Consideration of risks includes chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and Assessment book or download it from your eBook.

Each of the unknown solutions A to E contains one dissolved ionic compound.

Table 15.6.1 gives a brief description of precipitates formed when various cations and anions are allowed to react. This information will assist you in identifying the cation and anion present in the ionic compound in solution. In some cases the distinctive colour of a precipitate formed will be sufficient to identify one of the ions in solution. In other cases, you may need a few tests to isolate the ion.

TABLE 15.6.1 Formation of precipitates

- White, green, yellow and so on means the precipitate is visible on the bottom of the test-tube.
- Cloudy means the precipitate is uniformly suspended in the water.
- For blank spaces, the substance is assumed to be soluble, giving a clear (possibly coloured) solution.

cations used to identify the anion ↓	**Anions used to identify the cation**							
	nitrate NO_3^-	sulfate SO_4^{2-}	carbonate CO_3^{2-}	chloride Cl^-	iodide I^-	bromide Br^-	hydroxide OH^-	hydroxide (OH^-) with ammonia (NH_3) added
sodium Na^+								
potassium K^+								
barium Ba^{2+}		white	white				cloudy	
calcium Ca^{2+}			cloudy white				cloudy	
aluminium Al^{3+}			cloudy				cloudy (light)	cloudy
zinc Zn^{2+}			cloudy				cloudy	
iron(II) Fe^{2+}			grey-green				green	light green
iron(III) Fe^{3+}							brown	brown
copper Cu^{2+}			light blue		mustard		blue	
lead Pb^{2+}		white	white	cloudy	yellow	cloudy	white	white
silver Ag^+			light brown	white	light yellow	cream	brown	

Perform these steps to identify the anion present in each of the dissolved unknown ionic compounds.

1. Place about 20 drops of a selected unknown solution in a semi-micro test-tube.
2. Carefully add a few drops, drop-wise, of one of the test cation solutions to the test-tube.
3. In Tables 15.6.2a–e for the cation test, record:
   - **a** your observations and comment on the colour, lightness/darkness and density of any precipitate formed
   - **b** possible anions that would give results consistent with your test.
4. Choose a second test cation that will help you decide which of the possible anions identified in step 3 is present in your unknown solution. Test a fresh sample of unknown solution with this second cation and record your results in Tables 15.6.2a–e for the cation tests.
5. If necessary, repeat step 4 until you have clearly identified your unknown anion.
6. To identify the cation present in each of the unknown dissolved ionic compounds, repeat steps 1–5 using test anion solutions and putting your results in Tables 15.6.3a to 15.6.3e for the anion tests.
7. Deduce the formula of the ionic compound dissolved in each of the unknown solutions.

Variables

i Independent: the anion or cation tested

ii Dependent: the colour change of solution and presence/absence of a precipitate

iii Controlled: the unknown substance tested, concentration of all substances

Analysing

Raw data

1 Record your data for each of the unknown cation solutions A–E in Tables 15.6.2a to 15.6.2e.

TABLE 15.6.2a Cation test for unknown solution A

	Test 1	Test 2	Test 3	Test 4
Cation tested				
Observations				
Possible anions				

Anion for unknown solution A _______________

TABLE 15.6.2b Cation test for unknown solution B

	Test 1	Test 2	Test 3	Test 4
Cation tested				
Observations				
Possible anions				

Anion for unknown solution B _______________

TABLE 15.6.2c Cation test for unknown solution C

	Test 1	Test 2	Test 3	Test 4
Cation tested				
Observations				
Possible anions				

Anion for unknown solution C _______________

TABLE 15.6.2d Cation test for unknown solution D

	Test 1	Test 2	Test 3	Test 4
Cation tested				
Observations				
Possible anions				

Anion for unknown solution D _______________

TABLE 15.6.2e Cation test for unknown solution E

	Test 1	Test 2	Test 3	Test 4
Cation tested				
Observations				
Possible anions				

Anion for unknown solution E _______________

2 Record your data for each the unknown anion solutions A–E in Tables 15.6.3a to 15.6.3e.

TABLE 15.6.3a Anion test for unknown solution A

	Test 1	Test 2	Test 3	Test 4
Anion tested				
Observations				
Possible cations				

Cation for unknown solution A _______________

TABLE 15.6.3b Anion test for unknown solution B

	Test 1	Test 2	Test 3	Test 4
Anion tested				
Observations				
Possible cations				

Cation for unknown solution B ________________

TABLE 15.6.3c Anion test for unknown solution C

	Test 1	Test 2	Test 3	Test 4
Anion tested				
Observations				
Possible cations				

Cation for unknown C ________________

TABLE 15.6.3d Anion test for unknown solution D

	Test 1	Test 2	Test 3	Test 4
Anion tested				
Observations				
Possible cations				

Cation for unknown solution D ________________

TABLE 15.6.3e Anion test for unknown solution E

	Test 1	Test 2	Test 3	Test 4
Anion tested				
Observations				
Possible cations				

Cation for unknown solution E ________________

Processed data

3 Identify each of the unknown solutions and write its formula.

4 Write an ionic equation for each reaction that led to the formation of a precipitate.

Reflect and check that your data analysis demonstrates these characteristics.

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.
☐ Accurate application of algorithms, visual and graphical representations of data is demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

Analysis

5 A solution was thought to contain either copper ions or lead ions. Describe how you could identify which ions were present in the solution.

6 A solution was tested to determine the cation that was present, and the following results were obtained.

- The addition of dilute sulfuric acid (H_2SO_4) produced a white precipitate.
- The addition of dilute sodium hydroxide (NaOH) produced a white precipitate.
- The addition of potassium iodide produced a yellow precipitate.

Deduce the cation that was present.

7 A solution was thought to contain either Fe^{2+} or Fe^{3+} ions. Describe the minimum number of tests that could be conducted to identify the solution.

Reflect and check that your analysis demonstrates these characteristics.

☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.
☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.
☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Interpreting and communicating

Conclusion

1 **a** State your conclusion.

b What evidence have you collected that leads you to this conclusion?

Evaluation

2 Considering your analysis and conclusion, did the experiment provide an effective and efficient method of determining the positive and negative ions present in a series of unknown samples?

Improvements

3 If you were to repeat the experiment, identify the steps that you would do differently. You should include these points in your answer.

- **a** how you would change the methodology and how this might improve the results
- **b** how you performed the tasks and the skills that you need to improve on in your technique
- **c** how the collection of data (identification of correct colours) could be improved

Reflect and check that your evaluation demonstrates these characteristics.
☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis). ☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respect to the error analysis and limitations or sufficiency of the data.
☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment which are logically derived from the analysis of the evidence.

Chapter review

WS 2.2.7

KEY TERMS

aqueous
concentrated solution
crystallisation
crystallise
dilute solution
dilution
dipole
dissociation
dissolution
electronegativity
electrostatic attraction
full equation
homogeneous
hydrated
hydrides
hydrogen bond
hydrogen bonding
hydronium ion
intermolecular bond
intermolecular force
intramolecular bond
ion–dipole attraction
ion–dipole interaction
ionic equation
ionise
limestone
lone pair
miscible
molarity
non-polar
parts per billion (ppb)
parts per million (ppm)
polar
precipitate
precipitation reaction
saturated solution
seed crystal
solubility
solubility curve
solubility table
solute
solution
solvent
spectator ion
supersaturated solution
surface tension
unsaturated solution

KEY QUESTIONS

Retrieval

1 Hydrogen peroxide solutions for hair bleaching are sold as 120 mL solutions containing 5.00 g of H_2O_2 dissolved in water. Determine the concentration of hydrogen peroxide in the bleaching product.

A 0.81 %m/v
B 1.23 M
C 41.7 $g\,L^{-1}$
D $1.23 \times 10^{-3}\,mol\,L^{-1}$

2 Identify which of the following intermolecular forces gives water its unique physical properties.

A dispersion forces
B covalent bonds
C hydrogen bonds
D dipole–dipole interactions

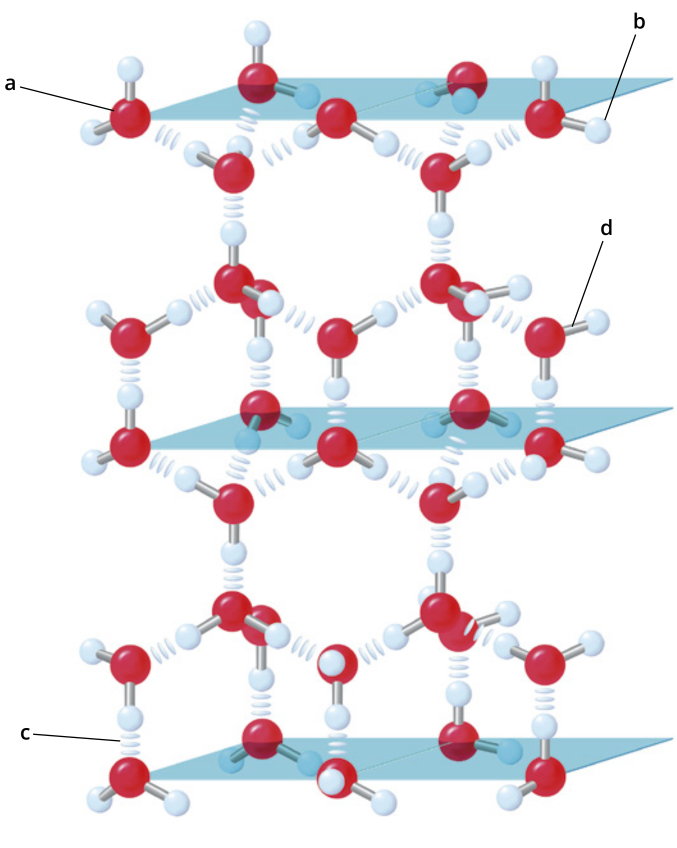
3 Determine which of the following would be insoluble in water.

A sodium nitrate
B ammonium chloride
C barium sulfate
D iron chloride

4 Identify which of the following would dissolve in water by forming hydrogen bonds.

A ammonia (NH_3)
B hexane (C_6H_{14})
C sodium chloride (NaCl)
D hydrogen chloride (HCl)

5 The image below shows the arrangement of water molecules in ice. Identify the parts of the diagram labelled **a**–**d**.



6 Many of the solvent properties of water depend on its structure and bonding. List all the bonds present in a sample of water from weakest to strongest. Remember to include all intramolecular and intermolecular bonding.

7 Describe what happens to the forces between solute and solvent substances when an ionic substance such as potassium bromide dissolves in water.

8 Explain why water is such a good solvent for polar and ionic substances.

9 Calculate the molar concentration of a 1.55 L solution that contains 1.223 g of $CaCl_2$.

Comprehension

10 Consider the following ions: SO_4^{2-}, PO_4^{3-}, Br^-, S^{2-}. Determine which of these would combine with Fe^{3+}(aq) to form a precipitate.

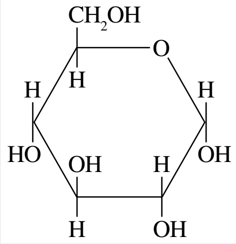
11 Write full balanced chemical equations and ionic equations for each of the following precipitation reactions.

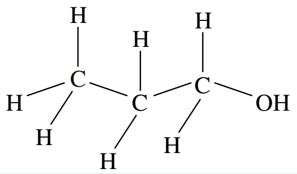
- **a** $NH_4Cl(aq) + AgNO_3(aq) \rightarrow$
- **b** $FeCl_2(aq) + Na_2S(aq) \rightarrow$
- **c** $Fe(NO_3)_3(aq) + KOH(aq) \rightarrow$
- **d** $CuSO_4(aq) + NaOH(aq) \rightarrow$
- **e** $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow$

12 Explain why water exhibits a much higher boiling point than the other group 16 hydrides.

13 Classify the following substances according to the way they dissolve in water: $C_6H_{12}O_6$, HI, I_2, C_2H_4, C_3H_7OH, HNO_3, CH_4.

The structures of $C_6H_{12}O_6$ and C_3H_7OH are shown below to assist you.





14 Explain, with reference to its structure and bonding, why octane does not dissolve in water.

15 Sugar is soluble in water. Explain the three steps involved when sugar molecules dissolve to form an aqueous sugar solution.

16 Water is added to 50.0 mL of 5.00 M HCl until the total volume is 250 mL. Calculate the concentration of HCl in the diluted solution.

17 A 1.50 mL sample of a 0.0500 M solution of calcium chloride ($CaCl_2$) was diluted with water to a volume of 10.0 L. Determine the concentration of Cl^- ions in the diluted solution in ppm.

18 In your own words, discuss the significance of polarity and hydrogen bonding in relation to the high boiling point of water.

Analysis

19 Calculate the concentration, in $g\,L^{-1}$, of a 60 mL solution that contains 5.0 g of sugar.

- **A** 300
- **B** 053
- **C** 12
- **D** 83

20 Nitrogen gas (N_2) and ethene (C_2H_6) are both insoluble in water, whereas ethanol (C_2H_5OH) is soluble. Refer to the bonding in each of the molecules to explain these differences.

21 DDT is a hazardous agricultural insecticide, which has been banned in many countries. It is only slightly soluble in water but is very soluble in fats and oils, so accumulates in the fat deposits of animals. Use the skills you have gained from this chapter to deduce the polarity of the DDT molecule from its solubility characteristics.

22 Use Figure 15.3.2 on page 468 to deduce the minimum temperature required to dissolve 40 g of $AgNO_3$ in 25 g of water.

23 Refer to Figure 15.3.2 on page 468 to answer the following.

A saturated solution of sodium nitrate was prepared by dissolving the maximum amount of solute in water at 50°C.

- **a** Determine what mass will dissolve in 150 g of water.
- **b** The solution was cooled to 30°C and some of the solute crystallised. Calculate the mass of crystals that would be formed.

24 Analyse what happens to the forces between solute and solvent substances when an ionic substance such as potassium bromide dissolves in water.

25 Determine the molar concentration of a saturated $CuSO_4 \cdot 5H_2O$ solution at 20°C. (Assume that 1 g of water gives 1 mL of solution.)

26 Compare and contrast the dissolution of ionic substances and molecular substances that ionise in water.

27 Identify the ions produced when the following compounds are added to water.

- **a** $Cu(NO_3)_2$
- **b** $ZnSO_4$
- **c** $(NH_4)_3PO_4$

28 Water boils at 100°C. However, a much higher temperature (1000°C) is needed to decompose water molecules into hydrogen gas and oxygen gas.

- **a** Using water as an example, explain the meaning of the terms ‘intermolecular’ and ‘intramolecular’ forces.
- **b** Deduce which of the two types of forces in part **a** is stronger. Justify your answer by using the information at the beginning of the question.

29 Calculate the molar concentration of a sodium sulfate solution that has a concentration of 4.26%(m/v).

30 Identify three sulfate compounds that have these characteristics.

- **a** soluble in water
- **b** insoluble in water

31 Calculate the volume of 10 M HCl needed to prepare 200 mL of 0.45 M HCl solution.

32 Categorise the following as soluble or insoluble.

- **a** copper chloride
- **b** strontium sulfate
- **c** barium hydroxide
- **d** lead sulfate
- **e** potassium nitrate

33 In each 5.0 mL dose of a cold and flu medicine, there is 2.00 mg of the active ingredient. Calculate the concentration of this chemical, in mg L^{-1} and %(m/v).

34 Explain the difference between a precipitate and a crystallised solute.

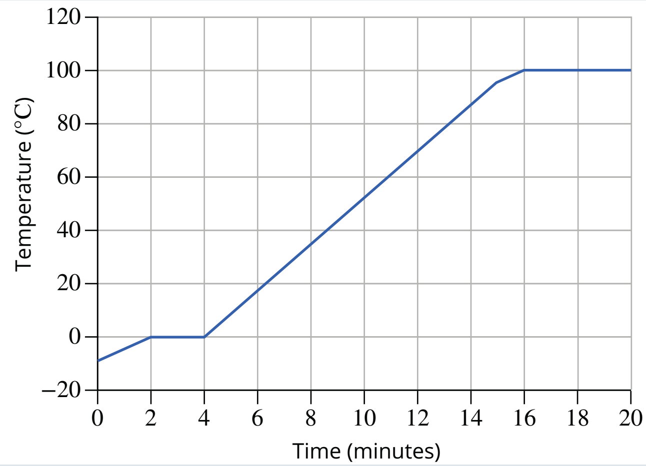
35 Calculate the final volume of a solution of NaOH if 200 mL of a 0.500 M solution was diluted to 0.200 M.

36 Calculate the concentration of a 0.00400 M solution of potassium nitrate (KNO_3) in ppm.

Knowledge utilisation

37 An evaporative air conditioner is used to cool the air inside some buildings. The cooler lowers the air temperature by using heat from the air to evaporate water. Select the property that is most important when water is used in an evaporative air conditioner.

38 A student was asked to record the temperature changes as a sample of ice was heated. The ice was placed in a beaker and heated with a Bunsen burner for 20 minutes. The graph below shows the temperature, in degrees Celsius, recorded at 1 minute intervals.

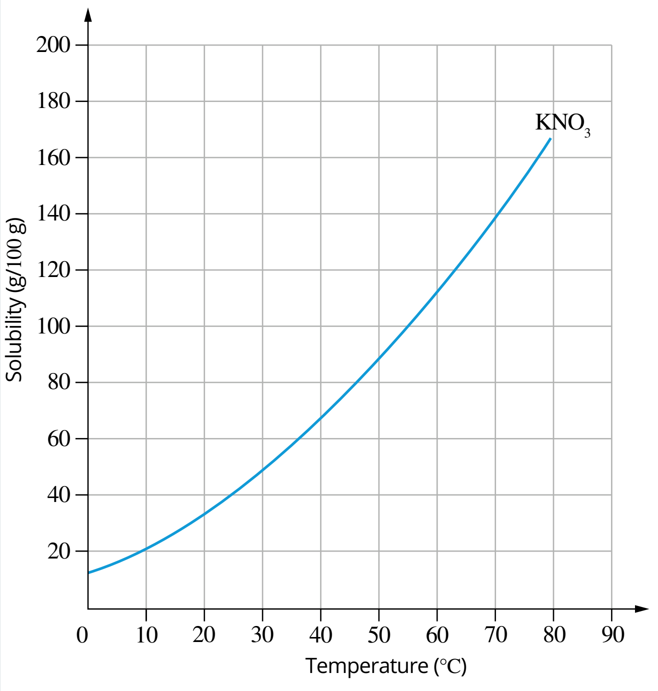


- **a** Explain what is happening, at a molecular level, between 2.0 and 4.0 minutes on the graph.
- **b** Even though heating is continued for 20 minutes, no further temperature rise is observed after 16 minutes. Discuss what happens to the added heat between 16 and 20 minutes.

39 A student conducts an experiment to investigate the solubility of CH_4 and CH_3OH molecules in three different solvents, X, Y and Z. The student finds that the CH_3OH molecules dissolve well in solvent Y but do not dissolve in solvent X. The student also finds that CH_4 partially dissolves in solvent Z and completely dissolves in solvent X. List each of the solvents in order from most polar to least polar. Explain your reasoning.

40 Glucose ($C_6H_{12}O_6$) and common salt (NaCl) both form clear colourless solutions when dissolved in water. Suggest an experiment you could perform in the laboratory to tell the two solutions apart.

41 The following question refers to the solubility curve of potassium nitrate (KNO_3) shown in the figure below.



- **a** Determine the mass of potassium nitrate that will dissolve in 500 g of water at 70°C.
- **b** If 50 g of water at 70°C contained 25 g of KNO_3, determine how much more KNO_3 must be added to make a saturated solution.
- **c** If 100 g of KNO_3 is added to 50 g of water at 40°C, calculate the mass of KNO_3 that will not dissolve.
- **d** Calculate the molar concentration of the potassium nitrate solution prepared in part **c**. (Assume that 1 g of water gives 1 mL of solution.)

Questions **42** and **43** refer to the table below.

42 A solution was tested to determine the cation present, and the results in the table below were obtained.

- The addition of dilute sodium carbonate (Na_2CO_3) gave a cloudy solution.
- The addition of potassium hydroxide (KOH) with ammonia (NH_3) added gave a cloudy solution.
- The addition of sodium hydroxide (NaOH) gave a cloudy solution.

Deduce which cation was present.

43 A solution was thought to contain Cu^{2+}, Pb^{2+} or Ag^+ ions. Give two anions that would not be helpful in determining what the cation present was.

Formation of precipitates

- White, green, yellow and so on means the precipitate is visible on the bottom of the test-tube.
- Cloudy means the precipitate is uniformly suspended in the water.
- For blank spaces, the substance is assumed to be soluble, giving a clear (possibly coloured) solution.

Colours produced in precipitate reactions

Cations used to identify the anion	**Anions used to identify the cation**							
	nitrate NO_3^-	sulfate SO_4^{2-}	carbonate CO_3^{2-}	chloride Cl^-	iodide I^-	bromide Br^-	hydroxide OH^-	hydroxide (OH^-) with ammonia (NH_3) added
silver Ag^+			light brown	white	light yellow	cream	brown	
copper Cu^{2+}			light blue		mustard		blue	
lead Pb^{2+}		white	white	cloudy	yellow	cloudy	white	white
calcium Ca^{2+}			cloudy white				cloudy	
aluminium Al^{3+}			cloudy				cloudy (light)	cloudy
zinc Zn^{2+}			cloudy				cloudy	
iron(II) Fe^{2+}			grey-green				green	light green

CHAPTER

16

An introduction to acids and bases

Acids and bases make up some of the household products in your kitchen and laundry. In this chapter, you will be introduced to a theory that explains the chemical properties of acids and bases, helping you to explain their usefulness within the home and industry. You will also look at how the acidity of a solution can be measured, therefore defining the solution as a strong or weak acid. Lastly, the reactions of acids and bases will be investigated and enable you to predict the products of a variety of reactions of acids and bases.

WS 2.3.1

Syllabus subject matter

Topic 2 • Aqueous solutions and acidity

pH

- recall that pH is dependent on the concentration of hydrogen ions in solution
- use the pH scale to compare the levels of acidity or alkalinity of aqueous solutions
- use the Arrhenius model to explain the behaviour of strong and weak acids and bases in aqueous solutions.

REACTION OF ACIDS

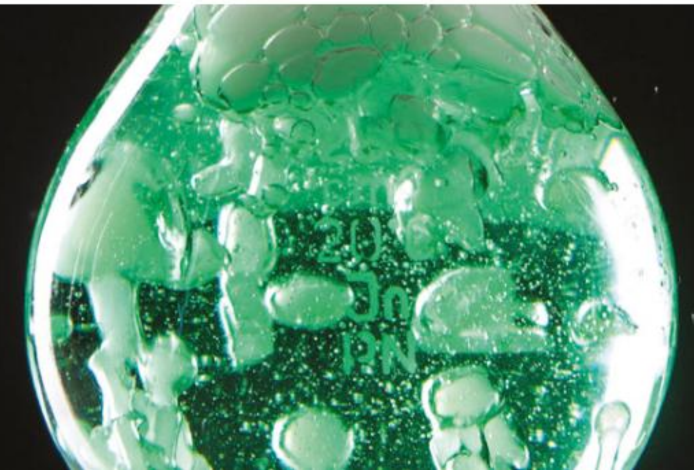
- understand and apply the reactions of acids with bases, metals and carbonates to determine reactants and products
- construct and use appropriate representations, including ionic formulas, chemical formulas and chemical equations, to symbolise, the reactions of acids and bases; and ionic equations to represent the reacting species and products in these reactions.

SCIENCE AS A HUMAN ENDEAVOUR

- **Acid rain:** Most sulfur dioxide released to the atmosphere comes from burning coal or oil in electric power stations.
- **Water quality:** Knowledge of the composition of water from different sources informs decisions about how that water is treated and used.

MANDATORY PRACTICAL

- Investigate the properties of strong and weak acids.



16.1 Introducing acids and bases

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain the properties of acids and bases using the Arrhenius model.

Acids are used in your home, in agriculture and in industry (Figure 16.1.1). They also have an important role in your body. In this module, you will be introduced to the Arrhenius model that explains the chemical properties of acids and bases, helping you to explain their usefulness within the home and industry. You will also look at how the **acidity** of a solution can be measured, therefore defining the solution as a strong or weak acid.

FIGURE 16.1.1 Common household products that contain acids, bases or salts

Table 16.1.1 gives the names, chemical formulas and uses of some common acids.

TABLE 16.1.1 Common acids and their everyday uses

Name	Formula	Uses
hydrochloric acid	HCl	• present in stomach acid to help break down proteins • as a cleaning agent for brickwork
sulfuric acid	H_2SO_4	• one of the most common chemicals manufactured • in car batteries and in the manufacture of fertilisers and detergents
nitric acid	HNO_3	• in the manufacture of fertilisers, dyes and explosives
ethanoic acid (acetic acid)	CH_3COOH	• found in vinegar • as a preservative
carbonic acid	H_2CO_3	• found in carbonated soft drinks and beer
phosphoric acid	H_3PO_4	• in some soft drinks and in the manufacture of fertilisers
citric acid	$C_6H_8O_7$	• found in citrus fruits
ascorbic acid	$C_6H_8O_6$	• found in citrus fruits (vitamin C)

Many cleaning agents used in the home, such as washing powders and oven cleaners, contain **bases**. **Solutions** of ammonia are used as floor cleaners, and sodium hydroxide is the major active ingredient in oven cleaner. Bases are effective cleaners because they react with fats or oils to produce water-soluble soaps. A soluble base is referred to as an **alkali**.

Table 16.1.2 gives the names, chemical formulas and uses of some common bases.

Aqueous solutions of bases are called alkalis. Calcium carbonate will react with acids. However, it is not considered an alkali because it is insoluble in water.

TABLE 16.1.2 Common bases and their uses

Name	Formula	Uses
sodium hydroxide (caustic soda)	NaOH	• drain and oven cleaners, and soap making
ammonia	NH_3	• household cleaners, fertilisers and explosives
calcium hydroxide	$Ca(OH)_2$	• found in cement and mortar • in garden lime to adjust soil pH
magnesium hydroxide	$Mg(OH)_2$	• key ingredient in some antacids, such as milk of magnesia, to overcome indigestion
sodium carbonate	Na_2CO_3	• in the manufacture of washing powder and glass

All acids have some properties in common. Bases also have common properties. The properties of acids and bases are summarised in Table 16.1.3.

TABLE 16.1.3 Properties of acids and bases

Properties of acids	Properties of bases
turn litmus indicator red	turn litmus indicator blue
tend to be corrosive	are caustic and feel slippery
taste sour	taste bitter
react with bases	react with acids
solutions have a relatively low pH	solutions have a relatively high pH
solutions conduct an electric current	solutions conduct an electric current

CHANGING IDEAS ABOUT THE NATURE OF ACIDS AND BASES

Over the years, there have been many attempts to define acids and bases. At first, acids and bases were defined in terms of their observed properties such as their taste, effect on **indicators** and reactions with other substances.

For example, in the seventeenth century, British scientist Robert Boyle described the properties of acids in terms of taste, their action as solvents and how they changed the colour of certain vegetable extracts (similar to red cabbage juice seen in Figure 16.1.2). He also noticed that alkalis (soluble bases) could reverse the effect that acids had on these extracts.

FIGURE 16.1.2 Natural acid–base indicators are found in plants such as red cabbage. Red cabbage extract turns a different colour in (from left to right) strong acid, weak acid, neutral solution, weak base and strong base.

It was not until the late eighteenth century that attempts were made to define acids and bases on the basis of the nature of their constituent elements. Antoine Lavoisier, a French chemist, thought that acidic properties were due to the presence of oxygen. While this explanation applied to sulfuric acid (H_2SO_4), nitric acid (HNO_3) and phosphoric acid (H_3PO_4), it did not explain why hydrochloric acid (HCl) was an acid.

In about 1810, Humphry Davy suggested that the acidic properties of substances were associated with hydrogen and not oxygen. He came to this conclusion after producing hydrogen gas by reacting acids with metals. Davy also suggested that acids react with bases to form salts and water.

This theory was further developed by the Swedish scientist Svante Arrhenius. In 1887, he defined acids and bases as follows.

- Acids are substances that **dissociate** (break apart) and **ionise** (form ions) in water. They produce hydrogen ions (H^+).
- Bases dissociate in water to produce **hydroxide ions** (OH^-).

THE ARRHENIUS MODEL OF ACIDS

In the Arrhenius model, an acid is defined as a substance that is ionised in water to produce hydrogen ions. For example, when hydrogen chloride gas is bubbled through water, the HCl molecules break apart and are ionised to form hydrochloric acid, which consists of hydrogen ions and chloride ions in the aqueous solution:

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

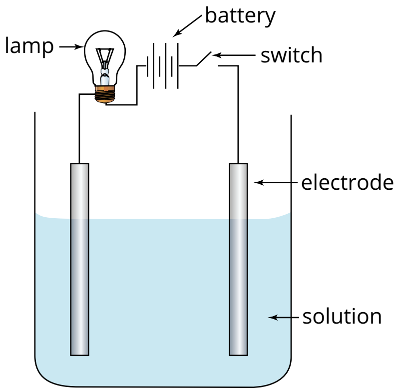
The Arrhenius model is useful because it explains the similarity in the reactions of different acids. The hydrogen ions that are present in acid solutions account for the common properties of acids. For example, as you will see later in this chapter, the hydrogen ions present in all acid solutions form hydrogen gas when the acid reacts with metals such as magnesium:

$$2H^+(aq) + Mg(s) \rightarrow H_2(g) + Mg^{2+}(aq)$$

ACIDS AND BASES AS ELECTROLYTES

Hydrochloric acid is formed when hydrogen chloride ionises in water to form a solution containing hydrogen ions and chloride ions. When a solution of hydrochloric acid is tested with the conductivity apparatus shown in Figure 16.1.3, the light globe will glow. The chloride ions are attracted to the positive electrode of the apparatus and the hydrogen ions to the negative electrode. The flow of current in the solution is a result of the movement of these ions.

Because sodium hydroxide dissociates into sodium and hydroxide ions in aqueous solution, a sodium hydroxide solution will also conduct an electric current.



If the solution conducts an electric current, the lamp glows.

FIGURE 16.1.3 Apparatus used to test the conductivity of acid solutions

THE ARRHENIUS MODEL OF BASES

Arrhenius defined a base as a substance that dissociates in water to form hydroxide ions. Sodium hydroxide is an ionic compound, containing sodium ions and hydroxide ions. When it is added to water it dissociates:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

Some bases dissociate to form more than one hydroxide ion:

$$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$$

The presence of the hydroxide ions in solutions of bases accounts for the common properties of bases. For example, the hydroxide ion reacts with the litmus indicator to form a blue-coloured compound.

According to the Arrhenius model, the stronger an acid is, the more readily it dissociates in water to produce hydrogen ions. These then immediately react with the water molecules to form the hydronium ion (H_3O^+). The stronger an Arrhenius base is, the more easily it dissociates to form the hydroxide (OH^-) ion. Strong acids and bases dissociate readily and therefore conduct electricity to a greater extent than weak acids and bases that do not dissociate in water to the same extent.

According to the Arrhenius model, acids dissociate and ionise in water to produce hydrogen ions (H^+), whereas bases dissociate in water to produce hydroxide ions (OH^-). The production of these charged particles in solution allows acids and bases to conduct electricity.

Limitations of the Arrhenius theory of acids and bases

The Arrhenius model explains the properties of acids in terms of the formation of hydrogen ions by **ionisation** of the acid molecule. However, the model does have some limitations.

The model does not explain why some substances that do not contain hydrogen form acidic solutions when mixed with water. For example, acidic solutions are formed when carbon dioxide, CO_2, or sulfur dioxide, SO_2, are dissolved in water. The model also does not explain why some substances such as ammonia, NH_3, or sodium hydrogen carbonate, $NaHCO_3$, form basic solutions when mixed with water, even though they do not contain hydroxide ions.

The Arrhenius model is also restricted to acids and bases that dissolve in water. The model does not explain acid–base behaviour in non-aqueous solutions.

An improved model for acid–base behaviour is the Brønsted–Lowry model, which describes an acid as a substance that can donate a hydrogen ion (proton) to a base, which accepts the proton. You will learn more about the Brønsted–Lowry model in Unit 3.

16.1 Review

SUMMARY

- An Arrhenius acid ionises in water to produce hydrogen ions.
- In an ionisation reaction, a molecular substance reacts with a solvent such as water to form ions in solution.
- The common properties of acids are due to their ability to form the hydrogen ion, H^+, in solution.
- An Arrhenius base dissociates in water, producing hydroxide ions, OH^-.
- In a dissociation reaction, positive and negative ions separate from an ionic lattice.
- The common properties of bases are due to the presence of the hydroxide ion, OH^-, in solution.
- A proton or hydrogen ion in solution can be represented by the hydronium ion, $H_3O^+(aq)$, or simply as $H^+(aq)$.

KEY QUESTIONS

Retrieval

1 List two common acids and two common bases found in your home.

2 Use the Arrhenius model of acids and bases to identify the ions formed when the following dissociate in water.

   a potassium hydroxide
   b ethanoic acid
   c sulfuric acid

Comprehension

3 Use suitable examples to explain the difference between a strong and a weak acid.

Analysis

4 Determine some of the limitations of the Arrhenius model of acids and bases using an example in your answer.

16.2 pH: a convenient way to measure acidity

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- relate the pH of a solution to the concentration of hydrogen ions within the solution
- compare the pH values for common acids and bases.

In Module 16.1, you learnt that the Arrhenius theory defines acids as substances that ionise in water to produce hydrogen ions. Conversely, the Arrhenius model also defines bases as substances that dissociate in water to produce hydroxide ions. In this module you will learn about the pH scale, which is a measure of acidity. You will also learn about the relationship between the concentration of hydronium and hydroxide ions in different solutions.

pH OF ACIDIC AND BASIC SOLUTIONS

Pure water undergoes **self-ionisation** to a very small extent. In this reaction, water behaves as both a very **weak acid** and as a very **weak base**, producing one **hydronium ion** (H_3O^+) for every one hydroxide ion (OH^-). Water is displaying amphiprotic properties.

Water molecules can react with each other as represented by the equation:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

The production of the H_3O^+ ion and OH^- ion within this reaction can be seen in Figure 16.2.1.

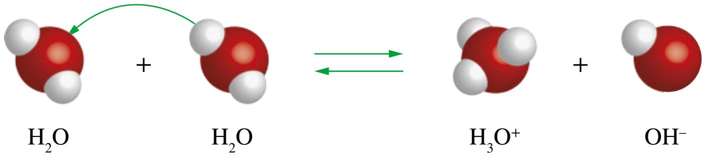


FIGURE 16.2.1 The ionisation of water molecules

IONIC PRODUCT OF WATER

As discussed earlier, pure water undergoes self-ionisation to a very small extent. The concentration of hydronium and hydroxide ions is very low. In pure water at 25°C, the H_3O^+ and OH^- concentrations are each 10^{-7} M. For each H_3O^+ ion present in a glass of water, there are 560 million H_2O molecules!

Experimental evidence shows that all aqueous solutions contain both H_3O^+ and OH^- ions, and that the product of their molar concentrations, $[H_3O^+][OH^-]$, is always 1.00×10^{-14} at 25°C. If either $[H_3O^+]$ or $[OH^-]$ in an aqueous solution increases, then the concentration of the other must decrease proportionally.

Square brackets [ ] are often used to represent molar concentration.

In solutions of acidic substances, H_3O^+ ions are formed by reaction of the acid with water, as well as from self-ionisation of water. So the concentration of H_3O^+ ions will be greater than 10^{-7} M at 25°C. Since the product $[H_3O^+][OH^-]$ remains constant, the concentration of OH^- ions in an **acidic solution** at this temperature must be less than 10^{-7} M.

The opposite is true for **basic solutions**. The concentration of OH^- ions in a basic solution is greater than 10^{-7} M and that of H_3O^+ ions is less than 10^{-7} M.

In summary, at 25°C, the ion concentrations for acid and base solutions are as follows.

- pure water and **neutral solutions**: $[H_3O^+] = [OH^-] = 10^{-7}$ M
- acidic solutions: $[H_3O^+] > 10^{-7}$ M and $[OH^-] < 10^{-7}$ M
- basic solutions: $[H_3O^+] < 10^{-7}$ M and $[OH^-] > 10^{-7}$ M

The higher the concentration of H_3O^+ ions in a solution, the more acidic the solution is.

pH: A CONVENIENT WAY TO MEASURE ACIDITY

Definition of pH

The range of H_3O^+ concentrations in solutions is so great that a convenient scale, called the **pH scale**, has been developed to measure acidity. The pH scale was first proposed by the Danish scientist Sören Sörenson in 1909 as a way of expressing levels of acidity. The pH of a solution is defined as:

$$[H_3O^+] = 10^{-pH}$$

The pH scale eliminates the need to write cumbersome powers of 10 when describing the hydrogen ion concentration. The use of pH greatly simplifies the measurement and calculation of acidity. Since the scale is based upon the negative logarithm of the hydrogen ion concentration, the pH of a solution *decreases* as the concentration of hydrogen ions *increases*.

Figure 16.2.2 shows a pH meter, which is used to accurately measure the pH of a solution.

FIGURE 16.2.2 A pH meter is used to measure the acidity of a solution.

Acidic, basic and neutral solutions can be defined in terms of their pH at 25°C.

- Neutral solutions have a pH equal to 7.
- Acidic solutions have a pH less than 7.
- Basic solutions have a pH greater than 7.

On the pH scale, the most acidic solutions have pH values slightly less than 0 and the most basic solutions have values of about 14. The pH values of some common substances are provided in Table 16.2.1.

TABLE 16.2.1 pH values of some common substances at 25°C

Solution	pH	$[H_3O^+]$ (M)	$[OH^-]$ (M)
1.0 M HCl	0.0	1	10^{-14}
lemon juice	3.0	10^{-3}	10^{-11}
vinegar	4.0	10^{-4}	10^{-10}
tomatoes	5.0	10^{-5}	10^{-9}
rain water	6.0	10^{-6}	10^{-8}
pure water	7.0	10^{-7}	10^{-7}
sea water	8.0	10^{-8}	10^{-6}
soap	9.0	10^{-9}	10^{-5}
oven cleaner	13.0	10^{-13}	10^{-1}
1.0 M NaOH	14.0	10^{-14}	10^{-0}

The reactions of acids and bases are important in a large variety of everyday applications. The high acidity of gastric juices is essential for protein digestion in the stomach. There is a complex system of pH control in your blood because even small deviations from the normal pH range of 7.35–7.45 for any length of time can lead to serious illness and death.

ℹ A solution with pH 2 has 10 times the concentration of hydronium ions as one of pH 3. The solution of pH 2 also has one-tenth the concentration of hydroxide ions as the one of pH 3.

Indicators

One of the characteristic properties of acids and bases is their ability to change the colour of certain plant extracts. Litmus is a purple dye obtained from lichen. In the presence of acids, litmus turns red. The colouring of rose petals, blackberries and red cabbage is also altered by acids and bases. Such plant extracts are called indicators. (Common indicators and their pH ranges can be seen in Figure 16.2.3.)

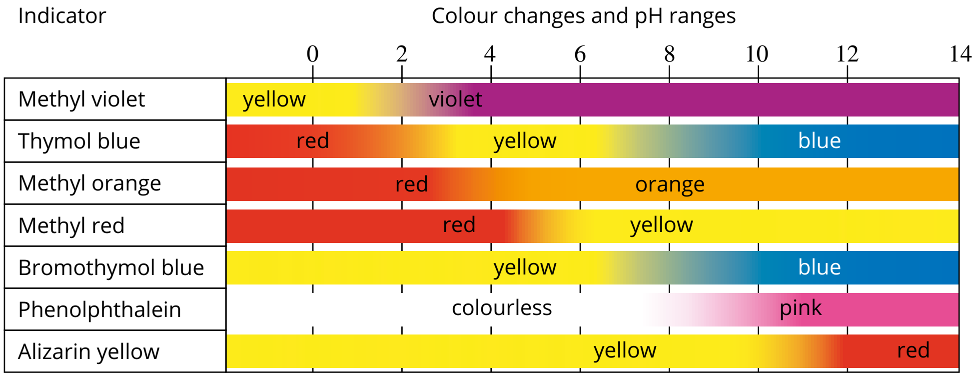


FIGURE 16.2.3 Common indicators and their pH ranges

Indicators themselves are weak acids or weak bases. The conjugate acid form of the indicator is one colour and the conjugate base form is another. The indicators that undergo a single colour change are also used for many analyses.

UNIVERSAL INDICATOR

Universal indicator (Figure 16.2.4) is widely used to estimate the pH of a solution. It is a mixture of several indicators and changes through a range of colours, from red through yellow, green and blue, to violet. If a more accurate measurement of pH is needed, you can use a pH meter instead of universal indicator.

FIGURE 16.2.4 Universal indicator pH scale. When universal indicator is added to a solution, it changes colour depending on the solution's pH. The tubes contain solutions of pH 0 to 14 from left to right. The green tube (centre) is neutral, pH 7.

CONCENTRATION OF ACIDS AND BASES

The **concentration** of acids and bases is usually expressed in units of mol L^{-1} or M. This is also referred to as molar concentration or **molarity**.

You will recall from Chapter 15 that the molar concentration of a solution, in mol L^{-1}, is given by the expression:

$$c = \frac{n}{V}$$

where c is the molar concentration (mol L^{-1}), n is the number of moles of **solute** dissolved in the solution (mol) and V is the volume of the solution (L).

The most convenient way of preparing a solution of a dilute acid is by mixing concentrated acid with water, as shown in Figure 16.2.5. This is known as a **dilution**.

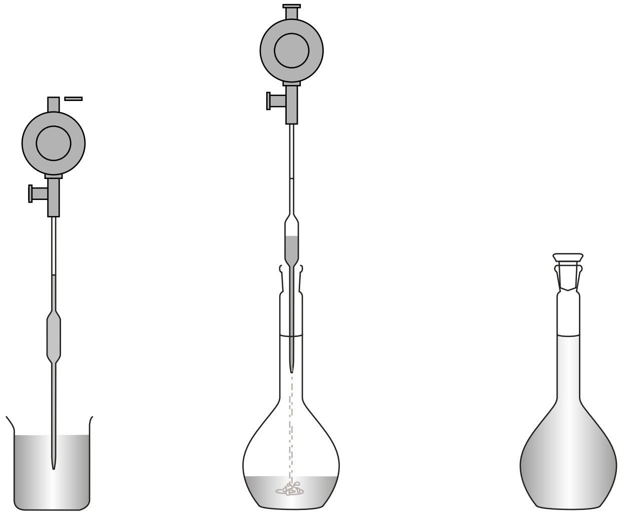


FIGURE 16.2.5 Preparing a 1.00 M HCl solution by diluting a 10.0 M solution. Heat is released when a concentrated acid is added to water, so the volumetric flask is partly filled with water before the acid is added. (Extra safety precautions would be required for diluting concentrated sulfuric acid.)

The amount of solute (in moles) in a solution does not change when a solution is diluted; the volume of the solution increases and the concentration decreases. The change in concentration or volume can be calculated using the formula:

$$c_1 V_1 = c_2 V_2$$

where c_1 and V_1 are the initial concentration and volume, and c_2 and V_2 are the concentration and volume after dilution.

You can calculate the concentration of a dilute acid if you know the:

- volume of the concentrated solution
- concentration (molarity) of the concentrated solution
- total volume of water added.

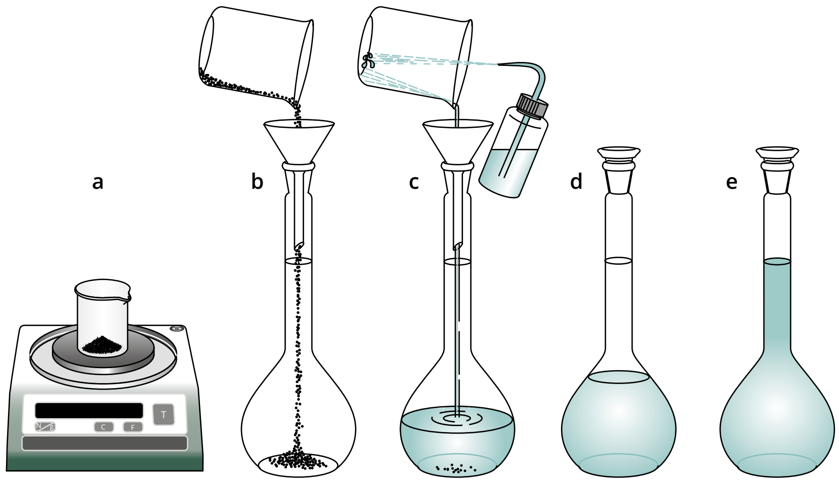
The molar concentrations of some concentrated acids are shown in Table 16.2.2.

TABLE 16.2.2 Molar concentrations of some concentrated acids

Concentrated acid (% by mass)	Formula	Molarity (M)
ethanoic acid (99.5%)	CH_3COOH	17
hydrochloric acid (36%)	HCl	12
nitric acid (70%)	HNO_3	16
phosphoric acid (85%)	H_3PO_4	15
sulfuric acid (98%)	H_2SO_4	18

In the laboratory, you can prepare solutions of a base of a required concentration by:

- diluting a more **concentrated solution**, or
- dissolving a weighed amount of the base in a measured volume of water, as shown in Figure 16.2.6.



(a) Accurately weigh out a mass of the base.
(b) Transfer the base to a volumetric flask.
(c) Ensure complete transfer of the base by washing with water.
(d) Dissolve the base in water.
(e) Add water to make the solution up to the calibration mark and shake thoroughly.

FIGURE 16.2.6 Preparing a solution by dissolving a weighed amount of base in a measured volume of water

Worked example 16.2.1

CALCULATING MOLAR CONCENTRATION AFTER DILUTION

<table>
<tr><td colspan="2">Calculate the molar concentration of hydrochloric acid when 10.0 mL of water is added to 5.0 mL of 1.2 M HCl.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>The number of moles of solute does not change during a dilution.
So, $c_1V_1 = c_2V_2$, where <i>c</i> is the concentration in M and <i>V</i> is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)</td><td>$c_1V_1 = c_2V_2$</td></tr>
<tr><td>Identify given values for concentrations and volumes before and after dilution.
Identify the unknown.</td><td>10.0 mL was added to 5.0 mL, the final volume is 15.0 mL. (In practice, small volume changes can occur when solutions are mixed; assume no volume changes occur.)
$c_1 = 1.2\,\text{M}$
$V_1 = 5.0\,\text{mL}$
$V_2 = 15.0\,\text{mL}$
You are required to calculate c_2, the concentration after dilution.</td></tr>
<tr><td>Transpose the equation and substitute the known values into the equation to find the required value.</td><td>$c_2 = \frac{c_1 \times V_1}{V_2}$
$= 1.2 \times \frac{5.0}{15.0}$
$= 0.40\,\text{M}$</td></tr>
</table>

➤ Try yourself 16.2.1

CALCULATING MOLAR CONCENTRATION AFTER DILUTION

Calculate the molar concentration of nitric acid when 80.0 mL of water is added to 20.0 mL of 5.00 M HNO_3.

Worked example 16.2.2

CALCULATING THE VOLUME OF WATER TO BE ADDED IN A DILUTION

Determine how much water must be added to 30.0 mL of 2.50 M HCl to dilute the solution to 1.00 M.

Thinking	**Working**
The number of moles of solute does not change during a dilution. So $c_1V_1 = c_2V_2$, where c is the concentration in M and V is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)	$c_1V_1 = c_2V_2$
Identify given values for concentrations and volumes before and after dilution. Identify the unknown.	$c_1 = 2.50$ M $V_1 = 30.0$ mL $c_2 = 1.00$ M You are required to calculate V_2, the volume of the diluted solution.
Transpose the equation and substitute the known values into the equation to find the required value.	$V_2 = \frac{c_1 \times V_1}{c_2}$ $= \frac{2.50 \times 30.0}{1.00}$ $= 75.0$ mL
Calculate the volume of water to be added.	Volume of dilute solution = 75.0 mL Initial volume of acid = 30.0 mL So 75.0 – 30.0 = 45.0 mL of water must be added.

➤ Try yourself 16.2.2

CALCULATING THE VOLUME OF WATER TO BE ADDED IN A DILUTION

Determine how much water must be added to 15.0 mL of 10.0 M NaOH to dilute the solution to 2.00 M.

Effect of dilution on pH of acids and bases

Consider a 0.1 M solution of hydrochloric acid. HCl is a **strong acid**, so the concentration of H_3O^+ ions in this solution is 0.1 M.

WS
2.2.5

If this 1.0 mL solution is diluted by a factor of 10 to 10.0 mL by the addition of 9.0 mL of water, the concentration of H_3O^+ ions decreases to 0.01 M and the pH increases to 2.0. A further dilution by a factor of 10 to 100 mL will increase pH to 3.0. However, note that when acids are repeatedly diluted, the pH cannot increase above 7.

Similarly, the progressive dilution of a 0.10 mL NaOH solution will cause the pH to decrease until it reaches very close to 7.

16.2 Review

SUMMARY

- Water self-ionises according to the equation:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

- At 25°C the pH of a neutral solution is 7. The pH of an acidic solution is less than 7 and the pH of a basic solution is greater than 7.
- Indicators are compounds that change colour in an acid and a different colour in a base so are commonly used to provide an indication of the pH of the solution.
- Amounts of acid or base in solution do not change during a dilution. If the volume of the solution increases then its concentration decreases.
- Solutions of acids and bases of a required concentration can be prepared by diluting more concentrated solutions using the formula:

$$c_1V_1 = c_2V_2$$

  where c_1 and V_1 are the initial volume and concentration, and c_2 and V_2 are the final volume and concentration after dilution.
- The pH increases when a solution of an acid is diluted.
- The pH decreases when a solution of a base is diluted.

KEY QUESTIONS

Retrieval

1 Refer to Figure 16.2.3 on page 510 to determine the colour of phenolphthalein indicator in a solution of pH 10.

2 Identify two plant materials whose colour is changed by the addition of an acid or a base.

3 Human blood has a pH of 7.4. Identify blood as acidic, basic or neutral.

Comprehension

4 The pH of a cola drink is 3 and the pH of black coffee is 5. Distinguish between the two solutions in terms of their acidity.

Analysis

5 Calculate the concentration of the diluted acid when 1.0 L of water is added to 3.0 L of 0.10 M HCl.

6 Calculate how much water must be added to 10 mL of a 2.0 M sulfuric acid solution to dilute it to 0.50 M.

7 Calculate the volume of water that must be added to dilute a 20.0 mL volume of 0.600 M HCl to 0.100 M.

8 The molarity of concentrated sulfuric acid is 18.0 M. Calculate the volume of concentrated sulfuric acid required to prepare 1.00 L solution of 2.00 M H_2SO_4.

16.3 Reactions of acids and bases

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand the different reactions involving acids and bases
- predict the products in reactions of acids and other compounds
- construct chemical equations to represent the reactions of acids and bases using correct formula and symbols.

Acids were originally grouped together because of their similar chemical behaviour. Chemists use indicators, such as litmus, to identify acidic solutions. Acids and bases react readily with many other chemicals, and some of the early definitions of acids and bases were derived from these reactions.

In this module, you will learn to use the patterns in the reactions of acids and bases to predict the products that are formed.

ACIDS AND REACTIVE METALS

When dilute acids are added to main group metals and some transition metals, bubbles of hydrogen gas are released and a salt is formed. In general, the equation for the reaction is:

acid + reactive metal → salt + hydrogen

Reactive metals include calcium, magnesium, iron and zinc, but *not* copper, silver or gold. For example, the reaction between dilute hydrochloric acid and zinc metal can be seen in Figure 16.3.1 and represented by the equation:

$$2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$$

This reaction can also be represented by an ionic equation. In an aqueous solution, the hydrochloric acid is ionised and the salt, zinc chloride (a soluble ionic compound), is dissociated. The ionic equation can be determined as shown in the Worked example 16.3.1 on page 518.

FIGURE 16.3.1 Hydrogen gas is produced from the reaction of zinc with dilute hydrochloric acid.

NEUTRALISATION REACTIONS

If a solution of a metal hydroxide is added to a solution of an acid (an **acid–base reaction**), the hydroxide ions will react with the hydronium ions. The acid and base are said to have been **neutralised** at the point when all the hydroxide ions have reacted with all the hydronium ions, forming water (H_2O).

Acids and metal hydroxides

Soluble metal hydroxides, such as NaOH, dissociate in water to form metal **cations** and hydroxide ions, $OH^-(aq)$. The products of a reaction of an acid with a metal hydroxide are an ionic compound, called a **salt**, and water.

The general rule for the reaction between acids and metal hydroxides can be expressed as:

acid + metal hydroxide → salt + water

For example, solutions of sulfuric acid and sodium hydroxide react to form sodium sulfate and water. This can be represented by the full (or overall) equation:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The salt formed in the reaction between sulfuric acid and sodium hydroxide is sodium sulfate. In water, the sodium and sulfate ions will dissociate, leaving Na^+ and SO_4^{2-} in solution. If water was evaporated from the reaction mixture, solid sodium sulfate would be left behind.

FIGURE 16.3.2 This statue has been heavily eroded by acid rain, which reacts with carbonate salts in limestone. Acid rain is formed when gases, such as sulfur dioxide and nitrogen oxides, dissolve in water to form acidic solutions.

ACIDS AND METAL CARBONATES

The weathering of buildings and statues (Figure 16.3.2) is due in part to the reaction between acid rain and the carbonate minerals in the stone.

Acids reacting with metal carbonates and metal hydrogen carbonates produce carbon dioxide gas, together with a salt and water. Metal carbonates include sodium carbonate (Na_2CO_3), magnesium carbonate ($MgCO_3$) and calcium carbonate ($CaCO_3$).

The general reaction for metal carbonates with acids can be summarised as:

$$\text{acid} + \text{metal carbonate} \rightarrow \text{salt} + \text{water} + \text{carbon dioxide}$$

For example, a solution of hydrochloric acid reacting with sodium carbonate solution produces a solution of sodium chloride, water and carbon dioxide gas. The reaction is represented by the equation:

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

The reaction between hydrochloric acid and sodium carbonate is represented in Figure 16.3.3.

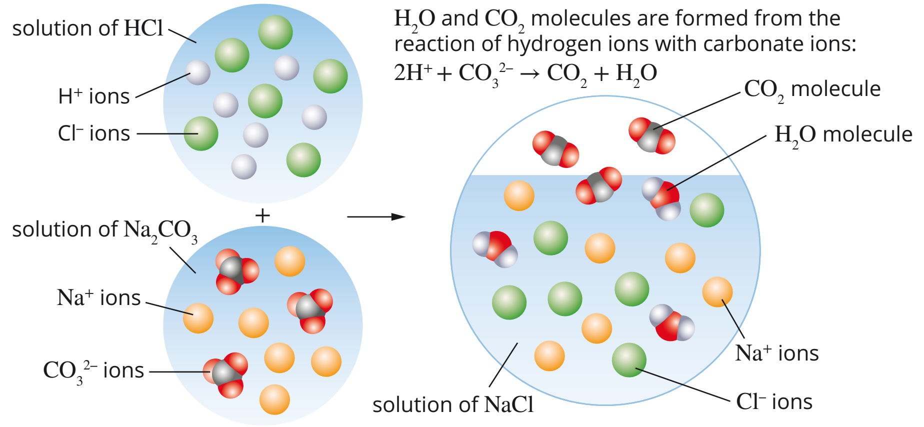


FIGURE 16.3.3 The reaction between solutions of sodium carbonate and hydrochloric acid

Metal hydrogen carbonates (also known as bicarbonates) include sodium hydrogen carbonate ($NaHCO_3$), potassium hydrogen carbonate ($KHCO_3$) and calcium hydrogen carbonate ($Ca(HCO_3)_2$). Acids added to metal hydrogen carbonates also produce carbon dioxide, together with a salt and water. The general reaction is:

$$\text{acid} + \text{metal hydrogen carbonate} \rightarrow \text{salt} + \text{water} + \text{carbon dioxide}$$

For example, when solutions of hydrochloric acid and sodium hydrogen carbonate are mixed, the following reaction occurs:

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

Salts

Salts consist of the positive ion (cation) from the base and the negative ion (**anion**) from the acid.

Table 16.3.1 lists the names of salts formed from some **neutralisation reactions** of acids with metal hydroxides. Note that the name of the positive ion is listed first and the name of the negative ion from the acid is listed second.

TABLE 16.3.1 Salts formed from some common neutralisation reactions

Reactants (acid + metal hydroxide)	Name of salt formed	Formulas of ions present in the salt solution
hydrochloric acid + potassium hydroxide	potassium chloride	$K^+(aq) + Cl^-(aq)$
hydrochloric acid + magnesium hydroxide	magnesium chloride	$Mg^{2+}(aq) + Cl^-(aq)$
nitric acid + sodium hydroxide	sodium nitrate	$Na^+(aq) + NO_3^-(aq)$
sulfuric acid + zinc hydroxide	zinc sulfate	$Zn^{2+}(aq) + SO_4^{2-}(aq)$
phosphoric acid + potassium hydroxide	potassium phosphate	$K^+(aq) + PO_4^{3-}(aq)$
ethanoic acid + calcium hydroxide	calcium ethanoate	$Ca^{2+}(aq) + CH_3COO^-(aq)$

Figure 16.3.4 is a diagrammatic representation of the ions in solutions of HCl and NaOH when mixed in a neutralisation reaction.

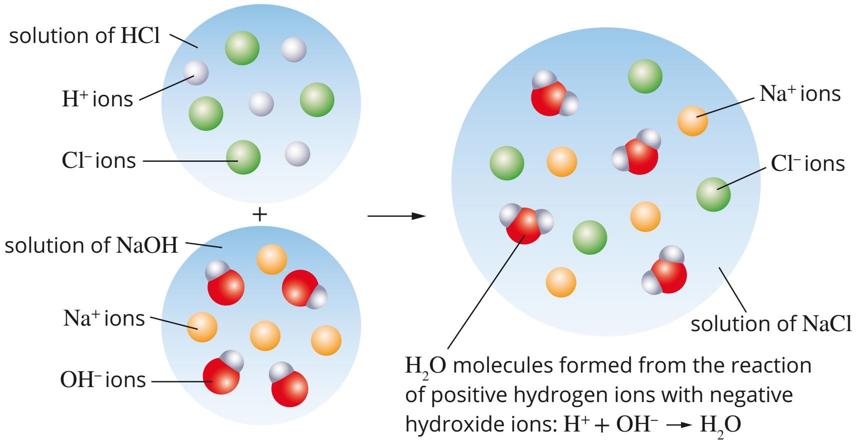


FIGURE 16.3.4 A representation of the reaction between $H^+(aq)$ and $OH^-(aq)$ ions that occurs when solutions of HCl and NaOH are mixed

TESTING FOR CARBONATE SALTS

Acids can be used to detect the presence of carbonate salts. Carbon dioxide is produced when an acid is added to a carbonate.

The **limewater test** is a simple laboratory test used to confirm the presence of carbon dioxide gas. Limewater is a saturated solution of calcium hydroxide ($Ca(OH)_2$). When carbon dioxide is bubbled through this solution, it will turn 'milky' or 'cloudy' in appearance (Figure 16.3.5) due to the precipitation of calcium carbonate:

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

FIGURE 16.3.5 Limewater test. Carbon dioxide is bubbled through limewater, turning the limewater cloudy.

WRITING EQUATIONS FOR REACTIONS OF ACIDS AND BASES

As you have read, acids and bases dissociate in water and react with each other in neutralisation reactions. When they dissociate and react with metals, hydrogen gas and a salt is formed. Examples of each of the reactions mentioned previously have been represented using chemical equations. They can also be represented using ionic equations. Writing equations and ionic equations for the different reactions will be demonstrated in a number of worked examples.

IONIC EQUATIONS

The hydroxide ions from metal hydroxides, such as sodium hydroxide (NaOH), calcium hydroxide ($Ca(OH)_2$) and magnesium hydroxide ($Mg(OH)_2$), react readily with the hydrogen ion ($H^+(aq)$) from acids.

The reaction between an acid and a metal hydroxide can be represented by an **ionic equation** as well as by an overall equation. You were introduced to ionic equations when studying precipitation reactions in Chapter 15.

When writing ionic equations, it is important to remember these points.

- Ionic equations are balanced with respect to both the number of atoms of each element and charge.
- **Spectator ions** are ions that are dissolved in the solution and are present as ions before and after the reaction, but are not involved in the reaction. These are omitted from the ionic equation.
- If a reactant or product is a solid, liquid or a gas, it cannot be written as ions and it must be present in the ionic equation.

When writing ionic equations for neutralisation reactions, it is also important to remember these points.

- Strong acids ionise in solution and are written as ions; for example, HCl in solution is written as $H_3O^+(aq) + Cl^-(aq)$. For the sake of simplicity, the hydronium ion ($H_3O^+(aq)$) can be written as a hydrogen ion ($H^+(aq)$).
- Metal hydroxides and salts are ionic and, if soluble, dissociate in solution and are written as ions; for example, KOH dissolving in water is written as $KOH(s) \rightarrow K^+(aq) + OH^-(aq)$.
- Water is a covalent molecular substance that does not ionise to any significant extent. It is written as $H_2O(l)$.

Worked example 16.3.1 indicates the steps to follow when writing ionic equations for the reactions of acids.

Worked example 16.3.1

WRITING AN IONIC EQUATION FOR AN ACID–BASE REACTION

Write an ionic equation for the reaction that occurs when hydrochloric acid is added to sodium hydroxide solution. A representation of this reaction is shown in Figure 16.3.5.

Thinking	**Working**
What is the general reaction? Identify the products formed.	acid + metal hydroxide → salt + water A solution of sodium chloride and water is formed.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: HCl(aq) is ionised in solution, forming $H^+(aq)$ and $Cl^-(aq)$. NaOH(aq) is dissociated in solution, forming $Na^+(aq)$ and $OH^-(aq)$. Products: Sodium chloride is dissociated and exists as $Na^+(aq)$ and $Cl^-(aq)$. Water is a molecular compound and its formula is $H_2O(l)$.
Write the equation showing all reactants and products, in ionised form where possible. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$
Identify the spectator ions: the ions that have an (aq) state both as a reactant and as a product.	$Na^+(aq)$ and $Cl^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to the number of atoms of each element and charge.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ Note that if hydronium ions are represented as $H_3O^+(aq)$, rather than as $H^+(aq)$, this reaction would be written as: $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$

➤ Try yourself 16.3.1

WRITING AN IONIC EQUATION FOR AN ACID–BASE REACTION

Write an ionic equation for the reaction that occurs when sulfuric acid is added to potassium hydroxide solution.

The reactions between acids and metal carbonates, and reactions between acids and metal hydrogen carbonates, can also be represented as ionic equations by following steps similar to the steps for writing reactions between acids and bases.

Worked example 16.3.2

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND METAL CARBONATES

Determine the products formed when a dilute solution of nitric acid is added to solid magnesium carbonate. Write an ionic equation for this reaction.

Thinking	**Working**
What is the general reaction? Identify the products.	acid + metal carbonate → salt + water + carbon dioxide Products of this reaction are magnesium nitrate in solution, water and carbon dioxide gas.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: Nitric acid is ionised in solution, forming $H^+(aq)$ and $NO_3^-(aq)$ ions. Magnesium carbonate is an ionic solid, $MgCO_3(s)$. Products: Magnesium nitrate is dissociated into $Mg^{2+}(aq)$ and NO_3^- (aq) ions. Water has the formula $H_2O(l)$. Carbon dioxide has the formula $CO_2(g)$.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + NO_3^-(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + NO_3^-(aq) + H_2O(l) + CO_2(g)$
Identify the spectator ions.	$NO_3^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to the number of atoms of each element and charge.	The equation with spectator ions omitted is: $H^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + H_2O(l) + CO_2(g)$ The balanced equation is: $2H^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + H_2O(l) + CO_2(g)$ Note that if hydronium ions are represented as $H_3O^+(aq)$, rather than as $H^+(aq)$, this reaction would be written as: $2H_3O^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + 3H_2O(l) + CO_2(g)$

➤ Try yourself 16.3.2

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND METAL CARBONATES

Determine the products formed when a solution of hydrochloric acid is added to a solution of sodium hydrogen carbonate. Write an ionic equation for this reaction.

PA 2.2.2

Worked example 16.3.3

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND REACTIVE METALS

Write an ionic equation for the reaction that occurs when dilute hydrochloric acid is added to a sample of zinc metal.	
Thinking	**Working**
What is the general reaction? Identify the products formed.	acid + reactive metal → salt + hydrogen Hydrogen gas and zinc chloride solution are produced.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: zinc is a solid, Zn(s). Hydrochloric acid is ionised, forming H^+(aq) and Cl^- (aq) ions. Products: hydrogen gas, H_2(g). Zinc chloride is dissociated into Zn^{2+}(aq) and Cl^-(aq) ions.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cl^-(aq) + H_2(g)$
Identify the spectator ions.	$Cl^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to the number of atoms of each element and charge.	$2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$

➤ Try yourself 16.3.3

WS 2.2.6

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND REACTIVE METALS

Write an ionic equation for the reaction that occurs when aluminium is added to a dilute solution of hydrochloric acid.

16.3 Review

SUMMARY

- Generalisations can be made about the likely products of reactions involving acids and bases:

  acid + metal hydroxide → salt + water

  acid + metal carbonate → salt + water + carbon dioxide

  acid + metal hydrogen carbonate → salt + water + carbon dioxide

  acid + reactive metal → salt + hydrogen
- Salts dissociate in solution to varying degrees dependent on their solubility. This will then affect whether they are included in full and ionic chemical equations.
- Each of these reactions can be represented by full and ionic equations.
- An ionic equation only shows those ions, atoms or molecules that take part in the reaction.
- Spectator ions (ions that do not take part in the reaction) are not included in ionic equations.
- Ionic equations are balanced with respect to both the number of atoms of each element and charge.

KEY QUESTIONS

Retrieval

1 Write full and ionic chemical equations for the reactions between:

  a magnesium and sulfuric acid

  b calcium and hydrochloric acid

  c zinc and ethanoic acid

  d aluminium and nitric acid

2 Name the salt produced in each of the reactions in Question **1**.

3 For each of the following reactions, write:

  i a full chemical equation to represent the reaction (remember to include states)

  ii an ionic equation

  a solid zinc oxide and sulfuric acid

  b solid calcium and nitric acid

  c solid copper(II) hydroxide and nitric acid

  d solid magnesium hydrogen carbonate and hydrochloric acid

  e solid tin(II) carbonate and sulfuric acid

Comprehension

4 Discuss why limewater turns cloudy in the presence of carbon dioxide gas.

Analysis

5 Predict the products of the following reactions and write full and ionic chemical equations for each.

  a A solution of sulfuric acid is added to a solution of potassium hydroxide.

  b Nitric acid solution is mixed with sodium hydroxide solution.

  c Hydrochloric acid solution is poured onto some solid magnesium hydroxide.

  d Blue copper(II) carbonate powder is added to dilute sulfuric acid.

6 Predict the products of the following reactions and write full equations for each.

  a Dilute hydrofluoric acid is mixed with a solution of potassium hydrogen carbonate.

  b Dilute nitric acid is added to a spoon coated with solid zinc.

  c Hydrochloric acid solution is added to some marble chips (calcium carbonate).

  d Solid bicarbonate of soda (sodium hydrogen carbonate) is mixed with vinegar (a dilute solution of ethanoic acid).

7 Dilute hydrochloric acid is added to a white solid in a test-tube. A colourless gas is produced. The gas turns limewater cloudy. Deduce the identity of the white solid.

16.4 Human impacts on water quality

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- explain the role that burning coal and oil in power stations plays in the contribution of sulfur dioxide, carbon dioxide and nitrogen oxide levels in the atmosphere
- understand the testing, monitoring and treatment strategies used to treat water from different sources.*

**The content covered in this module is not mandated by the Chemistry General Senior Syllabus. However, the content covered in the two Science as a Human Endeavour features in this module is included under syllabus guidance for Topic 2: Aqueous solutions and acidity.*

Since the start of the Industrial Revolution in the 1760s, there has been a dramatic increase in the combustion of fossil fuels. This has resulted in significantly increased levels of carbon dioxide and sulfur dioxide levels in the atmosphere.

In this module, you will apply your understanding of acid–base chemistry to study the effects of increasing sulfur dioxide concentrations on the environment, in particular water supplies. This is a significant environmental, social and economic issue. You will also investigate how water sources are monitored and treated to ensure it is safe for human and animal use and consumption.

ACID RAIN

Since the Industrial Revolution, human activity has produced a much greater quantity of acidic gases than the amount produced by natural processes. This has resulted in an increase in the acidity (decrease in pH) of rain falling near or downwind from industrial areas, as shown in Figure 16.4.1. This is known as acid rain. The role of sulfur dioxide emitted from power stations is investigated specifically in Science as a Human Endeavour on page 524.

FIGURE 16.4.1 Acid rain is formed downwind of industries that produce acidic gases such as sulfur dioxide. The acid rain leaches nutrients and toxic metals out of the soil, damaging vegetation.

The pH of unpolluted rainwater is between 5.5 and 6, but the pH of rainwater in industrial areas may be close to 4 and has even been measured to be as low as 2.4.

The internal combustion engine is responsible for the emission of large amounts of nitrogen(II) oxide (NO) and nitrogen(IV) oxide (NO_2). Nitrogen(IV) oxide dissolves in water, forming acid rain. This acid rain is a mixture of nitrous acid (HNO_2) and nitric acid (HNO_3):

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

The acid rain created by power stations and smelters is particularly severe in parts of the United States, Eastern Canada, Europe and Asia. Some industries, such as the smelting of ores containing metal sulfides, have produced localised acid rain in Australia. Fortunately, smelters are often located in remote areas that have little rainfall. The production of nitrogen oxides occurs mainly in cities due to greater traffic concentrations.

Effects of acid rain

The lower pH of acid rain has a number of undesirable environmental effects.

- Aquatic environments are upset when the pH of the water drops below 5.5, resulting in the death of many species of aquatic plants and animals (Figure 16.4.2).
- Acid rain leaches minerals such as calcium and magnesium from the soil. These minerals are essential to plant growth. Toxic aluminium ions are also leached from the soil in the following reaction:

  $$Al_2O_3(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2O(aq)$$

  These aluminium ions damage the roots of trees and block the transportation of water. These effects can result in the death of plant species and are a contributing cause of the forest dieback typically seen in acid rain-affected areas (Figure 16.4.3).
- Metal fittings and limestone used in buildings are corroded by acid rain. How the calcium carbonate in the stone reacts with the sulfuric acid in acid rain is shown in the following reaction:

  $$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$$

The calcium sulfate ($CaSO_4$) formed is commonly known as gypsum. Gypsum is a soft, slightly water-soluble material.

FIGURE 16.4.2 Aquatic plants and animals are killed when the pH falls too low.

FIGURE 16.4.3 These trees in Grayson Highlands National Park, Virginia, USA, have been damaged by acid rain.

SCIENCE AS A HUMAN ENDEAVOUR

Acid rain

Most sulfur dioxide released to the atmosphere comes from burning coal or oil in electric power stations. Sulfur dioxide gas is the main cause of acid rain. It is produced by the burning of sulfur-containing fossil fuels:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Sulfur dioxide is also produced in power stations and from the smelting of ores such as copper sulfide.

The sulfur dioxide is converted to sulfuric acid via a series of reactions that occur in the atmosphere:

$$SO_2(g) \rightarrow SO_3(g) \rightarrow H_2SO_4 \text{ (in rain)}$$

Overall equation:

$$H_2SO_4(l) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$$

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$$

Stricter government regulations and international treaties aimed at reducing the emissions of sulfur dioxide and nitrogen oxides have been adopted. These regulations have decreased the concentration of these gases over recent years.

The sulfuric acid in acid rain reacts with metal carbonates to produce a salt, carbon dioxide and water, and its effects are commonly seen in limestone buildings (Figure 16.4.4). Acid rain is also the cause of why even metal structures degrade more quickly in some areas of the world due to the more reactive water increasing the rate of oxidation.

Sulfur dioxide can be removed from waste gases in factories after combustion by treating them with limestone to form calcium sulfate. This reaction is the same as the reaction in which the limestone in Figure 16.4.4 forms calcium sulfate from acid rain. In the process where the sulfur dioxide is removed at the factory site, the product formed can be used to make plasterboard, which lines the internal walls of buildings.

The CSIRO states that because of Australia's sparse population and widely spread emission sources, Australia's acid rain levels are relatively low. Australian coal is also very low in sulfur, so this also aids in the lower levels of acid rain seen in this country.

Review

1 Write a balanced ionic equation for the reaction between sulfuric acid in acid rain and the limestone found in many statues and buildings.

2 Humans cannot be harmed directly by acid rain but discuss how it may still affect humans.

FIGURE 16.4.4 This limestone wall has been corroded by acid rain, causing the cavities.

MONITORING WATER QUALITY

To ensure high water quality and to test for contamination in water supplies, water samples are regularly collected for analysis. The results may be compared against water quality standards. This determines the water's suitability for its intended use and, if necessary, the treatment required to purify the water.

There are protocols for obtaining water samples to ensure that the samples are collected in a consistent manner and with the correct equipment. There are many different methods for obtaining water samples depending upon whether the samples are collected close to shore, from boats (Figure 16.4.5) or from bridges (Figure 16.4.6).

In a large dam or reservoir, it may be necessary to take samples from different depths. When comparing samples over time, samples need to be taken from the same location and in the same manner on each occasion, making it especially important to follow established protocols.

When collecting water samples for testing, it is important to consider the following questions.

- Which chemical is to be analysed and why is the testing to take place?
- What health risks are associated with the sampling?
- What equipment is required, and are sterile containers needed?
- What sample size is required for the selected test?
- What method is to be used to record the measurements?
- Where should the samples be taken and at what depths should the samples be taken?
- How will a representative sample be obtained?
- What are the labelling, storage and transport requirements for the sample?

FIGURE 16.4.5 Sampling water from a boat

FIGURE 16.4.6 Sampling water from a bridge

Protocols for water sampling

The containers used for sampling should not react with the sample. To ensure the quality of the sample, the following safeguards can be followed.

- The container can be rinsed with the sample before the final sample is taken.
- The container can be cleaned before taking another sample.
- Water used for cleaning the container can be tested to ensure there is no contamination between samples.
- If the water is to be tested for bacteria and other microorganisms, it is important that the container is sterile.

Sampling methods and equipment

The method used to avoid contamination of a water sample depends on where the water is being sampled from, and the conditions under which the sample is taken. Some sampling methods are outlined in Table 16.4.1.

TABLE 16.4.1 Sampling methods for collecting water samples

Conditions	Sampling method
A sample is taken from a well-mixed body of water.	Sampling near the surface is sufficient to obtain a representative sample. A sample should be taken about 10 cm below the surface and away from the water's edge.
A sample is taken from water contaminated by sediment (solid particles).	A sample should be drawn into the sample container by suction to avoid including the sediment.
A sample is taken from a river.	A sample should be taken upstream from where the person taking the sample stands.

EPA Queensland has published water-sampling protocols, which detail:

- the type of container
- how to sample and transport the sample
- how to preserve the sample
- the maximum holding time for the sample.

Sampling depths

To determine the depth at which a sample should be taken, the temperature is initially measured at every metre of depth.

- If the temperature is consistent, it can be assumed that the water is thoroughly mixed and a sample is taken halfway down.
- If there is a temperature variation, then a sample is taken in the middle of each temperature region.

When the water being sampled is more than 2 metres deep or samples are required from different depths, a Van Dorn sampler (Figure 16.4.7) can be used. This sampler can be lowered horizontally and pulled sideways before the sample is taken. Van Dorn samplers are made of a transparent polymer and are attached by a rope to the surface. After the sampler is raised to the surface, the water is drained and placed in a sterile sample bottle. The Van Dorn sampler is very useful when there are distinct layers in the water, such as in a deep lake or reservoir.

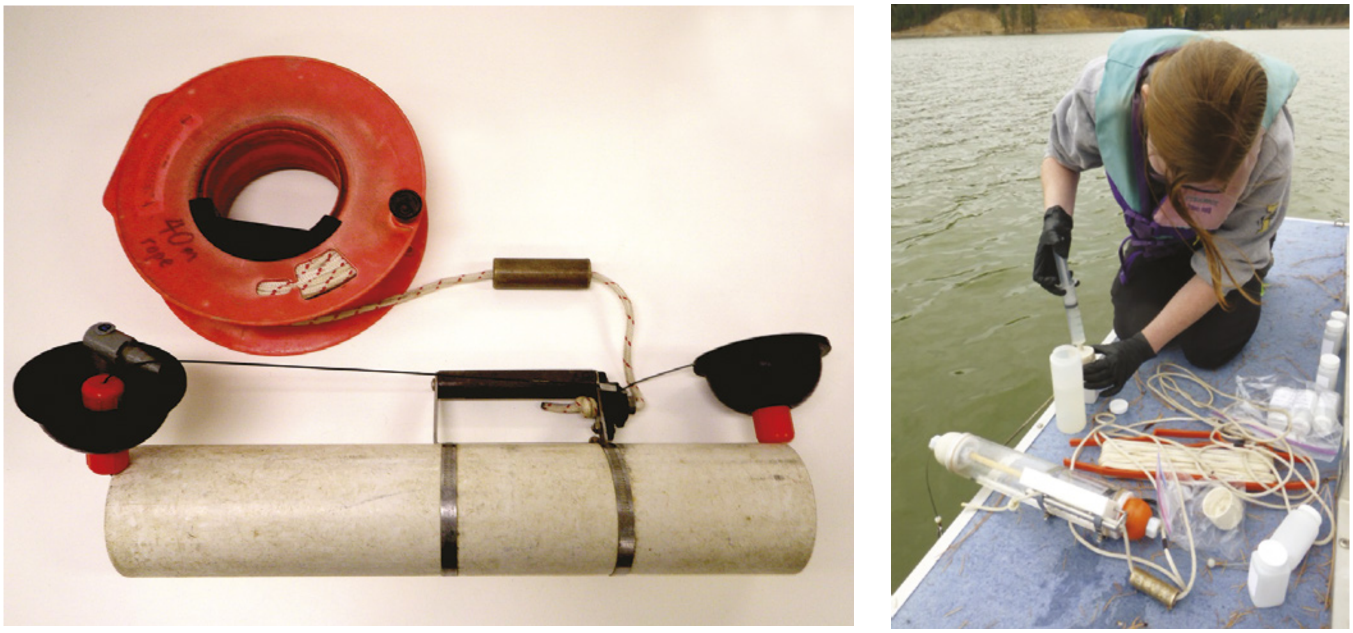


FIGURE 16.4.7 A Van Dorn sampler consists of a polymer tube with caps at each end.

Sampling locations

To develop a true picture of the water quality in a body of water, you need to take samples from many different locations.

For example, a water-quality monitoring program in the Tamar Estuary in Tasmania took samples from 20 sites divided into five zones, as shown in the map in Figure 16.4.8. The sites ranged from Launceston to where the lower estuary discharges into Bass Strait. The samples were taken monthly at specific depths and all samples were taken on the same day.

FIGURE 16.4.8 Sampling sites in the Tamar Estuary, Tasmania

SCIENCE AS A HUMAN ENDEAVOUR

Water quality

Knowledge of the composition of water from different sources informs decisions about how that water is treated and used.

In 2012, Chinese newspapers reported that 54% of China's drinking water was too polluted to drink. Just two years later, the situation in the country had worsened as the percentage of polluted drinking water rose to 60%. About 60000 people die each year in China from diseases caused by water pollution, such as cholera, dysentery and diarrhoea.

The quality of the drinking water around the world varies enormously. Most of the water in Australia is of high quality, but there have been instances where this is not the case.

Water can dissolve many different types of substances. This can cause problems when unwanted or toxic substances pollute water supplies. The dissolved materials may end up in rivers, reservoirs or groundwater, and therefore in Australia's water supplies.

In Australia, water authorities are required to provide water suitable for drinking purposes. This requires regular monitoring of water quality from catchment areas and throughout the various stages of treatment, storage and delivery.

The concentration of common ions like sodium, calcium and potassium as well as nutrients like phosphate and nitrogen are examples of some of the contaminants that are routinely tested for.

Salts are ionic compounds, and as such dissociate in water forming ion–dipole interactions with the water molecules. Pure water is a poor conductor of electricity as there are no charged particles available to carry a current. If the conductivity of a water sample is measured, the conductivity of water increases with more ions dissolved in the water, thus giving an indication of the number of dissolved salts in the water.

Gravimetric analysis is a technique that can be used to determine the amount of a particular salt present in a solution. Certain ionic compounds form precipitates in a double displacement reaction. Therefore, an aqueous solution that can react and form a stable precipitate can be used to determine the concentration of that salt using the stoichiometric ratios present in the reaction equation.

Review

1 In April 2017, the Brisbane River was contaminated with a fire-fighting foam known as perfluorooctanoic acid (PFOA), one of the per- and poly-fluoroalkyl substances (PFAS) (Figure 16.4.9). Research what impact this may have had on the water quality and how the concentration may be determined.

2 Research some of the important ways in which water is used in industry and how these might affect water quality.

FIGURE 16.4.9 Contamination in the Brisbane River, April 2017

16.4 Review

SUMMARY

- Natural and artificial processes affect the acidity of rainwater and waterways.
- Rainwater is naturally acidic due to the presence of dissolved carbon dioxide.
- Waste gases such as sulfur dioxide and nitrogen oxides that are produced by industry can cause acid rain.
- Unpolluted water in rivers and lakes has a natural pH between 6.5 and 8.5.
- Dissolved carbon dioxide tends to make river water acidic.
- In some rivers, the water is basic due to the reaction of dissolved carbonate ions from limestone with water.
- Protocols exist for sampling water for the analysis of water quality.
- All protocols need to consider:
    - the reason for testing
    - the chemical to be analysed
    - health risks associated with the sampling
    - specific requirements for the chemical being tested
    - selection of equipment and sample size
    - recording results and taking representative samples
    - labelling and storage requirements.

There are no review questions in this module as this content is not mandated by the syllabus and will not be assessed.

MANDATORY PRACTICAL 6

Relative strengths of acids

Research and planning

Aim

Measure the relative strengths of identical concentrations of some naturally occurring and manufactured acids in order to:

- determine the extent of ionisation of 1 M ethanoic acid (CH_3COOH) and 1 M hydrochloric acid (HCl)
- distinguish between some strong and weak acids.

Rationale (scientific background to the experiment)

Unlike HCl, CH_3COOH (aq) ionises only to a small extent in water:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

Because just 1% of the CH_3COOH molecules are present as CH_3COO^-, the concentration of H_3O^+ ions in 1 M CH_3COOH is about 0.01 M (pH 2). Methyl violet changes colour at a pH value of about 1.

Timing

35 minutes

Materials

- 10 mL 1 M hydrochloric acid
- 10 mL 1 M ethanoic acid
- 4 test-tubes
- 2 × 10 mL measuring cylinders
- methyl violet indicator
- safety glasses and lab coats

PRE-LAB SAFETY INFORMATION

Material used	Hazard	Control
1 M HCl	toxic by all routes of exposure; lung irritation	Wear eye and skin protection
0.1 M hydrochloric acid	corrosive to eyes and skin	Wear eye and skin protection

Please indicate that you have understood the information in the safety table.
Name (print): ______________________________________

I understand the safety information (signature): ____________

Method

Risk assessment

Consideration of risks includes chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and assessment book or download it from your eBook.

1 Prepare a solution of approximately 0.1 M HCl by pipetting 1 mL 1 M HCl to 9 mL water.

2 Prepare a 0.01 M HCl solution by diluting the 0.1 mL HCl solution in the same way.

3 Place solutions of 1 M, 0.1 M and 0.01 M HCl in test-tubes to a depth of about 5 cm.

4 To each, add three drops of methyl violet indicator. Record your pH values for each of them in the results table below.

5 Add three drops of methyl violet to a test-tube containing a similar quantity of 1 M CH_3COOH. Match the indicator colour in 1 M CH_3COOH to one of the HCl solutions.

Variables

i Independent: the concentration of acid

ii Dependent: the colour of the solution once the indicator has been added, which indicates the pH

iii Controlled: the volume of acid added, the amount of methyl violet indicator added, the individual making comparisons

Analysing

Raw data

1 Record your data in the results table.

Acid	pH using methyl violet indicator
1 M hydrochloric acid	
0.1 M hydrochloric acid	
0.01 M hydrochloric acid	
1 M ethanoic acid	

Processed data

2 Using your results, estimate the pH of the 1 M ethanoic acid by comparing the colour of the indicator in the ethanoic acid to the three HCl solutions.

Reflect and check that your data analysis demonstrates these characteristics

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.
☐ Accurate application of algorithms, visual and graphical representations of data is demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

Analysis

3 From your results, calculate the percentage of ethanoic molecules that has been ionised.

4 Rank the acids you tested according to increasing strength.

5 If HCl is considered a strong acid, state what you would class ethanoic acid as.

6 From your results, write a definition for a strong acid and a definition for a weak acid.

Reflect and check that your analysis demonstrates these characteristics

☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.
☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.
☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Interpreting and communicating

Conclusion

1 **a** State your conclusion.
  b State the evidence you have collected that leads you to this conclusion.

Evaluation

2 Considering your analysis and conclusion, discuss whether the experiment provided an effective and efficient method of determining the difference between strong and weak acids.

Improvements

3 If you were to repeat the experiment, identify the steps that you would do differently. You should include these points in your answer.
  a Explain how you would change the methodology and how this might improve the results.
  b Consider how well you performed the tasks and the skills that you need to improve on in your technique.
  c Explain how the collection of data could be improved or uncertainty reduced.

Reflect and check that your evaluation demonstrates these characteristics

☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis).
☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respect to the error analysis and limitations or sufficiency of the data.

☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment, which are logically derived from the analysis of the evidence.

Chapter review

WS 2.2.7 WS 2.2.8

16

KEY TERMS

acid
acid–base reaction
acidic solution
acidity
alkali
anion
base
basic solution
cation
concentrated solution
concentration
dilution
dissociate
hydronium ion
hydroxide ion
indicator
ionic equation
ionisation
ionise
limewater test
molarity
neutral solution
neutralisation reaction
neutralise
pH scale
salt
self-ionisation
solute
solution
spectator ion
strong acid
weak acid
weak base

KEY QUESTIONS

Retrieval

1 Determine which one of the following correctly identifies all the products formed when magnesium hydroxide reacts with hydrochloric acid.

A water
B chloride ions
C magnesium ions
D magnesium chloride precipitate
E water, magnesium ions and chloride ions

2 Identify which one of the following equations represents the reaction of a strong acid with water.

A $HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$
B $HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$
C $LiOH(aq) \rightarrow Li^+(aq) + OH^-(aq)$
D $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

3 Identify which of the following would be the likely products of the reaction of magnesium hydroxide and hydrochloric acid.

A water and magnesium hydroxide
B hydroxide ions and magnesium ions dissociated in water
C carbon dioxide, magnesium and hydrogen gas
D magnesium chloride and water

4 Identify the ionic equation that best describes the reaction that occurs when 0.5 M sulfuric acid solution is added to some solid copper(II) oxide.

A $CuO(s) + H^+(aq) \rightarrow CuOH^+(aq)$
B $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$
C $CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$
D $O^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l)$

5 Write balanced equations to show that in water:

a PO_4^{3-} acts as a base
b H_2S acts as an acid

6 Write a balanced equation for when H_2CO_3 is completely ionised in water.

Comprehension

7 Complete, and balance, the following chemical equations.

a $HNO_3(aq) + KOH(aq) \rightarrow$
b $H_2SO_4(aq) + K_2CO_3(aq) \rightarrow$
c $H_3PO_4(aq) + Ca(HCO_3)_2(s) \rightarrow$
d $HF(aq) + Zn(OH)_2(s) \rightarrow$

8 An acid solution is formed when hydrogen bromide gas is bubbled into water. Represent this ionisation reaction with a balanced equation.

9 Represent the reaction between an acid, HA, and a soluble metal carbonate, MCO_3, with a balanced ionic equation.

10 Describe the effect of acid rain on limestone.

11 Hydrogen is produced when dilute sulfuric acid reacts with aluminium metal.

a Represent this reaction with a balanced full equation.
b Represent this reaction with a balanced ionic equation.

Analysis

12 Consider the following statements.

i An acidic solution can be neutralised by the addition of a base.
ii A solution of hydrochloric acid conducts electricity.
iii Ammonia gas, NH_3, forms a basic solution when dissolved in water.

Determine which of these observations can be explained in terms of the standard Arrhenius theory of acids and bases.

13 Predict what would happen to the pH of a 1.0 M HCl solution with the addition of 1.0 M NaOH solution.

14 Create a concept map that demonstrates your understanding of the links between the following terms: acid, base, proton, hydrogen ion, hydronium ion.

Knowledge utilisation

15 Students were given four unlabelled bottles containing different acids. Each bottle contained a clear liquid and they all looked identical. They were also given several similar-sized pieces of zinc, a stopwatch and as many test-tubes as necessary.

- **a** Design an experiment that could qualitatively demonstrate the strengths of the acids.
- **b** Name the products you would expect from mixing the zinc with an acid.
- **c** Explain how you would determine the strength of the acid and then rank the acids in order of strength.

The results one student obtained are shown in the table.

Acid added to the zinc strip	**Observations**	**Time taken for reaction to go to completion (s)**
A		342
B		22
C		65
D		178

- **d** Predict the strongest acid, based on the results above.
- **e** Complete the table with the observations you might expect to see from this experiment.
- **f** Judge and rank the acids in order from weakest to strongest acid.
- **g** Propose a list of possible acids that could have been used in this experiment.

CHAPTER 17

Rates of chemical reactions

Chemical reactions occur at many different rates. The explosion of gunpowder and the combustion of petrol in a car's engine occur very quickly. On the other hand, the weathering of buildings, the ripening of fruit and the rusting of iron all occur quite slowly.

During chemical reactions, particles such as atoms, molecules and ions collide with each other and undergo a rearrangement to produce new substances. It is important to appreciate that collisions between reactant particles do not always result in a chemical reaction. For example, while a car's fuel tank is being filled with petrol, the hydrocarbon molecules in the fuel are colliding with oxygen molecules in the air without a reaction occurring.

By the end of this chapter, you will understand how rates of chemical reactions can be measured. You will also be able to describe how varying the conditions of chemical reactions can affect the rate of a reaction.

Using collision theory, you will learn to predict the effects of concentration of reactants, temperature, surface area, gas pressure and the presence of a catalyst on the rate of chemical reactions, and explain these effects.

The role of enzymes as catalysts will be considered with respect to reaction pathways and energy changes that occur during enzyme-catalysed reactions, and two models of enzymatic action will be presented.

Syllabus subject matter

Topic 3 • Rates of chemical reactions

■ RATES OF REACTIONS

- explain how varying the conditions present during chemical reactions, including temperature, surface area, pressure (gaseous systems), concentration and the presence of a catalyst can affect the rate of the reaction
- use the collision theory to explain and predict the effect of concentration, temperature, pressure and surface area on the rate of chemical reactions by considering the structure of the reactants and the energy of particles
- construct and explain Maxwell–Boltzmann distribution curves for reactions with and without catalysts
- recognise that activation energy (E_a) is the minimum energy required for a chemical reaction to occur and is related to the strength and number of the existing chemical bonds; the magnitude of the activation energy influences the rate of a chemical reaction
- sketch and use energy profile diagrams, including the transitional state and catalysed and uncatalysed pathways, to represent the enthalpy changes and activation energy associated with a chemical reaction

- explain how catalysts, including enzymes and metal nanoparticles, affect the rate of certain reactions by providing an alternative reaction pathway with a reduced activation energy, hence increasing the proportion of collisions that lead to a chemical change
- use appropriate mathematical representations to calculate the rate of chemical reactions by measuring the rate of formation of products or the depletion of reactants
- analyse experimental data, including constructing and using appropriate graphical representations of relative changes in the concentration, volume and mass against time.

SCIENCE AS A HUMAN ENDEAVOUR

- **The importance of enzymes**: Catalysts work in a variety of ways, and knowledge of the structure of enzyme molecules helps scientists to explain and predict how they are able to lower the activation energy for reactions.

MANDATORY PRACTICAL

- Investigate the rate of chemical reactions.

17.1 Investigating the rate of chemical reactions

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that chemical reactions can occur at significantly different rates
- identify the five major factors that affect the rate of chemical reactions: surface area of solid reactants; concentration of reactants in a solution; gas pressure for reactants in the gaseous state; temperature; and the presence of catalysts.

FIGURE 17.1.1 How quickly do the chemical reactions involved in baking occur?

The time it takes for a chocolate cake to cook in the oven (Figure 17.1.1) and the time taken for a fibreglass patch on a surfboard to set are related to the rate of chemical reactions.

In this module, you will learn how changes to reaction conditions affect reaction rates.

Chemical reactions are taking place all around us:

- in the soil and rocks beneath our feet
- in the air around and above us
- inside every plant and animal
- in our homes, schools and workplaces.

Some of these reactions are over in a flash. In a car accident when a car's airbag needs to be inflated, the chemical reactions producing the gas that expands the airbag need to happen extremely quickly. On the other hand, if the car's painted surface is scratched to expose the metal beneath, the rusting reactions take place at a very slow rate.

FACTORS THAT AFFECT REACTION RATES

Experimental investigations have shown that five main factors can change the rate of a **chemical reaction**:

- **surface area** of solid **reactants**
- **concentration** of reactants in a solution
- gas **pressure**
- temperature
- the presence of **catalysts**.

You can probably think of some examples of situations where one or more of these conditions are changed and a reaction becomes noticeably faster or slower.

Surface area

The surface area of solid reactants can have a significant effect on reaction rate. Smaller particles have a much larger surface area than the same mass of large particles. As a result, the smaller particles react much faster.

FIGURE 17.1.2 Particle size can be used to control the rate of reaction and create different effects during fireworks displays.

Manufacturers of fireworks modify the surface area of solid reactants to control the rate at which fuels in the fireworks burn and create different effects (Figure 17.1.2). For example, when very small pieces of aluminium are confined inside the shell they explode violently; however, if larger pieces of aluminium are used, the reaction is slower and the ejection of the burning metal pieces only creates sparks.

Concentration

The concentration of solutes dissolved in a solution can influence the rate of their reactions: higher concentrations usually lead to increased reaction rates.

Pollutants such as sulfur dioxide and nitrogen dioxide are released by cars and many industrial processes. When these compounds react with rainwater, acids such as sulfurous acid and nitric acid are formed. This causes the rainwater's hydrogen ion concentration to increase significantly and it is called **acid rain**. The increasing acidity of rain over the past 200 years has caused many famous marble buildings and statues to deteriorate much more rapidly due to the reaction between marble and the acids (Figure 17.1.3).

FIGURE 17.1.3 This limestone statue has become pitted in recent years. Limestone (calcium carbonate) reacts more rapidly with the increased concentration of hydrogen ions in rainwater.

Pressure

In reactions involving gases, increasing the pressure of the gases increases the rate at which the reaction takes place. Increasing the pressure at constant temperature will result in reactant particles becoming closer together. This will increase the frequency of collisions between gas molecules, and therefore increase the rate of reaction. The pressure of the reacting gas will increase if the amount of gas is increased, the temperature of the gas is increased or the volume of the container is decreased.

For this reason, engineers often employ high gas pressures in their design of chemical processes that use gas-phase reactions. An example is the production of ammonia gas by reacting hydrogen gas and nitrogen gas. Increasing the pressure ensures a faster rate of reaction.

Temperature

As every cook knows, the temperature of an oven affects the rate of the chemical reactions during baking. The higher the temperature, the more rapidly the reactions occur.

On the other hand, in hot weather it is wise to store fruit and vegetables in the refrigerator so that the chemical reactions that cause them to overripen and then spoil will be slowed down at the lower temperatures (Figure 17.1.4).

FIGURE 17.1.4 Food is stored at low temperatures in a refrigerator to slow down the rate of reactions that cause food spoilage.

Catalysts

Some chemical reactions occur much more rapidly if another substance is added to the reaction mixture. Such substances are called catalysts. Catalysts allow the reaction to follow a more energetically favourable pathway.

For example, if you chew a piece of dry biscuit or bread for several minutes, you may notice it tasting much sweeter. This happens because there is a catalyst present in your saliva that speeds up the breakdown of starch into sweet-tasting sugars.

In the following modules, you will learn how each of these factors can cause these changes in reaction rate.

17.1 Review

SUMMARY

- Rates of chemical reactions are affected by a number of experimental factors. The rate of a reaction may be increased by:
  - increasing the surface area of solid reactants
  - increasing the concentration of a reactant in solution
  - increasing the pressure of a gaseous reactant
  - increasing temperature
  - adding a catalyst.

KEY QUESTIONS

Retrieval

1 List the five experimental variables or factors that affect the rates of chemical reactions.

2 Identify the correct response in the statements about the four main ways in which reaction rates can be increased.

- **a** *increasing/decreasing* surface area of solid reactants
- **b** *increasing/decreasing* the temperature of a reaction mixture
- **c** *increasing/decreasing* the concentration of a reactant in solution
- **d** *increasing/decreasing* the pressure of gaseous reactants

Comprehension

3 For each of the following, determine whether the proposed change to experimental conditions will increase or decrease the rate of reaction between zinc metal and dilute hydrochloric acid.

- **a** increasing the temperature of the hydrochloric acid
- **b** decreasing the size of the pieces of zinc
- **c** decreasing the concentration of the hydrochloric acid

4 Explain why milk kept in the fridge is able to stay fresh for long periods of time but will quickly spoil at room temperature.

Analysis

5 Commercial epoxy glues consist of two viscous solutions: the epoxy resin and the hardener. The rate of hardening of epoxy glues depends on the amount of hardener that is mixed with the resin. Identify which of the five factors affecting rates of reaction this observation illustrates.

6 Assess how you could increase the rate of reaction of:

- **a** wood burning on a camp fire
- **b** a cake cooking in an oven

17.2 Quantifying rates of chemical reactions

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- define the rate of a chemical reaction
- recall experimental parameters that can be used to follow the rate of chemical reactions
- quantify the rates of chemical reactions in terms of the disappearance of a reactant or the appearance of a product
- differentiate between instantaneous rate, initial rate and average rate of reaction
- understand the relationship between reaction rate and the stoichiometry of a chemical reaction.

The **rate of reaction** is defined as the change in concentration of a reactant or product per unit time. The usual unit for the rate of reaction is $M\,s^{-1}$.

The rate of reaction is defined as the change in concentration of a reactant or product per unit time.

MEASURING RATES OF REACTION EXPERIMENTALLY

To experimentally determine the rate of reaction, either directly or indirectly, you need to measure how much of a reactant is being used up or how much of a product is being formed in a given time period.

When a reaction involves gaseous products, this might involve measuring changes in mass or gas volume with time. The graph shown in Figure 17.2.1 was obtained by measuring the mass of carbon dioxide produced in the reaction between marble chips and hydrochloric acid. The experiment was performed twice, first with large marble chips, then with small marble chips.

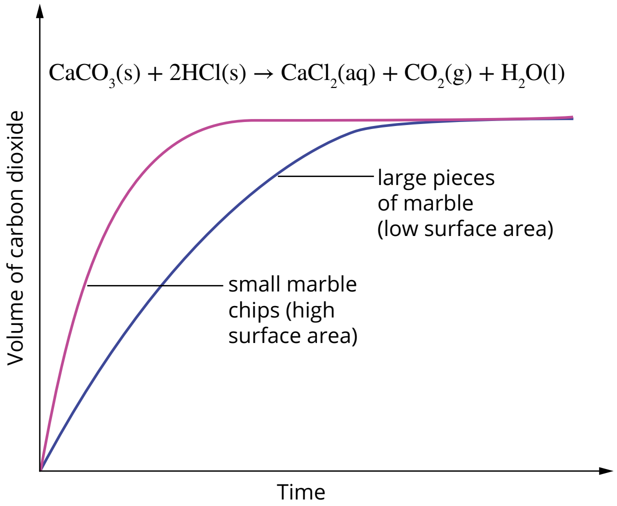


FIGURE 17.2.1 As carbon dioxide is produced from the reaction between marble chips and hydrochloric acid, it is collected in a gas syringe and its volume is recorded.

The steeper initial slope of the graph with small marble chips indicates that the rate of production of carbon dioxide gas is faster with the marble chips that have a larger surface area.

Colour changes and pH changes can also be used to follow the rate of some reactions, depending on the colour and acidity of reactants and products.

For example, in the above reaction of $CaCO_3$ with HCl, the rate of the reaction could be determined by following the change in pH as the hydrochloric acid is consumed. However, this is more complicated as the pH needs to be corrected for the small amount of water formed in the reaction to determine the amount of HCl consumed.

Quantifying rates of reaction

Consider the simplest hypothetical chemical reaction where a single reactant (A) is transformed into a single product (B) represented by the following chemical equation:

$$A \rightarrow B$$

It is important to note that the following discussion relates to the case of this simple hypothetical reaction where the stoichiometric coefficients of the reactants and products are one. A more complex reaction, in which stoichiometric coefficients are greater than one, is explained later in this module.

The chemical equation above tells us that as reactant A is being consumed, product B is being produced. It therefore follows that as the concentration of A decreases with time, the concentration of B increases with time. The course of this reaction can be followed using **concentration–time curves** (shown in Figure 17.2.2), which show how the concentrations of both A and B change with time. Concentration–time curves can be constructed by measuring the concentration of reactants and/or products at various times during the course of a chemical reaction.

Closer examination of Figure 17.2.2 shows that the initial slopes of both curves (at time $t = 0$) are quite steep compared to the slopes at later times. This indicates that the reaction is relatively fast at the beginning and slows down with time—a common feature of most chemical reactions.

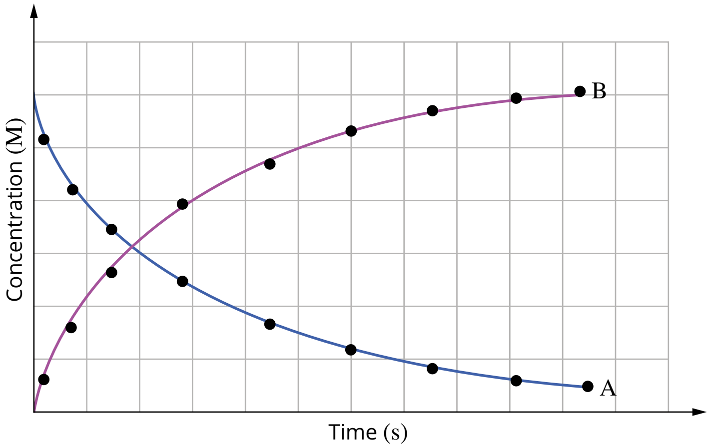


FIGURE 17.2.2 Concentration–time curves showing the changes in concentration over time for the reactant (A) and product (B) for the hypothetical reaction A ⟶ B. By convention, square brackets are used around chemical formula to refer to the concentration of that substance.

The slope of the tangent at any point along a concentration–time curve can be used to calculate the rate of a chemical reaction. Figure 17.2.3 shows that the rate of a reaction can be calculated in one of two general ways:

1 by considering the rate of disappearance of a reactant (Figure 17.2.3a)

2 by considering the rate of appearance of a product (Figure 17.2.3b).

Both are equally valid ways of quantifying the rate of reaction and both are represented mathematically by the following expressions:

$$\text{Rate of disappearance of A} = -\frac{\text{change in concentration of reactant A}}{\text{change in time}} = -\frac{\Delta[\text{A}]}{\Delta t}$$

$$\text{Rate of appearance of B} = \frac{\text{change in concentration of product B}}{\text{change in time}} = \frac{\Delta[\text{B}]}{\Delta t}$$

where Δ[A] = the change in concentration of A (i.e. [A] at t_2 − [A] at t_1), Δ[B] = the change in concentration of B (i.e. [B] at t_2 − [B] at t_1) and Δt = change in time ($t_2 - t_1$). Note that in these mathematical expressions, the 'Δ' sign means 'a change in'.

You will note the use of the negative sign in the first of these expressions. Reaction rates are always reported as positive values. Since the concentration of a reactant always decreases with time, the value $\frac{\Delta[\mathrm{A}]}{\Delta t}$ will always be a negative number. The negative sign is placed in front of the $\frac{\Delta[\mathrm{A}]}{\Delta t}$ term to ensure the rate will always be reported as a positive value.

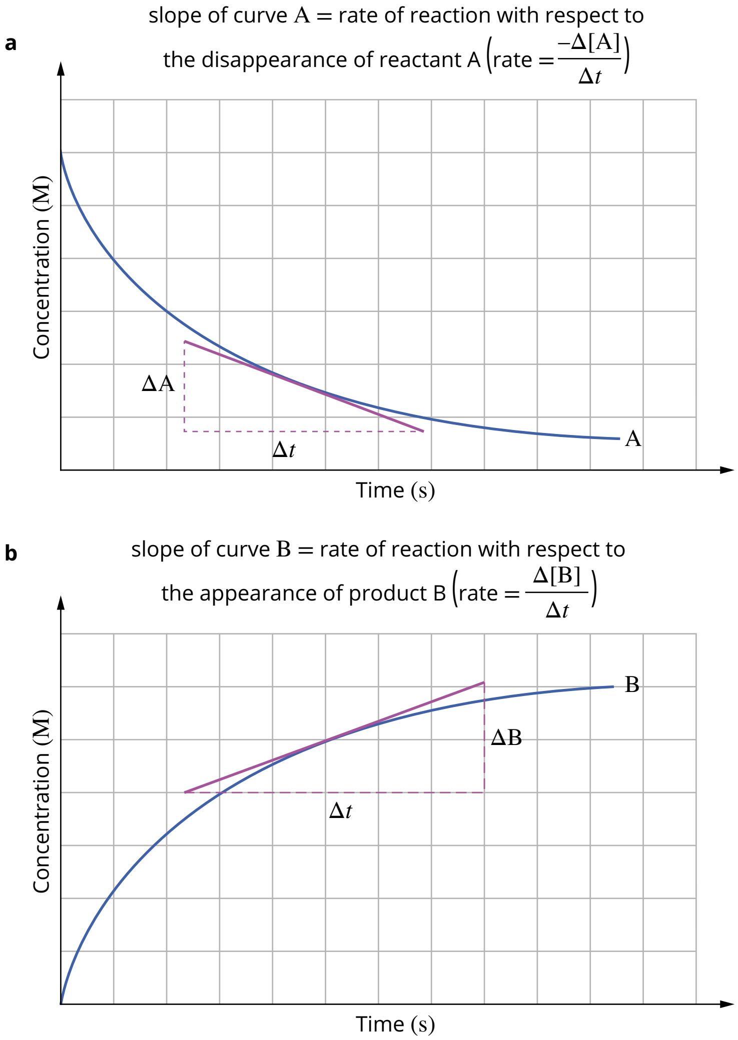


FIGURE 17.2.3 Concentration–time curves for the hypothetical reaction A ⟶ B. (a) The slope of the tangent at any point on curve A is equivalent in magnitude to the rate of disappearance of reactant A. (b) The slope of the tangent at any point on curve B is equivalent in magnitude to the rate of production of product B.

Instantaneous rate, initial rate and average rate of reaction

There are several different measures that can be used to report the rate of a chemical reaction. These include the **instantaneous rate of reaction**, the **initial rate of reaction** and the **average rate of reaction**.

The examples alluded to above and shown in Figure 17.2.3 are examples of instantaneous rates of reactions. As the name implies, the instantaneous rate of reaction is the rate of reaction at a given time during the chemical reaction. Instantaneous rates of reaction are measured as shown in Figure 17.2.3 by determining the slope of a concentration–time curve at any time point during a chemical reaction.

The initial rate of reaction refers to the instantaneous rate of reaction at time, $t = 0$. Figure 17.2.4 shows that the initial rate of reaction can be calculated by determining the initial slope of a concentration–time curve. Like the instantaneous rate, the initial rate can be calculated using the rate of disappearance of reactant (Figure 17.2.4a) or the rate of appearance of product (Figure 17.2.4b).

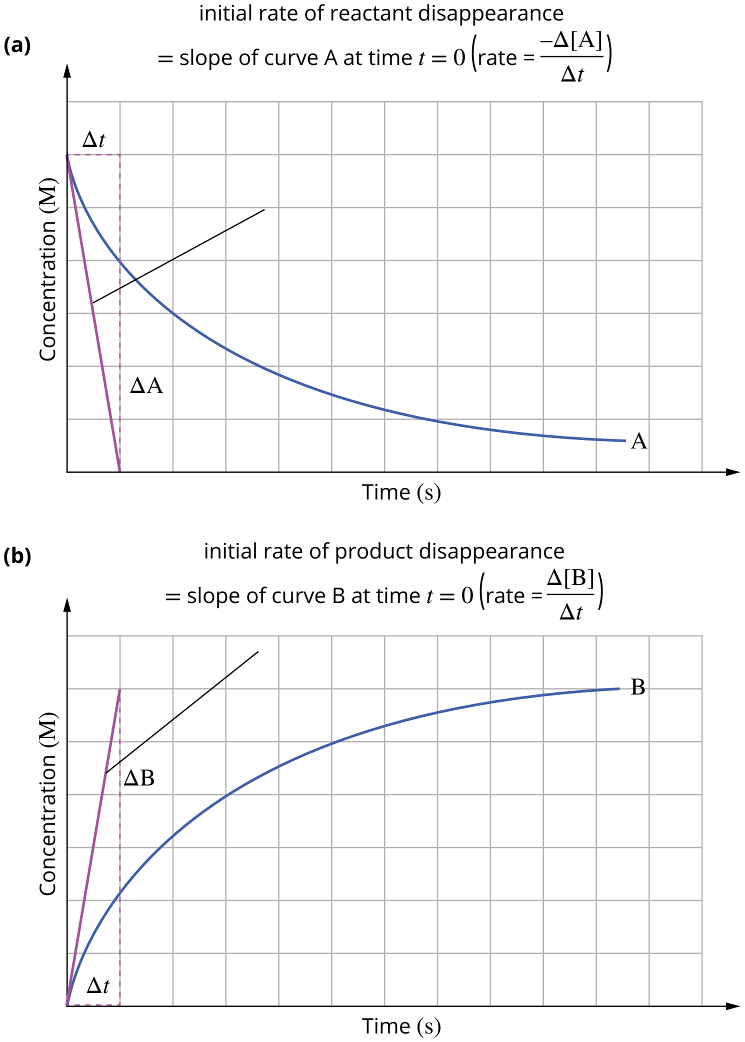


FIGURE 17.2.4 Concentration–time curves for the hypothetical reaction A ⟶ B. (a) The slope of the tangent at time $t = 0$ on curve A is equivalent in magnitude to the initial rate of reaction. (b) The slope of the tangent at time $t = 0$ on curve B is equivalent in magnitude to the initial rate of reaction.

In practice, measuring the instantaneous rate at any point in time can be difficult for many chemical reactions as it requires the construction of a complete concentration–time curve. The most common rate of reaction that is calculated is the average rate of reaction, which is determined by measuring a change in concentration between two different points in time. Figure 17.2.5 shows that if the concentration of a product (or reactant—not shown) is measured at two time points, this information can be used to calculate the average rate of reaction within that time period. Average rates of reactions should always be reported for the relevant time period.

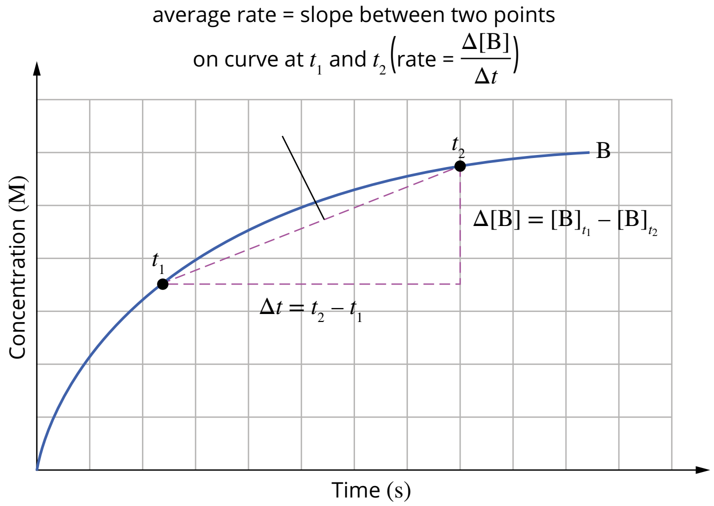


FIGURE 17.2.5 Concentration–time curves for the hypothetical reaction A ⟶ B showing that the average rate of reaction between two time points can be calculated from the slope of the line between these two points.

Worked example 17.2.1

CALCULATING AVERAGE RATES OF REACTION

The decomposition of nitrogen dioxide, NO_2, proceeds according to:

$$2NO_2(g) \longrightarrow 2NO(g) + O_2(g)$$

Experimental rate data for this reaction is given below and shows how the concentration of NO_2 decreases with time. Use this data to calculate the average rate of decomposition of NO_2 during the following time intervals.

a $t = 0\,s$ to $t = 100\,s$

b $t = 200\,s$ to $t = 300\,s$

Assume 3 significant figures for all measured values.

Time (s)	$[NO_2]$ (M)
0	0.500
50	0.394
100	0.325
150	0.275
200	0.241
250	0.210
300	0.190

Thinking	**Working**
The average rate of decomposition of NO_2 is given by: $\text{rate} = -\frac{\Delta[NO_2]}{\Delta t}$ where: $\Delta[NO_2]$ is the change in concentration of NO_2 during the stipulated time period; and Δt is the time period. Remember, the negative sign is present in the above expression because the concentration of the reactant is decreasing with time. The negative sign ensures that rate of decomposition will be reported as a positive value. The 'Δ' quantities are calculated by subtracting an initial quantity from a final quantity, i.e. final concentration minus initial concentration and final time minus initial time.	**a** In this part of the question, the relevant time period is between zero and 100 s. $\Delta NO_2 = [NO_2]_{t=100} - [NO_2]_{t=0}$ $= 0.325\,M - 0.500\,M$ $= -0.175\,M$ $\Delta t = 100\,s - 0\,s$ $= 100\,s$ Substituting these two values into the rate equation results in: $\text{rate} = -\frac{-0.175M}{100\,s} = 1.75 \times 10^{-3}\,M\,s^{-1}$ **b** In this part of the question the relevant time period is between 200 s and 300 s. $\Delta NO_2 = [NO_2]_{t=300} - [NO_2]_{t=200}$ $= 0.190\,M - 0.241\,M$ $= -0.051\,M$ $\Delta t = 300\,s - 200\,s$ $= 100\,s$ Substituting these two values into the rate equation results in: $\text{rate} = -\frac{-0.0510M}{100\,s} = 5.10 \times 10^{-4}\,M\,s^{-1}$ Note how the rate is quicker in the early part of the reaction compared to later on. This is typical of most chemical reactions.

➤ Try yourself 17.2.1

CALCULATING AVERAGE RATES OF REACTION

Consider the reaction A + B → C. Based on the table of data shown below, calculate the average rate of disappearance of A between 20 s and 40 s in $M\,s^{-1}$.

Time (s)	[A] (M)
0.0	0.224
10.0	0.210
20.0	0.188
30.0	0.173
40.0	0.154

Relationship between reaction rate and stoichiometry

The examples above show how different measures of the rate of a chemical reaction can be calculated for the simple, hypothetical case of a single reactant (A) producing a single product (B). You will now consider a real example of a chemical reaction containing multiple reactants and products where the stoichiometric coefficients take on various values.

The decomposition of dinitrogen pentoxide can be represented by the following chemical equation:

$$2N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$$

Reaction rates for this reaction can be expressed in three ways (see below) by considering the rate of disappearance of N_2O_5 as well as the rate of appearance of both NO_2 and O_2.

$$\text{Rate of disappearance of } N_2O_5 = -\frac{\Delta[N_2O_5]}{\Delta t}$$

$$\text{Rate of appearance of } NO_2 = \frac{\Delta[NO_2]}{\Delta t}$$

$$\text{Rate of appearance of } O_2 = \frac{\Delta[O_2]}{\Delta t}$$

Note: the negative sign in front of the $\frac{\Delta[N_2O_5]}{\Delta t}$ term indicates that N_2O_5 is a reactant and its concentration decreases with time.

The problem with this approach is that the three expressions for rate are not mathematically equivalent. When you look at the stoichiometric coefficients in the balanced chemical equation, you will note that for every 2 moles of N_2O_5 that is consumed in any time period, 4 moles of NO_2 will be produced. The rate of appearance of NO_2 is therefore twice as great as the rate of disappearance of N_2O_5. Similarly, for every 2 moles of N_2O_5 consumed, only 1 mole of O_2 will be produced in any given time period; the rate of appearance of O_2 is therefore one-half the rate of disappearance of N_2O_5.

As you can see, the rate of disappearance of reactants and the rate of appearance of products are not the same if the stoichiometric coefficients differ. It would be much more meaningful to define the value for the rate of reaction as being the same value regardless of the values of stoichiometric coefficients or whether you monitor the disappearance of reactants or the appearance of products.

To do this, you must take into account the value of the stoichiometric coefficient for each substance in the chemical reaction. In this case, the rate of reaction with respect to each substance in the reaction can be reported as:

$$\text{Rate of reaction} = -\frac{1}{2}\frac{\Delta[N_2O_5]}{\Delta t} = \frac{1}{4}\frac{\Delta[NO_2]}{\Delta t} = \frac{\Delta[O_2]}{\Delta t}$$

By expressing the rate of reaction in this way, each of the three expressions for rate are mathematically equivalent.

In general, for any chemical reaction that takes the form:

$$aA + bB \rightarrow cC + dD$$

where the lower-case letters represent the stoichiometric coefficients for each substance, the rate of reaction can be written as:

$$\text{Rate of reaction} = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$$

It is important to understand the difference between the rate of reaction and the rate of disappearance of reactants (or the rate of appearance of products). The rate of a chemical reaction at any point in time has only one value and is independent of the reaction stoichiometry. The rate of disappearance of reactants and the rate of appearance of products will be different to this value if the stoichiometric coefficients are different.

Worked example 17.2.2

REACTION RATE AND STOICHIOMETRY

The decomposition of HI proceeds according to:

$$2HI(g) \rightarrow H_2(g) + I_2(g)$$

a Write an expression for the rate of this chemical reaction showing the relationship between the rate of consumption of HI and the rate of production of both H_2 and I_2.

b Write an expression for the rate of decomposition of HI and show how it differs to the overall rate of reaction expression derived in part **a**.

Thinking	Working
a When writing an expression for the overall rate of a chemical reaction, you have to take into account the stoichiometric coefficients. This is particularly important when the coefficients are different for different substances in the balanced chemical equation. Using the approach outlined above, if the chemical equation takes the form of $aA + bB \rightarrow cC + dD$ the rate of reaction can be written as: $\text{Rate} = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$ where the lower-case letters represent the stoichiometric coefficients for each substance. Remember, you need to place a negative sign in front of the expression for any reactants.	**a** In this example the coefficient of HI is 2, while the coefficients of H_2 and I_2 are both 1. By multiplying by the inverse of the stoichiometric coefficients and multiplying the expression for any reactants by −1, the expression for the rate of reaction is $\text{Rate} = -\frac{1}{2}\frac{\Delta[HI]}{\Delta t} = \frac{\Delta[H_2]}{\Delta t} = \frac{\Delta[I_2]}{\Delta t}$ When written like this, the rate equation expression shows how the rate of disappearance of reactants relates to the rate of appearance of the products.
b It is important to understand the difference between the rate of a reaction and the rate of disappearance (or decomposition) of a reactant. The rate of a chemical reaction at any point in time has only one value and is independent of the reaction stoichiometry. The rate of disappearance of reactants and the rate of appearance of products will be different to this value if the stoichiometric coefficients are different.	**b** The expression for the rate of decomposition of HI is given by the change in concentration over time $\text{Rate of decomposition} = -\frac{\Delta[HI]}{\Delta t}$ The expression for the overall rate of reaction needs to take into account the stoichiometric coefficient of HI, i.e. $\text{Rate of reaction} = -\frac{1}{2}\frac{\Delta[HI]}{\Delta t}$ The rate of reaction for this reaction is therefore half of the rate of decomposition of HI.

➤ Try yourself 17.2.2

REACTION RATE AND STOICHIOMETRY

Chloryl fluoride is produced from the reaction of fluorine and chlorine dioxide. The chemical equation for this reaction is:

$$F_2(g) + 2ClO_2(g) \rightarrow 2ClO_2F(g)$$

Write an expression for the rate of reaction showing how the rate of formation of ClO_2F is related to the rate of consumption of F_2 and ClO_2.

17.2 Review

SUMMARY

- The rate of a chemical reaction is defined as the change in concentration of a reactant or product per unit time.
- The unit for rate of reaction is $M\,s^{-1}$.
- A range of experimental methods can be used to measure the rate of a reaction, including measuring the following at specific time intervals:
    - mass loss
    - volume of gas produced
    - colour change
    - concentration changes
    - pH change.
- The rate of a chemical reaction can be determined by examining the rate of disappearance of a reactant or by examining the rate of appearance of a product.
- There are three important measures of rates of reaction: the instantaneous rate of reaction; the initial rate of reaction; and the average rate of reaction between two time points. Concentration–time curves can be used to calculate each of these measures of reaction rates.
- The rate of a chemical reaction at any point in time has only one value and is independent of the reaction stoichiometry. For the general chemical reaction:

$$aA + bB \rightarrow cC + dD$$

the rate of reaction can be written as:

$$\text{Rate} = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$$

where the lower-case letters represent the stoichiometric coefficients for each substance in the balanced chemical equation.

KEY QUESTIONS

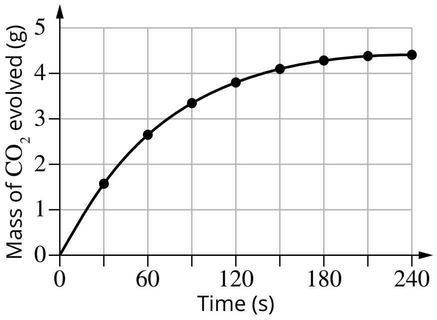
Retrieval

1 Name the unit for the rate of a chemical reaction.

2 List the three different types of reaction rates that can be obtained from the data in a concentration–time curve.

Comprehension

3 The graph below shows the mass of carbon dioxide gas produced during a 4-minute period of a reaction between marble chips (calcium carbonate) and 1.0 M nitric acid.



a Write a balanced chemical equation (including state symbols) for this reaction.

b Explain whether the rate of this chemical reaction is increasing or decreasing over time.

Analysis

4 Compare and contrast the instantaneous rate of a chemical reaction with the initial rate of a chemical reaction.

5 For the reaction of X and Y (represented by the following chemical equation) distinguish between the rate of consumption of X and the rate of reaction with respect to X.

$$2X + Y \rightarrow Z$$

6 Consider a reaction represented by the following chemical equation:

$$A + B \rightarrow C$$

Based on the concentration time data shown in the table below, calculate the average rate of reaction between each of the two times.

a 0 s and 20 s

b 20 s and 40 s

Time (s)	[A] (M)
0.0	0.290
10.0	0.220
20.0	0.178
30.0	0.156
40.0	0.148

continued over page

17.2 Review *continued*

7 The following data was obtained for the decomposition of sulfur dioxide, SO_2.

Time (s)	**$[SO_2]$ (M)**
0.0	0.852
10.0	0.636
20.0	0.462
30.0	0.326
40.0	0.218

The balanced chemical equation for this chemical reaction is:

$$2SO_2(g) \rightarrow 2SO \rightarrow + O_2(g)$$

a Calculate the average rate of decomposition of SO_2 between 10.0 s and 30.0 s.

b Calculate the average rate of reaction between 10.0 s and 30.0 s.

17.3 Collision theory

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- understand that the basis of collision theory is that reactant particles need to collide with sufficient energy and with the correct orientation to produce products
- describe how the activation energy affects the rates of chemical reactions
- differentiate between exothermic and endothermic reactions
- represent the potential energies of reactants, products and the transition state on an energy level diagram for both endothermic and exothermic reactions
- identify experimental factors that increase the rate of collisions and thereby increase the rate of reaction
- identify experimental factors that increase the energy of collisions and thereby increase the rate of reaction.

The chemical equation for a reaction indicates the nature of the reactants and products, but provides no information about the way in which the reaction proceeds.

Look at the equation for the decomposition of hydrogen peroxide:

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$

This equation gives no indication about whether the reaction proceeds quickly or slowly; nor can you tell how the products have been formed.

In fact, this reaction normally occurs very slowly, but when a catalyst, such as crystals of potassium permanganate, are added to the hydrogen peroxide solution, the reaction occurs rapidly, producing so much oxygen gas and heat that the reaction mixture foams and some of the liquid water vaporises (Figure 17.3.1).

Chemical reactions occur as a result of collisions between the reacting particles. This idea is part of the **collision theory** of reaction rates, which will be discussed in this module.

FIGURE 17.3.1 The decomposition of hydrogen peroxide is usually very slow, but the addition of crystals of potassium permanganate results in the rapid evolution of oxygen gas and water vapour.

COLLISION THEORY AND ACTIVATION ENERGY

During chemical reactions, particles (atoms, molecules or ions) collide and are rearranged to produce new particles. Consider the decomposition reaction of hydrogen peroxide:

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$

The collision that forms the first step of the reaction occurs between the two hydrogen peroxide molecules. If this collision is to result in the formation of molecules of water and oxygen, the collision must occur in such a way that the covalent bonds in the hydrogen peroxide break. To break bonds, energy is required.

The collision theory of reactions explains why some collisions result in reactions and others do not. According to collision theory, for a reaction to occur, the reactant particles must:

- collide with each other
- collide with sufficient energy to break the bonds within the reactants
- collide with the correct orientation to break the bonds within the reactants and so allow the formation of new products.

If a collision does not meet all of these requirements, then no reaction occurs. In fact, most collisions do not result in a chemical reaction. Collision theory explains why this is the case.

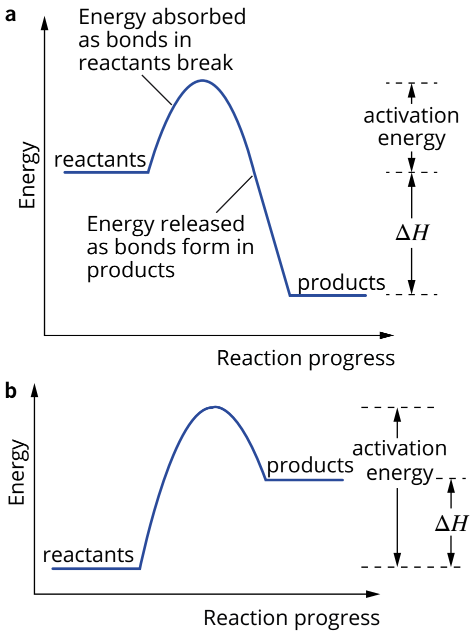


FIGURE 17.3.2 The characteristic shape of an energy profile diagram for: (a) an exothermic reaction, and (b) an endothermic reaction

Reactant particles must have energy equal to or greater than the activation energy before a reaction can occur.

Activation energy

For a reaction to occur between reactant molecules, the molecules must collide with a certain minimum amount of energy. Unless this minimum amount of energy is met or exceeded, the colliding molecules will rebound and simply move away from each other without reacting.

The minimum energy that a collision must possess for a reaction to occur is called the **activation energy**, E_a. When the energy of a collision is equal to or greater than the activation energy, a reaction can occur.

Activation energy can be represented on an **energy profile diagram**. Energy profile diagrams for both **endothermic** and **exothermic** reactions are shown in Figure 17.3.2. An energy profile diagram represents the potential energies of the reactants and the products over the course of the reaction.

The energy profile diagram for an exothermic combustion reaction like the one shown in Figure 17.3.2a indicates that the enthalpy of the products is always less than the enthalpy of the reactants. Overall, energy is released and so the ΔH value is negative. The energy profile also shows that, even in exothermic reactions, the activation energy must first be absorbed to start the reaction.

The energy profile diagram for an endothermic reaction (Figure 17.3.2b) shows that the enthalpy of the products is greater than the enthalpy of the reactants. Overall, energy is absorbed and so the ΔH value is positive. The energy profile also shows the absorption of the activation energy before the release of energy as bonds form in the products.

Energy profile diagrams for both exothermic and endothermic reactions have a peak that represents the activation energy. This is sometimes referred to as the activation energy barrier, and represents the minimum energy that must be absorbed in order to break the bonds of reactants so that a chemical reaction can progress. The activation energy is measured from the energy of the reactants to the top of the peak.

You will recall that an exothermic reaction releases more heat energy during the reaction than it absorbs. An endothermic reaction absorbs more heat energy during the reaction than it releases. This is represented on the diagram as ΔH and is the difference in energy between the reactants and the products.

Transition state

When the activation energy is absorbed, a new arrangement of the atoms known as the **transition state** occurs. The transition state occurs at the stage of maximum **potential energy** in the reaction: the activation energy (Figure 17.3.3). Bond breaking and bond forming are both occurring at this stage, and the arrangement of atoms is unstable. The atoms in the transition state rearrange into the products as the reaction progresses.

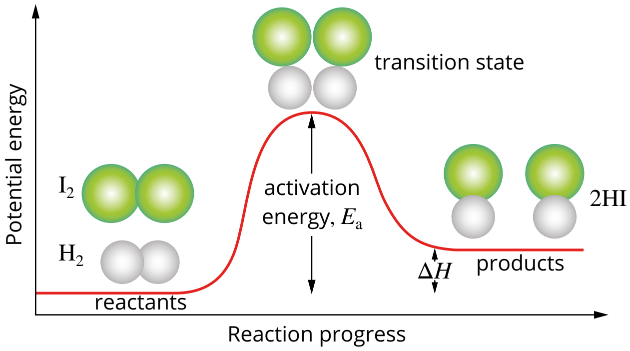


FIGURE 17.3.3 Energy profile diagram for the endothermic reaction $H_2(g) + I_2(g) \rightarrow 2HI(g)$

Activation energy and reaction rate

The size of the activation energy determines how easy it is for a reaction to occur and therefore what proportion of collisions results in a successful reaction. For this reason, the reaction rate is dependent upon the activation energy.

The existence of an activation energy for a reaction means that collisions between reactants do not always result in a chemical change. For example, nitrogen (N_2) and oxygen (O_2) molecules collide frequently in the air around us at room temperature. However, it is only when energy is provided by a spark, such as in car engines or during a lightning strike, that the energy of the collisions is increased enough to overcome the activation energy barrier. This allows nitrogen monoxide to be produced. The nitrogen monoxide formed in this process can then react to form brown nitrogen dioxide (NO_2), a poisonous gas that is a major contributor to the formation of the **photochemical smog**, as seen in Figure 17.3.4.

FIGURE 17.3.4 Photochemical smog, such as seen here over Barcelona, Spain, in December 2013, is caused by NO_2, a poisonous gas.

Orientation of colliding molecules

For a reaction to occur, reactants need to collide with enough energy to provide the activation energy. Reacting molecules must also collide with each other in the correct orientation in such a way that particular bonds in the reactants are broken and new bonds are formed in the products.

Figure 17.3.5 shows the importance of collision orientation. The decomposition of hydrogen iodide gas to produce hydrogen gas and iodine gas can be represented by the following chemical equation:

$$2HI(g) \rightarrow H_2(g) + I_2(g)$$

For this reaction to proceed the two hydrogen iodide molecules must collide with the hydrogen and iodine atoms orientated towards each other, for a reaction to possibly occur. If the collision orientation is incorrect, the particles simply bounce off each other, and no reaction occurs.

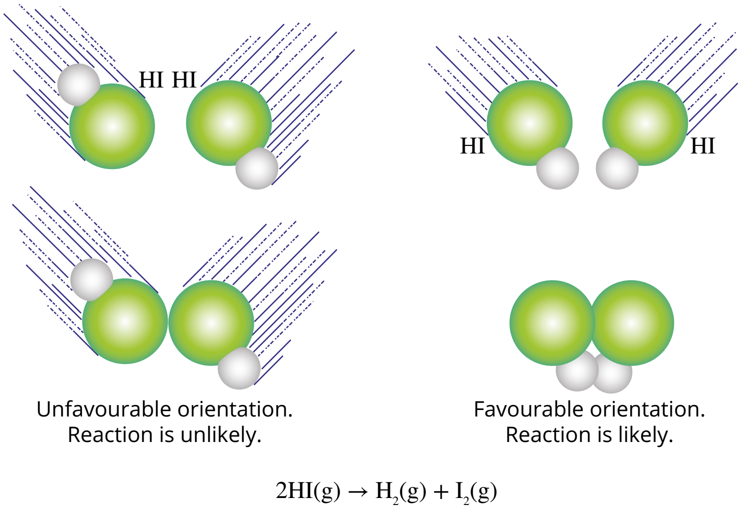


FIGURE 17.3.5 A reaction between colliding molecules is more likely to occur if the orientation of the collision is favourable.

INCREASING COLLISION FREQUENCY AND COLLISION ENERGY

Earlier in Module 17.1, you saw that the rate of a reaction can be changed by the:

- surface area of a solid reactant
- concentration of reactants in a solution
- pressure of any gaseous reactants
- temperature of the reaction
- presence of a catalyst.

In any given reaction mixture, only a certain proportion of the collisions are successful. To increase a reaction rate, you can increase the proportion of successful collisions by:

- increasing the frequency of successful collisions by increasing the number of collisions that can occur in a given time
- increasing the proportion of collisions that have energy that is equal to or greater than the activation energy (i.e. overcome the activation energy barrier).

In this module, you will consider how various changes to conditions can affect the proportion of successful collisions that occur between reactant particles and hence change the rate of reaction.

Increasing frequency of collisions

The rate of a reaction increases as the frequency of collisions increases. The frequency of collisions between reactants can be increased by:

- increasing the concentration of the reactants—collisions occur more frequently when particles are closer together
- increasing the surface area of a solid reactant.

Changing concentration or pressure

In Module 17.1, you learnt that the rate of a reaction can be increased by increasing the concentration of a reacting solution or the pressure of a reacting gas. This can be explained by collision theory.

The rate of a reaction increases when the frequency of collisions between reactants increases. When the concentration of a solution increases, there are more reactant particles moving randomly in a given volume of solution (Figure 17.3.6). The frequency of collisions consequently increases and so more successful collisions occur in a given time.

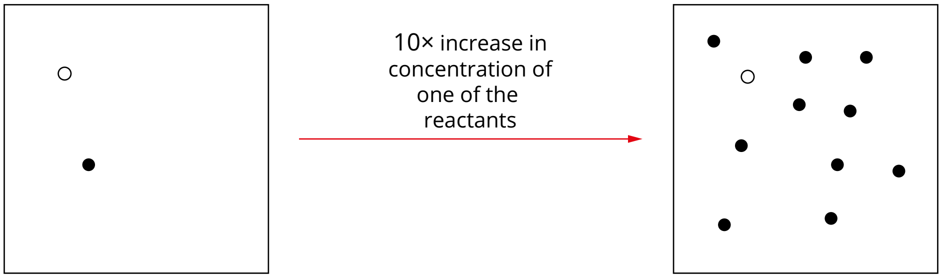


FIGURE 17.3.6 On the left, the concentration of both reactants is low. On the right, the concentration of one of the reactants has been increased 10-fold, resulting in an increase in collision frequency.

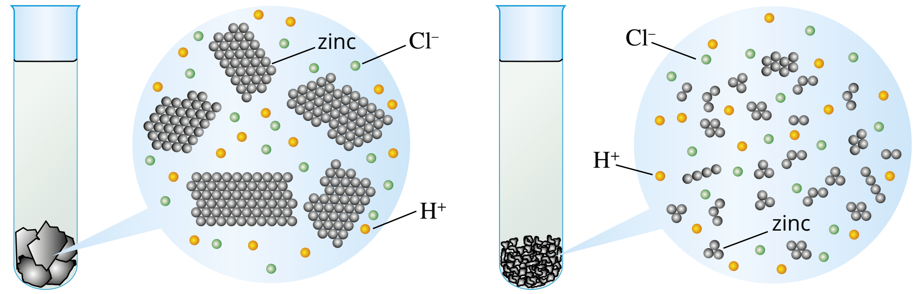
When using collision theory to explain the effect of concentration, pressure and surface area on the rate of a chemical reaction, you must discuss the effect on either collision frequency or the number of successful collisions per unit time.

For a reaction in the gas phase, the pressure of the gases can be increased either by adding more gas to a fixed volume container or by decreasing the volume of a container with a variable volume, such as a gas syringe. Increasing the pressure increases the concentration of gas molecules, causing more frequent collisions and increasing the number of successful collisions in a given time.

Changing surface area

When a solid is involved in a reaction, only the particles at the surface of the solid are involved in the reaction. The number of particles at the surface depends on the **surface area** of the substance. As you can see in Figure 17.3.7, breaking a solid into smaller parts means that more particles are present at the surface and available to react. The surface area has increased.

As a consequence of the greater number of exposed particles, the frequency of collisions between these particles and the other reactant particles increases, and so the reaction occurs more rapidly.



granules of Zn

Granules of zinc metal react slowly with HCl.

powdered Zn

Zinc powder with a large surface area reacts rapidly with HCl.

FIGURE 17.3.7 The reaction of hydrochloric acid and zinc. As the surface area of zinc increases, the rate of reaction with hydrochloric acid increases. (The H_3O^+ ion is represented as H^+ in the diagrams.)

The effect of increasing surface area can be seen when setting up an open fire at home or on a camp. It is best to first light a pile of kindling rather than trying to directly light large logs. The kindling has a larger surface area than the logs, so it catches fire more easily and will then burn rapidly, providing enough sustained heat energy to make the logs catch fire as well.

Worked example 17.3.1

USING COLLISION THEORY TO EXPLAIN CHANGES IN RATES OF REACTIONS

There have been many explosions in underground coal mines due to the presence of coal dust. Explain this observation in terms of collision theory.

Thinking	**Working**
Consider the state of the reactants.	Coal is a solid. In the mine, there would be lumps of coal and also powdered coal.
Relate the state of the reactant to the factor that affects the reaction rate and explain in terms of collision theory.	The surface area of powdered coal is greater than that of solid coal. When the surface area increases, the frequency of collisions increases and so the rate of reaction increases.
Return to the question to complete your answer.	An explosion is a very fast reaction. The very large surface area of the coal dust allows for an increase in the frequency of collisions with reacting particles, which increases the reaction rate so much that explosions occur.

➤ Try yourself 17.3.1

USING COLLISION THEORY TO EXPLAIN CHANGES IN RATES OF REACTIONS

WS 2.3.2

Iron anchors recovered from shipwrecks at considerable depths can show little corrosion after years in the sea, whereas anchors recovered from shallow water are badly corroded. Explain this observation in terms of collision theory.

Increasing the energy of collisions

As you have seen, a reaction can be made to occur more rapidly by increasing the concentration of the reactants or, for solids, increasing the surface area. A change in temperature can also have a major effect on the rate of a reaction. An increase in temperature not only increases the frequency of collisions, but also increases the **kinetic energy** of the particles and hence the energy of their collisions.

Maxwell–Boltzmann distribution

At any particular temperature, the particles in a substance have a range of kinetic energies. Although most of the particles have similar kinetic energies, there are always some particles with a high energy or a low energy. This range of energies is shown on a graph called a **Maxwell–Boltzmann distribution curve**, also known as a **kinetic energy distribution diagram**. Figure 17.3.8 shows how the distribution of energies is represented in a Maxwell–Boltzmann distribution curve.

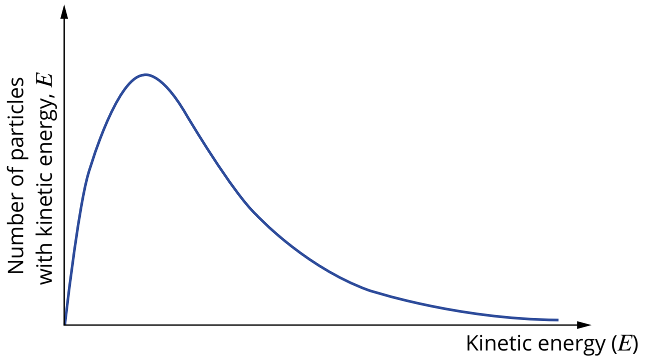


FIGURE 17.3.8 This Maxwell–Boltzmann curve shows the distribution of energies of particles in a sample at a particular temperature. The peak of the graph corresponds to the energy of the greatest number of particles.

During a reaction at a given temperature, only a small proportion of the reactant particles have kinetic energy that is equal to or greater than the activation energy and so are able to react. You can see this in Figure 17.3.9 as a small shaded area to the right of the activation energy, E_a.

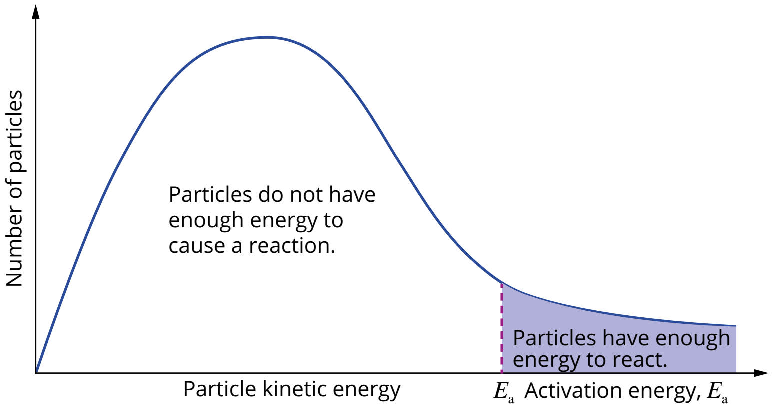


FIGURE 17.3.9 Only a small number of higher energy particles (represented by the shaded area) have sufficient energy to overcome the activation energy barrier.

Effect of temperature on rate of reaction

The relationship between kinetic energy and velocity is given by the formula $\text{KE} = \frac{1}{2}mv^2$. As the temperature of a reaction system increases, the average kinetic energy of the particles increases and the average speed of the particles in the system increases as well.

This is illustrated in Figure 17.3.10 in which the range of kinetic energies for a gas at three different temperatures is shown. Note that the area under the curve, which is equal to the total number of particles in the sample, stays constant when the temperature is changed. As the temperature increases, the increasing average kinetic energy of the particles can be seen by the movement to the right of the peak in the Maxwell–Boltzmann curve.

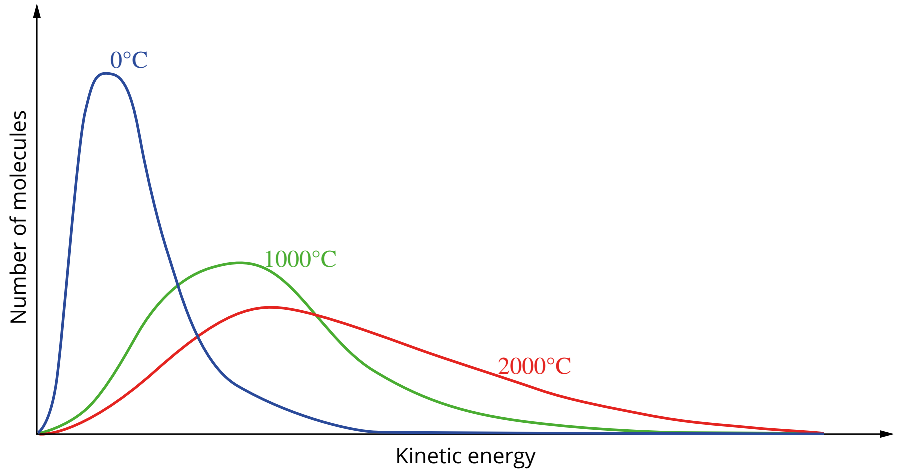


FIGURE 17.3.10 Kinetic energy distribution for a gas at a range of temperatures

As the temperature of a reaction increases, the increased speed of particles causes more collisions, increasing the frequency of successful collisions and the rate of reaction.

The increased kinetic energy of the particles also means that collisions occurring at higher temperatures have greater energy than those at lower temperatures. More particles will have energies that are greater than or equal to the activation energy and so the proportion of successful collisions also increases.

When the temperature increases, the increase in reaction rate due to the increased energy of the collisions significantly outweighs the increase in reaction rate due to the increased frequency of collisions.

A temperature increase of just 10°C causes the rate of many reactions to double (an increase of 100%), but it can be shown that this is not due to the increased frequency of collisions. The frequency of collisions only increases by about 3% when the temperature increases by 10°C. The main reason why the reaction rate increases is that more particles have sufficient energy to overcome the activation energy barrier of the reaction.

ℹ The effect of increasing temperature on the rate of reaction is mainly through increasing the proportion of reacting particles that have energies equal to or greater than the activation energy for the reaction. This increases the proportion of successful collisions.

17.3 Review

SUMMARY

- Collision theory is a theoretical model that accounts for the rates of chemical reactions in terms of collisions between particles during a chemical reaction.
- According to collision theory, for a reaction to occur, the reactant particles must:
  - collide with each other
  - collide with sufficient energy to break the bonds within the reactants
  - collide with the correct orientation to break the bonds within the reactants and so allow the formation of new products.
- The activation energy of a reaction is the minimum amount of energy required to break reactant bonds to allow a reaction to proceed. It is the minimum amount of energy that a collision between reactant particles must possess for a reaction to occur.
- The transition state is an arrangement of atoms that occurs when the activation energy is absorbed. The transition state is an unstable state in which bonds in the reactants are being broken and bonds in the products are starting to form.
- An energy profile diagram shows the activation energy as a peak of highest potential energy between the energy of the reactants and the energy of the products.
- Collision theory can be used to explain the increase in rate of reaction by an increase in:
  - concentration of a reactant solution
  - pressure of a gaseous reactant
  - surface area of a solid reactant
  - temperature.
- Increase in concentration, pressure or surface area results in an increase in the:
  - frequency of collisions between reactants
  - number of successful collisions in a given time.
- Increase in temperature results in an increase in the:
  - frequency of collisions between reactants
  - number of successful collisions in a given time
  - (most importantly) energy of collisions, so an increased proportion of collisions has an energy larger than the activation energy for the reaction.

KEY QUESTIONS

Retrieval

1 Most collisions between reacting particles are not successful in terms of forming products. Name the other factors that are required for successful collisions between reactants.

2 Define the activation energy of a chemical reaction.

3 Identify the appropriate term for the following statements.

a In an exothermic reaction the energy of the products is *higher/lower* than the energy of the reactants.

b In an endothermic reaction the energy of the products is *higher/lower* than the energy of the reactants.

Comprehension

4 The reaction between magnesium (Mg(s)) and hydrochloric acid (HCl(aq)) can be represented by the following chemical equation:

$$\text{Mg(s)} + 2\text{HCl(aq)} \longrightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$

Use collision theory to explain why this reaction proceeds faster with 1.0 M HCl compared to 0.1 M HCl.

5 Use the appropriate term to explain the following observation.

A chemical reaction is observed to be occurring very slowly.

a The activation energy may be *very large/very small*.

b The temperature may be *high/low*.

c The frequency of collisions may be *too low/too high*.

Analysis

6 Apply your knowledge of the collision model of particle behaviour to explain the following observations:

a Surfboard manufacturers find that fibreglass plastics set within hours in summer but may remain tacky for days in winter.

b A bottle of fine aluminium powder has a caution sticker warning ‘Highly flammable, dust explosion possible’.

c A potato cooks much more slowly in a pot of boiling water on a trekking holiday in Nepal than a potato boiled in a similar way in the Australian bush.

7 When 1 mole of methane gas burns completely in oxygen, the process of bond breaking uses 3380 kJ of energy and 4270 kJ of energy is released as new bonds form.

- **a** Write a balanced chemical equation for the reaction.
- **b** Calculate the value of the heat of reaction, ΔH, for the reaction.
- **c** Draw and label a diagram to determine the changes in energy during the course of the reaction.

8 The formation of hydrogen iodide from its elements is represented by the equation:

$$H_2(g) + I_2(g) \rightarrow 2HI(g)$$

This reaction has an activation energy of +167 kJ mol^{-1} and the heat of reaction, ΔH, is +28 kJ mol^{-1}. Determine the activation energy for the reverse reaction, the decomposition of 2 moles of hydrogen iodide.

9 A number of experiments involving the reaction between 100 mL of hydrochloric acid and 5 g of calcium carbonate were carried out. Sort experiments A–F and place them in increasing order of rate of reaction (slowest to fastest).

- **A** Powdered $CaCO_3$ and 1 M HCl are mixed at 40°C.
- **B** Small pieces of $CaCO_3$ and 1 M HCl are mixed at 15°C.
- **C** Powdered $CaCO_3$ and 1 M HCl are mixed at 25°C.
- **D** Large pieces of $CaCO_3$ and 0.5 M HCl are mixed at 15°C.
- **E** Powdered $CaCO_3$ and 1 M HCl are mixed at 15°C.
- **F** Small pieces of $CaCO_3$ and 0.5 M HCl are mixed at 15°C.

17.4 Catalysts

BY THE END OF THIS MODULE, YOU SHOULD BE ABLE TO:

- recall that a catalyst increases the rate of chemical reaction by providing an alternative reaction pathway with a lower activation energy compared to the uncatalysed reaction
- recall that catalysts are not consumed in chemical reactions
- differentiate between a homogeneous catalyst and heterogeneous catalyst
- construct energy profile diagrams for catalysed reactions
- recall that enzymes are large biological molecules (typically proteins) that catalyse a wide range of biochemical processes
- understand the mechanism of enzyme catalysis in terms of two models: the lock-and-key model; and the induced-fit model
- describe the three-step reaction scheme for enzyme-catalysed reactions
- construct energy profile diagrams for enzyme-catalysed reactions.

In this module, you will learn how catalysts increase the rate of reaction.

The change in reaction rate when a catalyst is present is often very substantial. The action of a catalyst can be understood using collision theory and the changes in energy that occur during a chemical reaction. Catalysts play an important role in industrial chemistry, in limiting air pollution and in controlling biochemical processes, as shown in the examples in Figure 17.4.1.

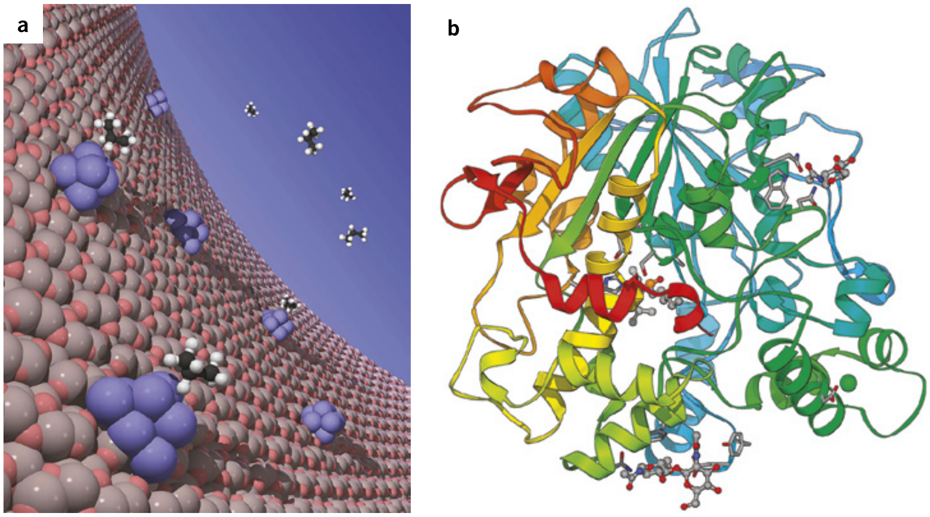


FIGURE 17.4.1 (a) A representation of the structure of a synthetic zeolite catalyst that is widely used in petroleum refineries to break down large hydrocarbon molecules into smaller, more useful molecules. Zeolite is a silica–alumina mineral. (b) A model of a lipase enzyme. Lipase is a catalyst that breaks down fats in the digestive system of the human body.

CATALYSTS AND ACTIVATION ENERGY

You have already learnt that the potential energy changes associated with a reaction can be represented as an energy profile diagram of the reaction. An energy profile diagram for an exothermic reaction is shown in Figure 17.4.2.

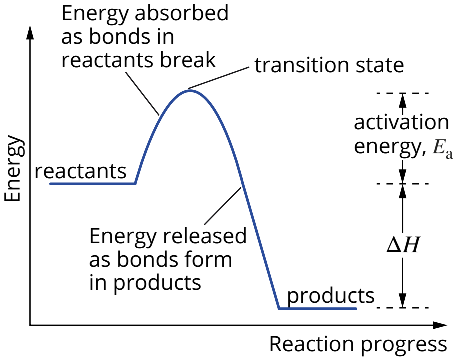


FIGURE 17.4.2 Energy profile diagram of an exothermic reaction such as burning natural gas

The activation energy is the minimum amount of energy required for a reaction to take place. On the energy profile diagram, the activation energy is measured from the energy of the reactants to the peak of the energy profile diagram.

The enthalpy change, ΔH, can also be represented on an energy profile diagram and is equal to the difference in energy between the products and the reactants.

Some reactions occur very readily because they have very small activation energies, E_a. These reactions need only a small amount of energy to be absorbed for bonds in the reactants to be broken.

Many reactions occur much more rapidly in the presence of a particular element or compound. These substances are known as catalysts. Catalysts are not consumed during the reactions they speed up, and so do not appear as either reactants or products in chemical equations.

Catalysts are able to cause a reaction to occur more quickly because they provide a new **reaction pathway**, which causes the activation energy barrier of the overall reaction to be dramatically reduced, as seen in the energy profile diagram in Figure 17.4.3.

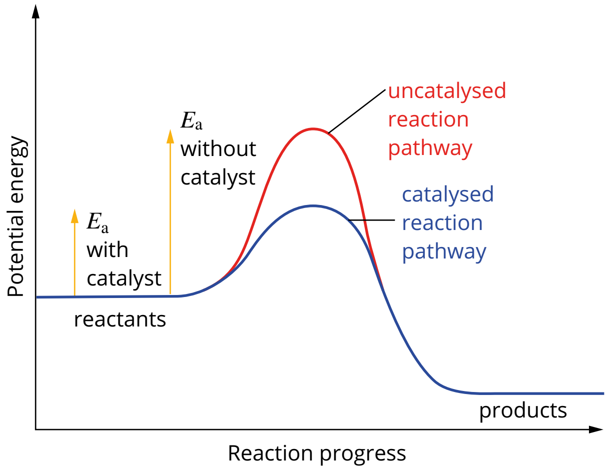


FIGURE 17.4.3 Energy profile diagrams of a catalysed and uncatalysed reaction

Catalysts only lower the activation energy for a reaction. There is no change to ΔH for the reaction.

With the catalyst present and a lower activation energy, the colliding particles are more likely to have energies that exceed this lower barrier, causing the bonds in the reactants to be broken more frequently. As a result, a greater proportion of collisions are 'successful'; that is, they lead to the formation of products. Thus, the reaction rate is increased.

The Maxwell–Boltzmann curve in Figure 17.4.4 shows the smaller number of particles with energies that exceed the activation energy in an uncatalysed reaction (shaded in red) compared with the number of particles with energies that exceed the activation energy in a catalysed reaction (sum of the regions shaded in red and blue).

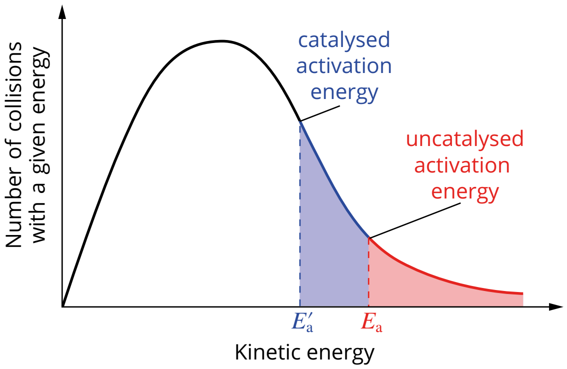


FIGURE 17.4.4 A catalyst provides a reaction pathway with a lower activation energy, increasing the proportion of collisions that exceed the activation energy and lead to a reaction.

Catalysts lower the activation energy by providing an alternative reaction pathway for the reaction. This increases the proportion of reactant particles with energies equal to or greater than the lower activation energy. This increases the proportion of successful collisions.

TYPES OF CATALYSTS

Depending on the physical state of the chemicals involved in the reaction and the catalyst, catalysts can be divided into two groups.

- **Homogeneous catalysts** are in the same physical state as the reactants and products of the reaction.
- **Heterogeneous catalysts** are in a different physical state from the reactants and products of the reaction.

An example of homogeneous **catalysis** occurs in the upper atmosphere and has contributed to the depletion of the ozone layer. Chlorine atoms in the gaseous state act as catalysts in the decomposition of ozone gas into oxygen gas. The chlorine atoms may have come from chlorofluorocarbons (CFCs) released into the atmosphere from refrigerators or air conditioners.

You may be familiar with the catalysed decomposition of a hydrogen peroxide solution using the black powder manganese(IV) oxide (MnO_2). This is an example of the use of a heterogeneous catalyst.

The chemical industry uses catalysts extensively. Chemists prefer to use heterogeneous catalysts for industrial processes because they are:

- more easily separated from the products of a reaction
- much easier to reuse
- able to be used at high temperatures.

Particles at the surface of some solids of high surface area tend to adsorb (form a bond with) gas molecules that strike the surface. **Adsorption** distorts bonds in the gas molecules or may even break them completely. This allows a reaction to proceed more easily than it would if the solid were absent.

These solid surfaces provide a new way for the reaction to occur (a new reaction pathway) that has a significantly lower activation energy.

A powdered or sponge-like form of a solid catalyst is often used to provide the greatest possible surface area. With a larger surface area, more reactant molecules can be adsorbed and the reaction is even faster.

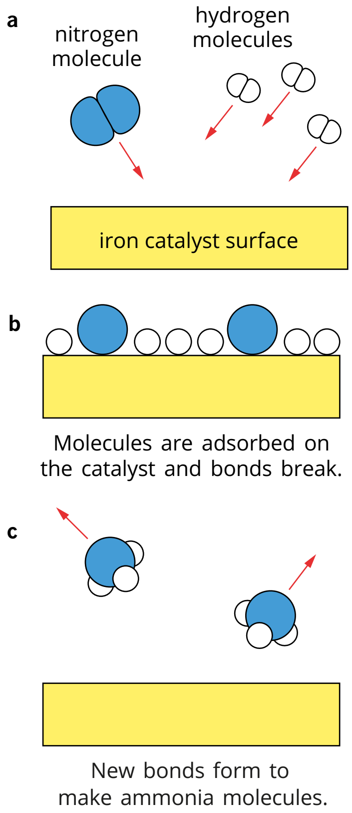


FIGURE 17.4.5 Behaviour of a catalyst in ammonia production. (a) Nitrogen and hydrogen molecules approach the iron catalyst surface. (b) The nitrogen and hydrogen molecules adsorb on the surface of the catalyst and their covalent bonds are broken. (c) The hydrogen and nitrogen atoms readily combine to form ammonia molecules. The molecules then leave the catalyst surface, and the catalyst remains unaltered by the reaction.

Heterogeneous catalysis and the Haber process

The Haber process is a very important commercial reaction that produces ammonia gas, which is used to make fertilisers, nylon, explosives and some pharmaceuticals. In the Haber process, hydrogen and nitrogen gas are converted to ammonia (NH_3), using a catalyst of powdered iron.

The reaction is represented by the equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \quad \Delta H = -91\ \text{kJ mol}^{-1}$$

Hydrogen and nitrogen molecules both adsorb onto the iron surface (see Figure 17.4.5a). As they attach themselves to the surface, the covalent bonds within their molecules break (Figure 17.4.5b).

The hydrogen and nitrogen atoms now readily combine to form ammonia molecules and move away from the iron surface (Figure 17.4.5c). The catalyst remains unaltered by the reaction.

Without a catalyst, temperatures over 3000°C are needed for the reaction to occur. The catalyst allows the manufacture of ammonia to proceed at an economical rate using a temperature of about 500°C.

The iron catalyst provides an alternative reaction pathway that dramatically reduces the activation energy barrier—the energy needed to break the covalent bonds in the nitrogen and hydrogen molecules. Even though collisions are less frequent at 500°C than at 3000°C, a greater proportion of colliding particles have sufficient energy to successfully react.

Catalytic converters in cars

All new cars sold in Australia have a catalyst fitted between the engine and the exhaust pipe. The purpose of the catalyst is to clean the exhaust gases from the engine and reduce the air pollution that could be caused if these gases entered the atmosphere.

Catalysts in cars convert carbon monoxide and nitrogen oxide, formed in the engine, to the non-toxic gases carbon dioxide and nitrogen.

$$2NO(g) \rightarrow N_2(g) + O_2(g)$$
$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

Unburnt hydrocarbons are also converted by the catalyst to carbon dioxide and water:

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

The catalyst is usually a mixture of platinum, palladium and rhodium metals and aluminium oxide, and it is mounted on a honeycomb-shaped support (Figure 17.4.6). Millions of tiny pores in the metals provide a large surface area.

Exhaust gases enter the catalyst chamber, pass quickly over the metals and leave the exhaust pipe in a purified condition. The catalyst is unchanged by the reaction and can be used without replacement for many years.

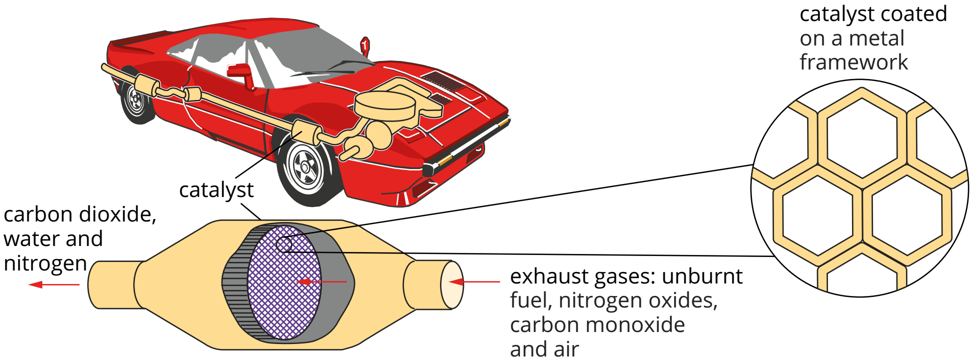


FIGURE 17.4.6 A catalyst fitted in the exhaust system of a car reduces air pollution.

ENZYME CATALYSIS

Enzymes are large biological molecules (typically proteins) that catalyse a wide range of biochemical processes in living organisms. They are found almost everywhere in our bodies and assist in everything from digesting our food to copying our DNA. Enzymes are known to catalyse over 5000 different types of chemical reactions and they are essential for pretty much all biochemical reactions in our cells.

The importance of enzymes is perhaps best understood when you see the results of not having them (or having them in less than optimal amounts). For example, many people find it difficult to digest dairy products and can become quite ill even digesting small amounts. This is known as lactose intolerance and is primarily due to a lack of the enzyme, lactase, which catalyses the breakdown of lactose in the small intestine to produce glucose and galactose (Figure 17.4.7).

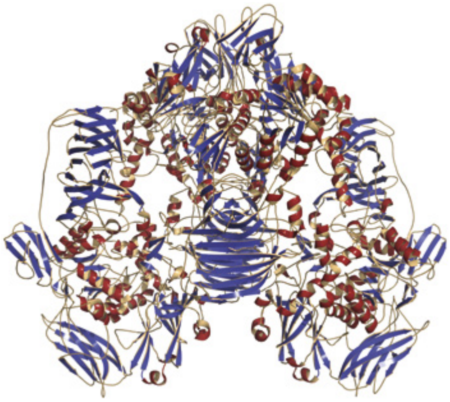


FIGURE 17.4.7 Computer-generated image of the large globular protein structure of lactase. Lactase is found in the small intestine and catalyses the breakdown of the disaccharide sugar, lactose, into the two monosaccharide sugars: glucose and galactose.

WS 2.3.4

Table 17.4.1 shows other examples of medical conditions caused by the lack of certain enzymes required to catalyse crucial metabolic reactions.

TABLE 17.4.1 Examples of medical conditions caused by a lack of catalysis enzymes

Disease	Defective enzyme or system	Symptoms	Treatment
phenylketonuria (PKU)	phenylalanine hydroxylase	severe mental retardation	screening; dietary modification
type 2 tyrosinemia	tyrosine aminotransferase	irritation to the corneas of the eyes	diet with reduced phenylalanine and tyrosine content
galactosemia	galactose-1-phosphate uridyl transferase	liver failure in infancy	newborn screening; milk avoidance
Gaucher disease (types I and II)	cerebrosidase	enlargement of the spleen and liver; painful and crippling effects on the bones; severe brain disease and death (type II)	enzyme replacement (type I)
Tay–Sachs disease	beta-hexosaminidase A	neurological disorders; enlarged head; death in early childhood	—
Fabry disease	α-galactosidase	severe pain; renal failure; heart failure	enzyme replacement
Hurler syndrome, Hunter syndrome	α-iduronidase (Hurler syndrome); iduronate sultatase (Hunter syndrome)	enlargement of the liver and spleen; skeletal deformities; coarse facial features; stiff joints; mental retardation; death within 5–15 years	enzyme replacement
Morquio syndrome	galactose 6-sulfatase; β-galactosidase	truncal dwarfism; severe skeletal deformities; potentially life-threatening susceptibility to cervical spine dislocation; valvular heart disease	—
albinism	tyrosinase	lack of skin hair and eye pigments	—

Like other catalysts, enzymes are not consumed in chemical reactions. Also, like other catalysts, they increase reaction rates by lowering the activation energy of chemical reactions. However, enzymes tend to be incredibly efficient at doing this, often increasing rates of reaction by millions or even billions of times. An example is the breakdown of hydrogen peroxide to oxygen and water:

$$2H_2O_2 \longrightarrow O_2 + 2H_2O$$

The activation energy for the uncatalysed breakdown of hydrogen peroxide to oxygen and water is $+76\,\text{kJ}\,\text{mol}^{-1}$. In the presence of the enzyme catalase, the activation energy is reduced to $+30\,\text{kJ}\,\text{mol}^{-1}$. This modest reduction in activation energy increases the rate of reaction by 10^8, enough to convert a reaction time that would take years to one that is measured in seconds.

Enzyme catalysis can involve both homogeneous catalysis (where the enzyme and its reacting molecules are freely soluble in solution) or heterogeneous catalysis where the enzyme molecule is embedded into the cell wall structure of living tissue (with the reactant molecules soluble in solution).

Mechanism of enzyme action

The major difference between enzymes and other types of catalysts is their specificity. Enzymes tend to act on specific molecules and therefore tend to catalyse a single chemical reaction or a small number of closely related reactions.

The molecules that enzymes act on are called **substrates**. Enzymes convert substrates into other molecules, forming the products of the chemical reaction. At the end of the reaction the free enzyme becomes available to bind to other substrate molecules to continue the catalytic reaction.

Figure 17.4.8 shows a schematic representation of the enzyme-catalysed breakdown of sucrose to produce glucose and fructose. You can see that the overall process involves the binding of the substrate molecule followed by formation of two intermediate complexes prior to the formation of products and regeneration of the free enzyme.

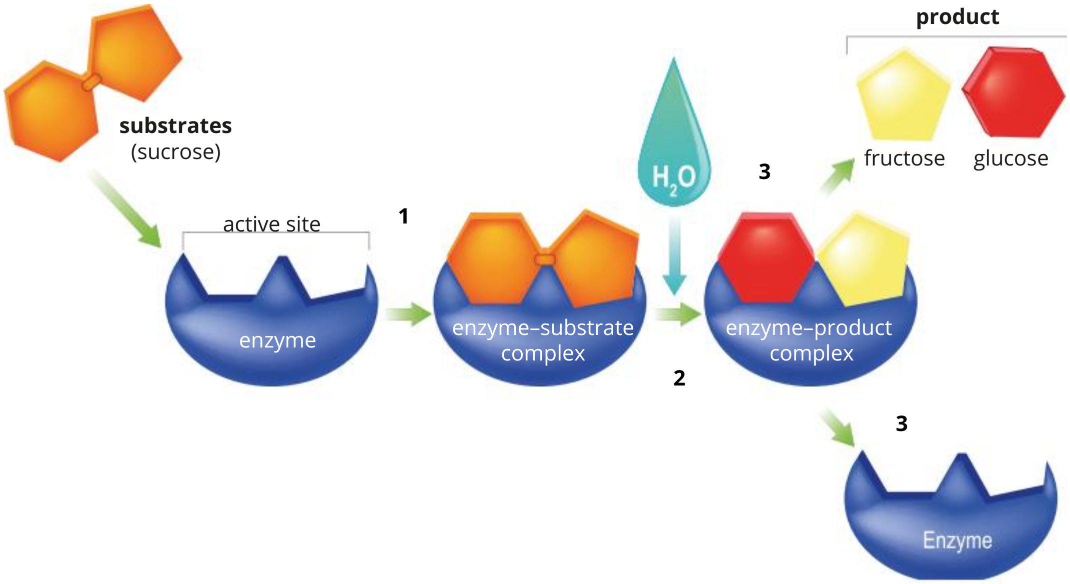


FIGURE 17.4.8 Schematic representation of the catalytic breakdown of sucrose by the enzyme sucrase, showing the three major steps of enzyme catalysis: 1) binding of the substrate within the active site of the enzyme to form the enzyme–substrate complex; 2) conversion of enzyme–substrate complex into the enzyme–product complex; 3) release of the products and regeneration of the free enzyme.

This process can also be represented using this generalised enzyme-catalysed reaction scheme:

$$E + S \rightarrow E\text{–}S \rightarrow E\text{–}P \rightarrow E + P$$

where E is the enzyme, S is the substrate molecule(s), P is the product molecule(s), E–S is the enzyme–substrate complex and E–P is the enzyme–product complex.

Active site binding

Figure 17.4.8 shows that enzymes need to bind to the substrate molecule before any catalysed reaction can take place. The binding site within the enzyme is a small section of the overall protein structure known as the **active site**. As stated above, enzymes are very specific in terms of the substrates they bind. This specificity arises from an active site having complementary size, shape and charge characteristics with the substrate molecule allowing for specificity of bonding. Every enzyme has a unique active site specific for a particular substrate molecule. It turns out that the types of intermolecular attractions you saw in Chapter 12 involving dipole–dipole interactions, such as hydrogen bonding, become very important in the initial enzyme–substrate binding process.

Lock-and-key model of enzymatic action

An explanation for the specificity of enzymes was first proposed by Emil Fischer in 1894 who suggested that a substrate molecule acts as a 'key' to the 'lock' of the active site. This **lock-and-key model** suggests that substrates and enzymes have specific complementary geometric shapes that fit together like pieces of a jigsaw puzzle, as represented in Figure 17.4.8. While this model explains the high specificity of enzyme-catalysed reactions, it fails to explain why the energy of the transition state is lowered compared to the uncatalysed reaction—in other words, it fails to explain why the activation energy of the overall reaction decreases in the presence of an enzyme.

Induced-fit model of enzymatic action

A modification to the lock-and-key model was proposed by Daniel Koshland in 1958. The **induced-fit model** does not assume a perfect geometric fit between the substrate and the active site; rather, it proposes that the substrate induces changes to the shape of the flexible enzyme structure and, in particular, induces changes to the shape of the active site. In this model, the active site is moulded into the precise shape required for the enzyme to carry out its catalytic function. In some cases, the shape of the substrate molecule may also change to allow optimal (lower energy) conditions for the reaction to proceed more rapidly. Figure 17.4.9 shows a pictorial representation of the induced-fit model.

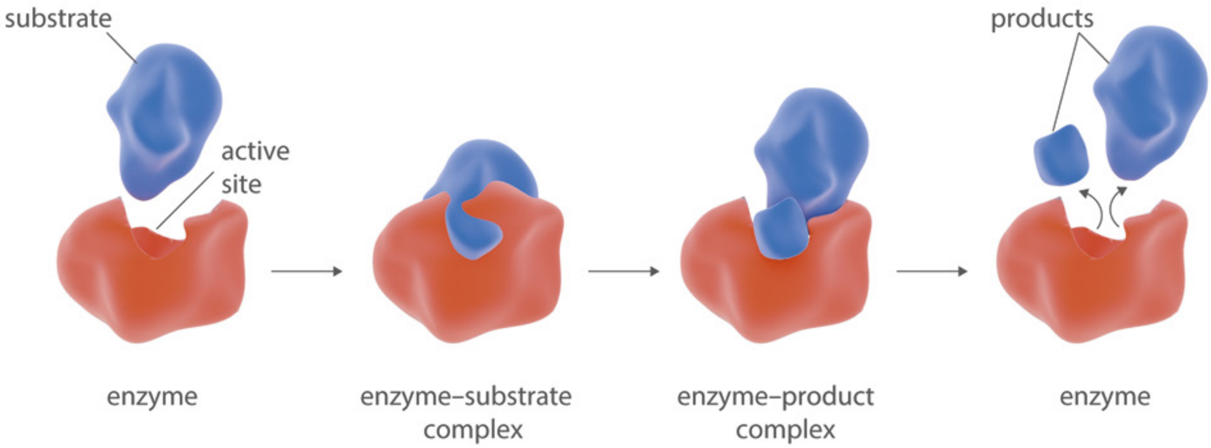


FIGURE 17.4.9 Induced-fit model of enzyme catalysis. Initially the shape of the active site is not an exact match for the substrate. The flexible enzyme structure is able to mould itself into the correct orientation and shape to allow the catalytic reaction to proceed.

The induced-fit model is better able to explain how enzymes can accelerate chemical reactions in one of several ways, all of which assist in lowering the activation energy.

- Enzymes distort the bound substrate molecule(s) to stretch and ultimately break the bonds required for the reaction to proceed.
- By better orientating the substrate molecule(s) into an optimal arrangement, enzymes contribute to a lowering of the energy required to stretch and break bonds.
- In some cases, the enzymes form temporary covalent bonds with the substrate molecule(s) to provide a lower energy transition state.
- The active site of the enzyme moulds itself to provide an energetically favourable environment to stabilise the transition state at a lower energy.

It is possible that enzymes may use several of these mechanisms simultaneously.

Energy profile diagrams of enzyme-catalysed reactions

Figure 17.4.10 shows a typical energy profile diagram for an enzyme-catalysed reaction. A characteristic feature of this diagram is the presence of multiple transition states (and multiple energy barriers) corresponding to the formation of the intermediate complexes and the products. The overall activation energy, E_a, is represented by the highest energy level of the three transition states and this is normally associated with the energy required to transition from the enzyme–substrate complex to the enzyme–product complex.

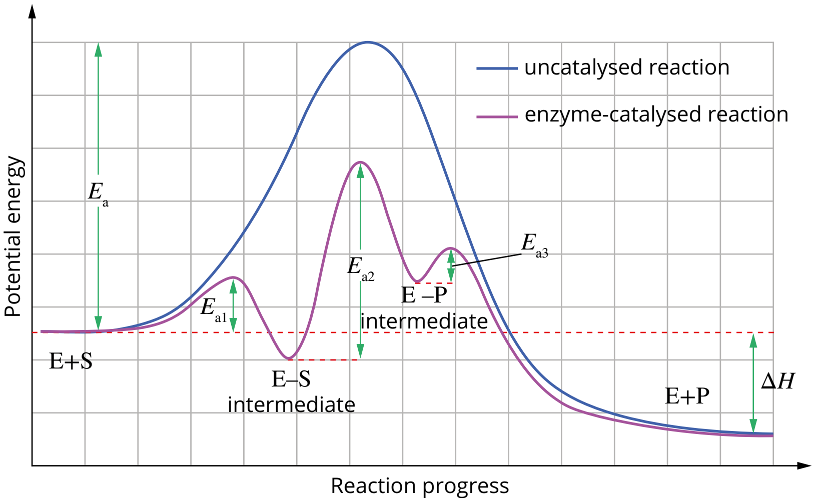


FIGURE 17.4.10 Energy profile diagrams for an uncatalysed reaction and a typical enzyme-catalysed reaction showing the relative energy levels of reactants, products and intermediate complexes as well as the relative energy levels of the various transition states. The overall activation energy for the enzyme-catalysed reaction (E_{a2}) is significantly less than the activation energy for the uncatalysed reaction (E_a).

As described above, the overall reaction scheme for a typical enzyme-catalysed reaction can be broken into three steps, each of which can be related to different parts of the energy profile diagram in Figure 17.4.10.

Step 1: E + S → E–S

Initial binding of the substrate with the active site to form the enzyme–substrate complex (E–S). A relatively small amount of energy is required for this step as shown by the small energy barrier E_{a1}. The energy of the E–S complex is lower than the reactants because it involves energy release, whether it be stabilisation via intermolecular forces or formation of temporary covalent bonds.

Step 2: E–S → E–P

Conversion of the enzyme–substrate complex (E–S) to the enzyme–product complex (E–P). This process involves bond stretching and bond breaking and so typically requires the highest energy of the three-step process. The relatively high energy required is shown by the energy barrier E_{a2}, which is still significantly lower than the overall energy required for the uncatalysed reaction (E_a).

Step 3: E–S → E + P

The breakdown of the enzyme–product complex (E–P) leading to the formation of products and regeneration of the free enzyme, thus completing the catalytic cycle.

WS 2.3.5

WS 2.3.3

SCIENCE AS A HUMAN ENDEAVOUR

From lock-and-key to induced fit—changing scientific ideas

Scientific theories, ideas and models change over time. Mostly they change as new observations and new evidence come to light that can better explain current observations or better predict future events. This was the case for the shift from the previously accepted lock-and-key model to the currently accepted induced-fit model for enzyme catalysis.

The development of the lock-and-key model and the induced-fit model of enzyme catalysis is a good example of how scientific theories, ideas and models change over time, as new evidence becomes available and as greater understanding develops.

Emil Hermann Fischer (Figure 17.4.11) was a remarkable chemist. Born in 1852 in the Cologne region of Germany, he attended the University of Bonn and the University of Strasbourg as an undergraduate student. He worked as a research chemist at several universities before taking on professorial roles at the University of Wurzburg and the University of Berlin, where he remained until his death in 1919.

He was awarded the Nobel Prize in Chemistry in 1902 for his work on identifying the 16 enantiomers (asymmetric forms of the same molecule) of the aldo-hexases, the family of sugars that glucose belongs to. His work on biomolecules and the precise arrangement of the molecules in the different forms of these biomolecules led him to propose the lock-and-key model of enzyme interaction.

This model was the accepted model for describing the mechanism of enzyme catalysis until 1958 when American Daniel Koshland (Figure 17.4.12) proposed the induced-fit model of enzyme interactions.

As with the evolution of most scientific models, Koshland was able to build on the work of others, including Fischer. In the years between Fischer's work and the work of Koshland, researchers developed a greater understanding of transition state chemistry, activation energy and the science of the energy of enzyme-catalysed reactions. Fischer's lock-and-key-model adequately explained that enzymes are specific—each can only bind to one or a very few kinds of substrates in enzyme-catalysed reactions; however, it was unable to explain the specifics of the energy changes of these reactions—an area of research that had grown considerably in the intervening years.

Koshland described his induced-fit model as more like a hand-in-a-glove rather than a lock-in-a-key. He proposed that the enzyme and its active site were flexible entities, not rigid, and could change shape on binding of the substrate. Figure 17.4.13 shows a simplified representation of Fischer's lock-and-key-model and Koshland's induced-fit model.

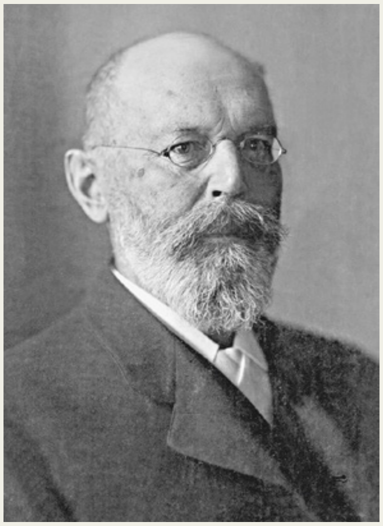


FIGURE 17.4.11 Emil Hermann Fischer (1852–1919) proposed the lock-and-key model of enzyme–substrate interactions.

FIGURE 17.4.12 Daniel Koshland (1920–2001) further developed Fischer's lock-and-key model by proposing the revised induced-fit model of enzyme–substrate interactions.

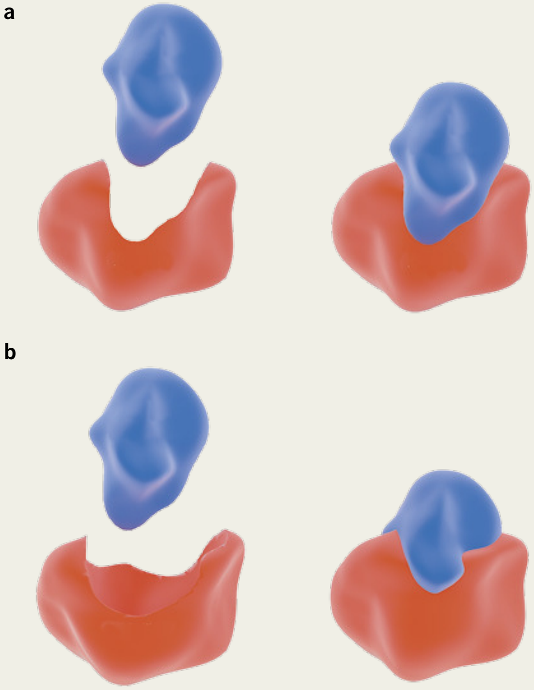


FIGURE 17.4.13 The two models of enzyme interaction proposed by (a) Fischer and (b) Koshland. The development of these models is an example of how scientific ideas change over time, being refined and adapted as new insights and discoveries are made.

Review

1 Judge what you think was the compelling piece of scientific observation or evidence that led to the acceptance of the induced-fit model over the lock-and-key model.

2 New observations, new evidence and better understanding of scientific concepts are reasons for the evolution of scientific ideas. Explain reasons why scientific ideas might persist in the face of new observations, evidence or understandings that would otherwise overturn the existing ideas.

17.4 Review

SUMMARY

- The rate of a reaction can be increased by using a catalyst.
- A catalyst provides an alternative reaction pathway that has a lower activation energy.
- Energy profile diagrams (Figure 17.4.14), which can include catalysed and uncatalysed pathways, may be used to represent the enthalpy changes and activation energy associated with a chemical reaction.

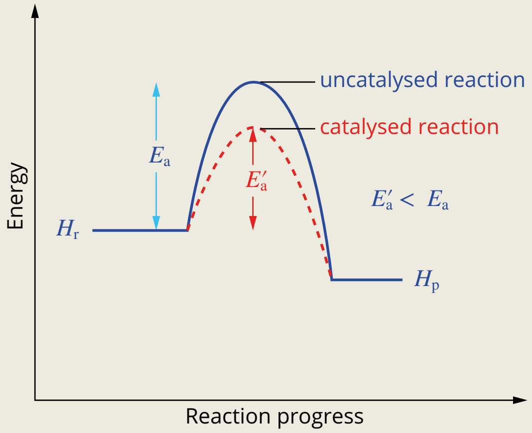


FIGURE 17.4.14 Energy profile diagram of a reaction showing the effect of a catalyst

- A catalyst provides a new reaction pathway and it is not used up in the reaction.
- When a catalyst is present, a greater proportion of the collisions between particles exceeds the activation energy barrier of the reaction and therefore leads to a more rapid chemical change.
- Enzymes are macromolecules that catalyse a wide range of biochemical reactions in living organisms.
- A key feature of enzyme-catalysed reactions is their specificity. Enzymes tend to act on specific molecules and therefore tend to catalyse a single chemical reaction or a small number of closely related reactions.
- Enzyme-catalysed reactions involve the formation of two intermediate complexes (the enzyme–substrate complex and the enzyme–product complex), which serve to provide an alternative reaction pathway with a decreased activation energy.
- Two models of enzyme action have been proposed: the lock-and-key model and the induced-fit model. The lock-and-key model can explain the specificity of enzyme action but not the reason for a decrease in activation energy. The induced-fit model, a modified form of the lock-and-key model, is able to explain both the specificity of enzyme reactions and the decrease in activation energy.

KEY QUESTIONS

Retrieval

1. Define the term 'catalyst'.
2. State what an enzyme is.
3. List the three steps in an enzyme-catalysed reaction.
4. List the four ways in which an enzyme can assist in lowering the activation energy of a chemical reaction.

Comprehension

5. In terms of the energy of colliding reactant particles, explain how a catalyst increases the rate of a chemical reaction.
6. **a** Explain why surface properties are important to the operation of catalysts.
   b Many industrial catalysts are made into porous pellets. Determine the reason for this.

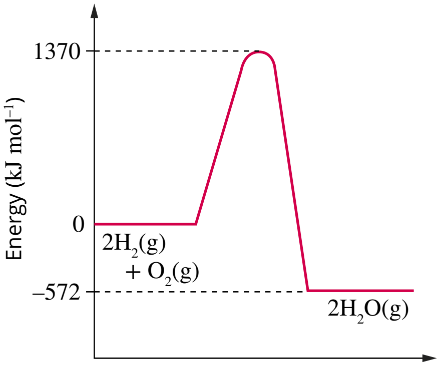
Analysis

7. The rate of a chemical reaction can be increased by increasing temperature and/or by adding a catalyst. With reference to collision theory, compare and contrast these two approaches to increasing reaction rates.
8. If a sugar cube is held in the flame of a candle, the sugar melts and browns but does not burn. However, the cube burns if salt is first rubbed into it, even though the salt does not react. Determine the effect of the salt on the activation energy of this combustion reaction.

9 Many major car makers have plans for hydrogen-powered cars. In the fuel cells of these cars, hydrogen reacts with oxygen from the air to produce water.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

Energy changes for the reaction are shown in the energy profile graph below.



a Determine the magnitude of the activation energy of this reaction.

b Calculate ΔH for this reaction.

Several groups of scientists have claimed to have split water into hydrogen and oxygen using a molybdenum catalyst:

$$2H_2O(g) \xrightarrow{\text{Mo}} 2H_2(g) + O_2(g)$$

c Sketch energy change graphs for this reaction with and without the presence of a catalyst.

d Determine the value of ΔH for this water-splitting equation.

MANDATORY PRACTICAL 7

Factors affecting the rate of reaction

Research and planning

Aim

To investigate the ways in which changes in temperature and concentration of reactants affect the rate of a reaction

Rationale (scientific background to the experiment)

Hydrochloric acid reacts with sodium thiosulfate to produce sulfur, sulfur dioxide and water in the following reaction:

$$S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(g) + H_2O(l)$$

The sulfur precipitate causes the reaction mixture to become cloudy. The rate at which the solution becomes cloudy indicates the rate of reaction: the shorter the period before the cloudiness obscures a cross on a sheet of paper placed under the beaker, the more rapid the rate of reaction. In this practical activity, the effects of temperature and concentration of one of the reactants on the rate of reaction are studied. Therefore, students should see an inverse relationship between the rate of reaction and time taken for the cross to disappear.

Timing

50 minutes

Materials

- 40 mL 2 M hydrochloric acid
- 100 mL 0.25 M sodium thiosulfate solution
- 10 mL measuring cylinder
- 100 mL measuring cylinder
- 6 × 100 mL beakers
- thermometer, –10 to 110°C
- Bunsen burner, tripod stand and gauze mat
- bench mat
- stopwatch
- black marking pen
- safety glasses

As directed by your teacher, complete the risk assessment and management table by referring to the hazard labels on the reagent bottles or material safety data sheets (MSDS) or your teacher's risk assessment for the activity.

Method

Risk assessment

Consideration of risks includes chemical and physical risks. Before you commence this practical activity, you must conduct a risk assessment. Complete the template in your Skills and Assessment book or download it from your eBook.

PART A: EFFECT OF TEMPERATURE

1. Mark a cross on a sheet of paper with a black marking pen.
2. Place a 100 mL beaker on top of the cross. Pour 10 mL of 0.25 M sodium thiosulfate solution and 35 mL of water into the beaker.
3. Record the temperature of the solution in Table 17.5.1.
4. Add 5 mL of 2 M hydrochloric acid and commence timing. Measure and record in the table the time taken for the cross to disappear when viewed from above the beaker.
5. Pour 10 mL of 0.25 M sodium thiosulfate solution and 35 mL of water into another beaker. Heat the solution to approximately 40°C. Remove the beaker from the source of heat, place it on top of the cross, and immediately add 5 mL of hydrochloric acid. Again, measure and record the time taken for the cross to disappear.
6. Repeat step **5** with a third, fourth and fifth beaker, heating the solutions to approximately 50°C, 60°C and 70°C respectively.
7. Share your results with the class, to enable an average time to be determined for each temperature used.

PART B: EFFECT OF CONCENTRATION OF REACTANTS

1. Place a 100 mL beaker on top of a cross marked on a sheet of paper. Pour 10 mL of 0.25 M sodium thiosulfate solution and 35 mL of water into the beaker. Add 5 mL of 2 M hydrochloric acid and commence timing. Measure the time taken for the cross to disappear when viewed from above the beaker, and record this in Table 17.5.2.

PRE-LAB SAFETY INFORMATION

Material used	Hazard	Control
2 M HCl	toxic by all routes of exposure; lung irritation	Wear eye and skin protection.
Sulfur dioxide gas	toxic by inhalation; causes burns; risk of serious damage to the eyes	Avoid breathing gas; work in a well-ventilated area and dispose of reaction mixtures using a sink in a fume cupboard. Wear safety glasses.

Please indicate that you have understood the information in the safety table.
Name (print): ____________________
I understand the safety information (signature): ____________________

2 Place another 100 mL beaker on top of the cross. Pour 15 mL of the sodium thiosulfate solution and 30 mL of water into the beaker. Add 5 mL of hydrochloric acid and commence timing. Measure the time taken for the cross to disappear and record this in Table 17.5.2.

3 Place another 100 mL beaker on top of the cross. Pour 25 mL of the sodium thiosulfate solution and 20 mL of water into the beaker. Add 5 mL of hydrochloric acid and commence timing. Measure the time taken for the cross to disappear and record this in Table 17.5.2.

4 Place another 100 mL beaker on top of the cross. Pour 35 mL of the sodium thiosulfate solution and 10 mL of water into the beaker. Add 5 mL of hydrochloric acid and commence timing. Measure the time taken for the cross to disappear and record this in Table 17.5.2.

5 Place another 100 mL beaker on top of the cross. Pour 40 mL of the sodium thiosulfate solution and 5 mL of water into the beaker. Add 5 mL of hydrochloric acid and commence timing. Measure the time taken for the cross to disappear and record this in Table 17.5.2.

6 Share your results with the class to enable an average time to be calculated for the different concentrations of reactants used.

Variables

PART A: EFFECT OF TEMPERATURE

i independent: the temperature of the solution

ii dependent: the rate of reaction as determined by the time taken for the cross to disappear

iii controlled: the volume and concentration of reactants, the size of the cross, the distance of viewing above the beaker, the person looking for the cross

PART B: EFFECT OF CONCENTRATION OF REACTANTS

i independent: the concentration of reactants by varying the volume of each used

ii dependent: the rate of reaction as determined by the time taken for the cross to disappear

iii controlled: the temperature, the size of the cross, the distance of viewing above the beaker, the person looking for the cross

Analysing

Raw data

1 Record your data for each experiment.

TABLE 17.5.1 Experimental results for effect of temperature

Temperature (°C)	**Time for the cross to disappear (s)**

TABLE 17.5.2 Experimental results for effect of concentration of reactants

Solution	**Concentration of sodium thiosulfate in solution (M)**	**Time for the cross to disappear (s)**
10 mL $Na_2S_2O_3$ solution, 35 mL water and 5 mL HCl		
15 mL $Na_2S_2O_3$ solution, 30 mL water and 5 mL HCl		
25 mL $Na_2S_2O_3$ solution, 20 mL water and 5 mL HCl		
35 mL $Na_2S_2O_3$ solution, 10 mL water and 5 mL HCl		
40 mL $Na_2S_2O_3$ solution, 5 mL water and 5 mL HCl		

Processed data

2 Calculate the concentrations of the sodium thiosulfate in each of the reaction mixtures made up in Part B and enter those in the second column of Table 17.5.2.

3 Calculate the average time for the cross to disappear using all the class results for each temperature and each concentration. The class results tables should look similar to your results tables; however, the columns should read ‘Average time for the cross to disappear (s)’.

Reflect and check that your data analysis demonstrates these characteristics

☐ Effective investigation of phenomena is demonstrated by the collection of sufficient and relevant raw data.
☐ Accurate application of algorithms, visual and graphical representations of data is demonstrated by appropriate processing and presentation of data to aid the analysis and interpretation of data.

Analysis

4 Identify any differences between your group results and the class results for either of the tests. Deduce why this might have occurred.

5 After analysing the results, infer what effect an increase in temperature has on the rate of this reaction.

6 Conclude why an increase in temperature leads to a change in the rate of reaction.

7 After analysing the results, judge what effect increasing the concentration of sodium thiosulfate has on the rate of this reaction.

8 Explain why the rate of a reaction depends upon the concentration of the reactants.

Reflect and check that your analysis demonstrates these characteristics

- ☐ Systematic and effective analysis of evidence is demonstrated by a thorough and appropriate error analysis.
- ☐ Systematic and effective analysis of evidence is demonstrated by a thorough identification of relevant trends, patterns and relationships.
- ☐ Insightful and valid interpretation of evidence is demonstrated by drawing a valid and defensible conclusion based on the analysis.

Interpreting and communicating

Conclusion

1 **a** State your conclusions. Ensure that you summarise the effects that a change in temperature and in the concentration of the reactants have on the rate of reaction.

b Identify the evidence that you have collected that leads you to make these conclusions.

Evaluation

2 Assess your analysis and conclusions. Judge whether the experiment provided an effective and efficient method of investigating the ways in which the changes in temperature and concentration of reactants affect the rate of a reaction.

Improvements

3 If you were to repeat the experiment, identify the steps that you would do differently. You should include these points in your answer.

a Explain how you would change the methodology and how this might improve the results.

b Consider how well you performed the tasks and the skills that you need to improve on in your technique.

c Explain how the collection of data could be improved or uncertainty reduced.

Reflect and check that your evaluation demonstrates these characteristics

- ☐ Critical evaluation of processes is demonstrated by a discussion of the reliability and validity of the experimental process supported by evidence such as the quality of the data (as quantified in the error analysis).
- ☐ Critical evaluation of the conclusion is demonstrated by a discussion of the veracity of the conclusions with respect to the error analysis and limitations or sufficiency of the data.
- ☐ Insightful evaluation of processes and conclusions is demonstrated by a suggestion of improvements or extensions to the experiment that are logically derived from the analysis of the evidence.

Chapter review

WS 2.3.6

WS 2.3.7

17

KEY TERMS

acid rain
activation energy
active site
adsorption
average rate of reaction
catalysis
catalyst
chemical reaction
collision theory
concentration
concentration–time curve
endothermic
energy profile diagram
enzyme
exothermic
heterogeneous catalyst
homogeneous catalyst
induced-fit model
initial rate of reaction
instantaneous rate of reaction
kinetic energy
kinetic energy distribution diagram
lock-and-key model
Maxwell–Boltzmann distribution curve
photochemical smog
potential energy
pressure
rate of reaction
reactant
reaction pathway
substrate
surface area
transition state

KEY QUESTIONS

Retrieval

1 Identify which one of the following is the correct definition of rate of reaction.

A the time it takes for all of a reactant to be used up
B how fast a reaction is going at the end of 1 minute
C how much a reaction is bubbling
D the change in concentration of reactants or products over time

2 Identify which of the following is the correct unit for measuring the rate of a reaction.

A $mol^{-1}\,L\,s^{-1}$
B $mol\,L^{-1}\,s^{-1}$
C $mol^{-1}\,L^{-1}\,s$
D $mol\,L\,s^{-1}$

3 Identify which of the following changes would decrease the rate of the reaction between zinc metal and dilute hydrochloric acid.

A increasing the temperature of the hydrochloric acid
B decreasing the size of the pieces of zinc
C decreasing the concentration of the hydrochloric acid
D decreasing the volume of hydrochloric acid used

4 According to collision theory, select which one of the following is *not* essential for a reaction to occur.

A Molecules must collide to react.
B The reactant particles should collide with the correct orientation.
C The reactant particles should collide with enough energy to overcome the activation energy barrier.
D The reactant particles should collide with double the energy of the activation energy.

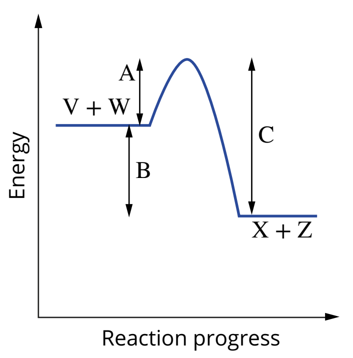
5 According to collision theory, state what must happen for a reaction to occur.

Comprehension

6 Consider the reaction between solutions V and W that produces X and Z according to the equation:

$$V(aq) + W(aq) \rightarrow X(aq) + Z(aq)$$

The energy profile diagram for this process is shown below:



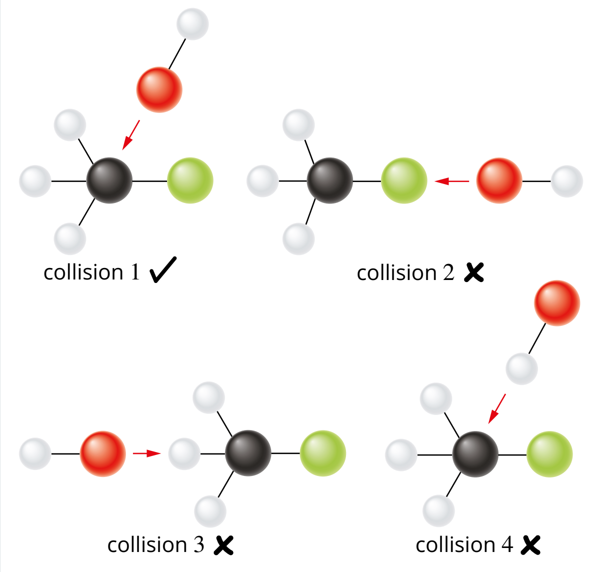
Determine which one of the following alternatives describes the change that a catalyst produces to increase the reaction rate.

A Only B is decreased.
B Only A is decreased.
C A, B and C are decreased.
D A and C only are decreased.

7 The following changes are made to a reaction mixture. Determine which one of the following changes will lead to a decrease in reaction rate.

A Smaller solid particles are used.
B The temperature is decreased.
C A catalyst is added.
D The concentration of an aqueous reactant is increased.

8 The schematic diagram below depicts the substitution reaction between chloromethane and a hydroxide ion. Using collision theory, explain why collision 1 might be successful, while collisions 2–4 will not be successful.



9 Explain why the rate of reaction between 1 M $CuSO_4$ and powdered zinc is greater than that with an equal mass of large zinc pieces.

10 Summarise what must occur when reactant particles collide and react.

Analysis

11 Compare and contrast the average rate of reaction with the instantaneous rate of reaction. Identify how the instantaneous rate and initial rate are similar.

12 Calcium carbonate ($CaCO_3$) chips and 2 M HCl were the starting reactants for the chemical reaction represented by the following chemical equation:

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

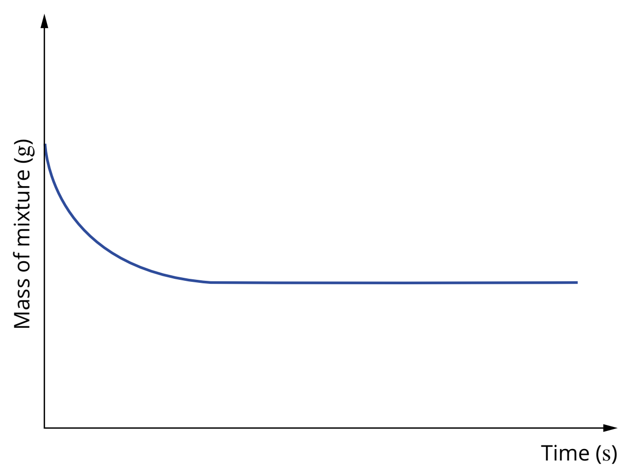
Predict what will happen to the rate of reaction (and give your reasons) if the starting reactants were changed in each case.

- **a** $CaCO_3$ chips and 1 M HCl
- **b** $CaCO_3$ powder and 2 M HCl
- **c** $CaCO_3$ powder and 1 M HCl

13 A 5.00 g piece of copper was dissolved in a beaker containing 500 mL of 2.00 M nitric acid. The equation for the reaction that occurred is:

$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$

The changing mass of the mixture was observed for a period of time, and the graph is shown below:



- **a** Identify why the mass of the mixture initially decreases with increasing time.
- **b** Based on the information provided, determine which reactant is limiting.
- **c** Redraw the above graph, then sketch in the expected curve if 500 mL of 1.00 M nitric acid had been used instead. Label your new graph line. Identify why there is a difference in shape.
- **d** Redraw the above graph, then sketch in the expected curve if 5.00 g of powdered copper had been used instead. Label this new graph line. Identify why there is a difference in shape.

14 A reaction is represented by the following chemical equation:

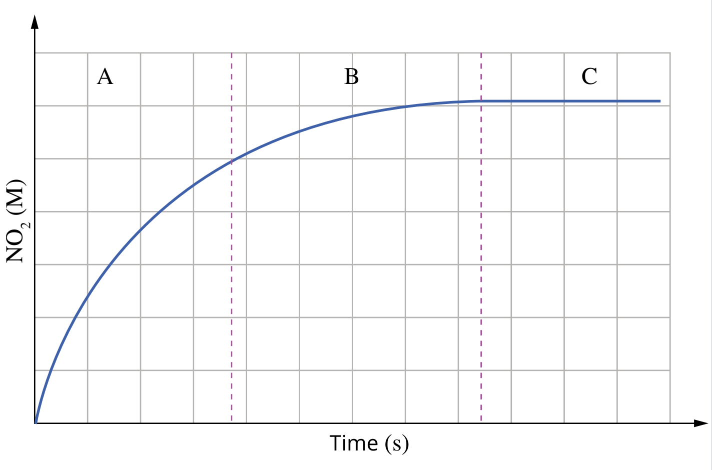
$$3A + B \rightarrow 2C$$

- **a** Create rate expressions for the rate of disappearance of A, the rate of disappearance of B and the rate of appearance of C.
- **b** Using the rate expressions you wrote down in part **a**, if you calculated the average rates of appearance and disappearance over the same time period, predict whether the rates be equal. If not, rank the rates in order from fastest to slowest.
- **c** If your answer to part b was no, create rate expressions for all three substances that would be equal.

15 The production of nitrogen dioxide, NO_2, from nitric oxide, NO, can be represented by the following chemical equation:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

A concentration–time curve for the production of NO_2 is shown below:



a Deduce which region of the curve shows a constant value for the rate of reaction—A, B or C. Justify your answer.

b Judge which region of the curve indicates the fastest rate—A, B or C. Justify your answer.

16 The following data was obtained for the decomposition of hydrogen peroxide, H_2O_2:

Time (s)	**$[H_2O_2]$ (M)**
0.0	0.600
10.0	0.300
20.0	0.150
30.0	0.0750

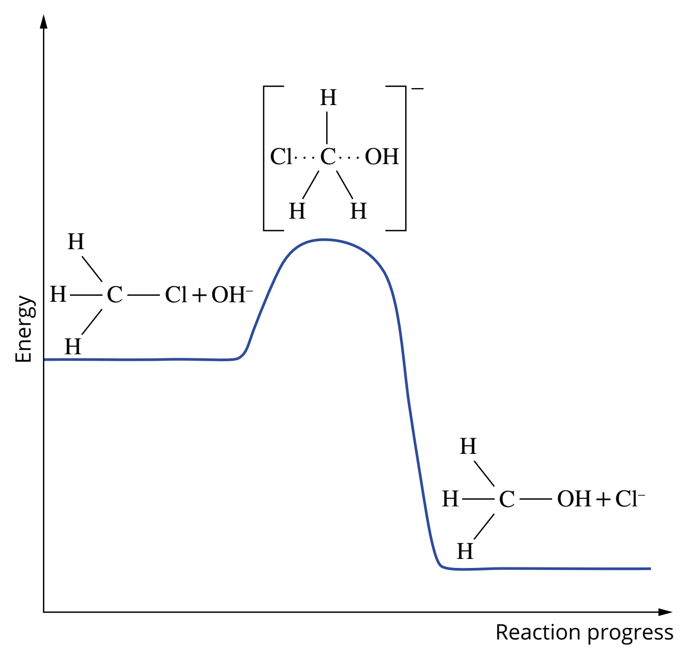
The balanced chemical equation for this chemical reaction is: $2H_2O_2 \rightarrow 2H_2O + O_2$

a Calculate the average rate of disappearance of H_2O_2 between 10.0 s and 20.0 s

b Calculate the average rate of reaction between 10.0 s and 20.0 s.

c Compare and contrast these measures of rate.

17 The figure below is an energy profile diagram for the substitution reaction between chloromethane and hydroxide ion.



a Copy and analyse this diagram. Predict then label on the diagram the positions of ΔH and the activation energy.

b Explain what is meant by the term 'activation energy'.

c Predict whether the reaction is endothermic or exothermic.

d Deduce where the transition state occurs and label this position on your diagram.

e Identify the bonds that are beginning to be broken and formed to produce the transition state.

18 Compare and contrast homogeneous catalysts and heterogeneous catalysts. Identify if enzymes are homogeneous or heterogeneous catalysts.

19 The decomposition of hydrogen peroxide can be represented by the following chemical equation:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

For each of the following, predict whether the rate of reaction will increase or decrease. Provide an explanation in each case.

a decreasing the concentration of hydrogen peroxide

b increasing the temperature of hydrogen peroxide

c adding a potassium iodide catalyst

20 **a** Identify the five factors that influence the rate of a reaction.

b Classify the five factors from part **a** according to whether they increase the proportion of successful collisions by:

i increasing collision frequency

ii increasing the proportion of collisions that have energy equal to or greater than the activation energy

21 The Haber process involves the reaction of nitrogen gas and hydrogen gas to make ammonia gas. Identify two ways the rate of this reaction could be increased, at constant temperature. Using collision theory, explain why the rate is increased in each case.

22 Lumps of limestone, calcium carbonate, react readily with dilute hydrochloric acid. Four large lumps of limestone, total mass 10.0 g, were reacted with 100 mL 0.100 M acid.

a Create a balanced equation to represent the reaction.

b Identify which reactant is in excess. Use a calculation to support your answer.

c Describe a technique that you could use in a school laboratory to measure the rate of the reaction.

d A 10.0 g sample of small lumps of limestone will react at a different rate from four large lumps. Predict if the rate of reaction with the smaller lumps will be faster or slower. Explain your answer in terms of collision theory.

e Predict two other ways in which the rate of this reaction can be altered. Explain your answer in terms of collision theory.

Knowledge utilisation

23 Hydrogen reacts explosively with oxygen to form water.

a Decide what chemical bonds are broken in the reaction.

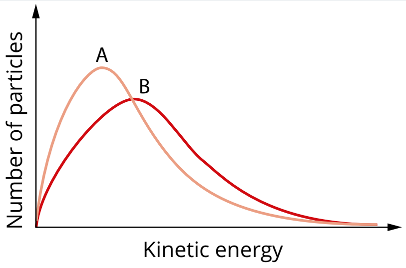
b Decide what chemical bonds are formed.

c Develop an explanation for how the energy changes during bond breaking and bond forming affect the overall energy change for the reaction.

d Judge why there is no reaction until the reaction mixture is ignited.

24 The figure shows the distribution of energies of particles in a substance at two different temperatures, 40°C and 60°C.

a Indicate the temperatures represented by graphs A and B.

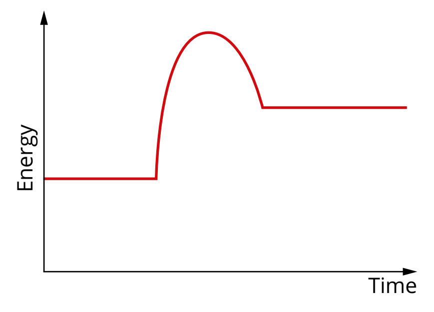


b Copy this diagram for temperature B and use it to develop an explanation to show the effect of a catalyst on a reaction.

c Use the diagram you have drawn in part **b** to predict in terms of collision theory how a catalyst increases the rate of a reaction.

25 The figure below shows the energy profile diagram for the reaction of hydrogen and iodine to form hydrogen iodide:

$$H_2(g) + I_2(s) \rightarrow 2HI(g)$$



a Copy the diagram and label it using the following: $H_2(g) + I_2(s)$; HI(g); ΔH; activation energy.

b Judge whether the reaction is endothermic or exothermic.

c Adapt the diagram to reflect the use of a catalyst in the reaction.

26 The reaction between sodium thiosulfate ($Na_2S_2O_3$) and hydrochloric acid (HCl) results in the formation of elemental sulfur (S), sulfur dioxide (SO_2), sodium chloride (NaCl) and water (H_2O).

a Develop a balanced ionic equation for this reaction.

b Predict one way in which, experimentally, the course of this reaction can be monitored by over time.

c Predict the effect an increase in temperature would have on the rate of reaction.

d Predict the effect an increase in HCl concentration would have on the rate of the reaction.

UNIT 2 • REVIEW

REVIEW QUESTIONS

Molecular interactions and reactions

TR 2.1 TR 2.2 TR 2.3

Topic 1: Intermolecular forces and gases

Multiple-choice questions

1 State which of the following gives the correct shape for each of the molecules listed.

	Linear	Bent	Tetrahedral
A	CO_2	H_2S	CH_4
B	H_2	CO_2	NH_3
C	HF	H_2O	NH_3
D	H_2O	NH_3	CH_4

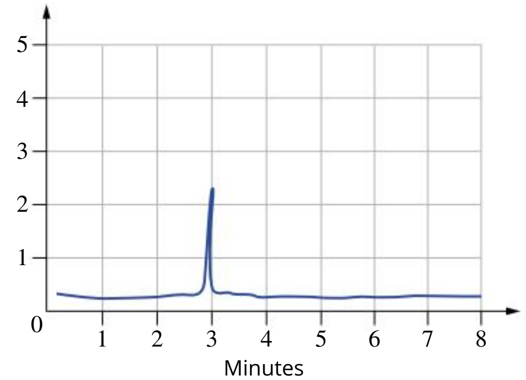
2 Identify which of the following groups contains only polar molecules.

A NH_3, H_2S, HCl
B CO_2, CH_4, H_2O
C HF, O_2, H_2
D H_2O, NH_3, CH_4

3 Identify which molecules have dispersion forces as the only intermolecular attraction.

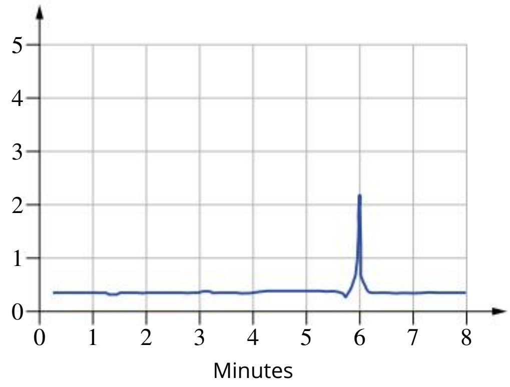
A HF
B OF_2
C NF_3
D CF_4

4 The HPLC chromatogram of a solution containing 4 ppm wolfram, a pesticide, is shown below.

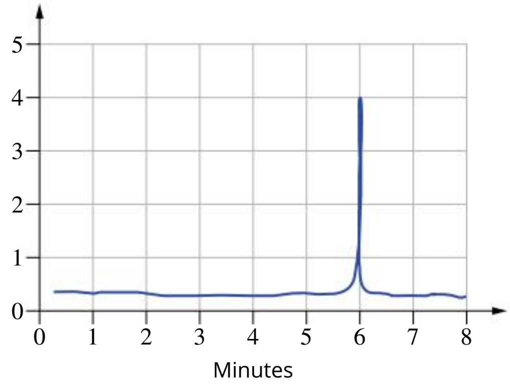


Identify which one of the following graphs correctly represents the chromatogram of an 8 ppm wolfram solution. Assume that all the chromatograms were obtained under identical conditions.

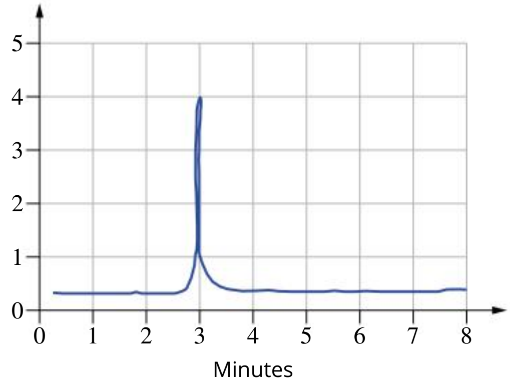
A



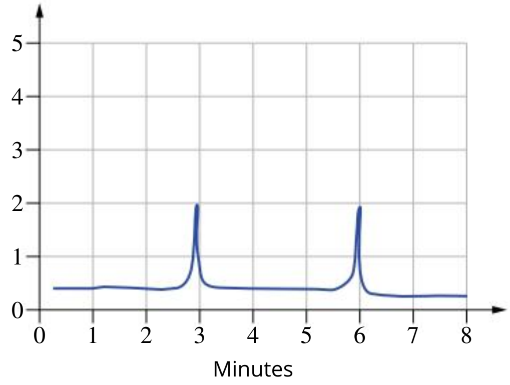
B



C



D



5 Identify which two intermolecular forces dominate in the bonding between molecules of HBr.

A hydrogen bonding and dispersion forces
B hydrogen bonding and ion–dipole attractions
C dipole–dipole attractions and dispersion forces
D dipole–dipole attractions and hydrogen bonding

6 Identify which of the following correctly represents SO_2, N_2, C_3H_8 and C_2H_5OH in order of increasing melting point.

A N_2, C_3H_8, SO_2, C_2H_5OH
B N_2, SO_2, C_3H_8, C_2H_5OH
C N_2, SO_2, C_2H_5OH, C_3H_8
D C_3H_8, N_2, C_2H_5OH, SO_2

7 Identify in which of these compounds you would expect hydrogen bonding to play a significant role in determining the boiling point.

$$H_2O,\ C_2H_6,\ CH_3F,\ C_2H_5OH,\ CH_3NH_2$$

A H_2O only
B H_2O, CH_3F, and C_2H_5OH only
C H_2O, C_2H_5OH and CH_3NH_2 only
D all except C_2H_6

8 Identify which one of the following classes of contaminants found in water can be quantitatively determined using HPLC.

A dioxins
B insecticides
C pesticides
D all of the above

9 A gas is confined in a cylinder with a movable piston at one end. Assuming the temperature remains constant, if the piston is pushed in, identify which of the following statements about the volume and pressure of the gas is correct.

A The pressure will remain unchanged.
B The pressure will increase because the volume is increasing.
C The pressure will decrease due to the increase in volume.
D The pressure will increase because of the decrease in volume.

10 A flask contains equal masses of two gases, hydrogen and nitrogen, at a temperature of 30°C. Determine which of the following statements about these gases is NOT correct.

A The average kinetic energy $\left(\frac{1}{2}mv^2\right)$ of the molecules of both gases is the same.
B The average velocity of the H_2 molecules is greater than the average velocity of the N_2 molecules.
C There are more H_2 molecules than N_2 molecules in the flask.
D The pressure exerted by the H_2 molecules in the flask is less than the pressure exerted by the N_2 molecules.

11 Determine which of the following gases would occupy the largest volume.

A 0.10 mol of CO_2 at 50°C and 1.0 kPa pressure
B 0.10 mol of H_2 at 40°C and 1.0 kPa pressure
C 0.20 mol of CH_4 at 50°C and 1.0 kPa pressure
D 0.10 mol of O_2 at 50°C and 0.40 kPa pressure

12 A 1.03 kg mass of a gas occupies 520 L at 155°C and 270 kPa pressure. Determine the approximate molar mass of the gas in $g\,mol^{-1}$.

A 9
B 26
C 32
D 44

13 Pentane undergoes complete combustion to produce carbon dioxide and water according to the equation:

$$C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$$

Determine the volume of carbon dioxide, in L, generated by the reaction of 100 g of pentane at STP.

The molar mass of pentane is 72 $g\,mol^{-1}$.

A $\frac{100 \times 24.5}{5 \times 72}$
B $\frac{72 \times 3}{24.5 \times 100}$
C $\frac{100 \times 5 \times 24.5}{72}$
D $\frac{100 \times 72}{5 \times 24.5}$

14 20.0 mL of NH_3 and 20.0 mL of O_2 measured at STP react. This reaction may be represented by the following equation:

$$4NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(l)$$

After the reaction, gas volumes are again measured at STP. Calculate the volume of oxygen gas that remains unreacted.

A 0.0 mL
B 5.0 mL
C 15.0 mL
D 20.0 mL

Short-answer questions

1 Consider the compounds carbon dioxide, CO_2, and hydrogen sulfide, H_2S.

Predict whether or not these compounds are polar or non-polar. Justify your prediction in terms of the nature of bonding within each molecule and the shape of each molecule. Illustrate your answer by using appropriate diagrams.

2 The grid below includes formulas for 12 different substances.

hydrogen sulfide H_2S	graphite C	fluorine F_2
carbon dioxide CO_2	methanol CH_3OH	nitrogen N_2
methane CH_4	ammonia NH_3	chloromethane CH_3Cl
hydrogen fluoride HF	oxygen O_2	ethene C_2H_4

Identify substances from this grid that meet the following descriptions.

Note that:

- the same substance may be used more than once
- a complete answer may require the inclusion of more than one substance.

a with non-polar tetrahedral molecules

b with polar tetrahedral molecules

c with non-polar linear molecules

d with one or more atoms covalently bonded to exactly three other atoms

e molecules are attracted to one another by dispersion forces only

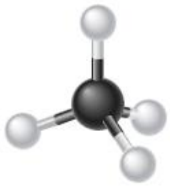
f have intermolecular hydrogen bonds

g have molecules of the same shape as water molecules

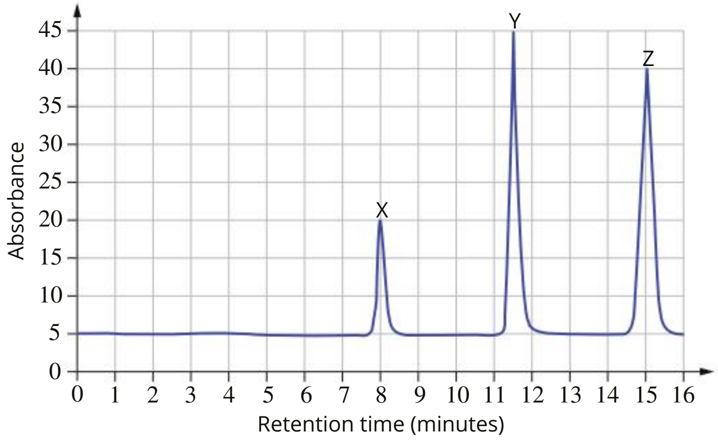
3 **a** Describe the intermolecular forces between pure samples of F_2, HF and CH_3F.

b Sort these substances from lowest to highest in the relative order of boiling points. Explain your answer.

4 Ball-and-stick models of some molecules are shown in the following table. Complete the table by identifying the shape of each molecule, stating the approximate bond angle, and providing an example of a substance whose molecules have the shape indicated.

Ball-and-stick model	Shape	Bond angle	Example
			
			
			
			
			

5 A sample of pond water was analysed for organic compounds X, Y and Z using HPLC. The following chromatogram was obtained.



a Recall some of the organic compounds that can find their way into waterways.

b **i** Identify which of the three compounds travelled through the HPLC column fastest. Explain your choice.

ii Identify which compound is present in the largest quantity. Explain your choice.

c The concentration of compound X needs to be determined.

i Explain why a calibration curve is required in order to determine the concentration of X.

ii Describe how you would prepare such a calibration curve.

d Another sample of water contained half of the concentration of X and Z but no Y or any other compound. Sketch the predicted chromatogram for this sample of water.

6 A student sets up a paper chromatogram and places a spot of green food dye on the origin. In one experiment the solvent moved 12 cm and a blue spot moved 9 cm from the origin. When the experiment was repeated the blue spot moved 15 cm. Determine how far (in cm) from the origin the solvent is likely to be in the second experiment.

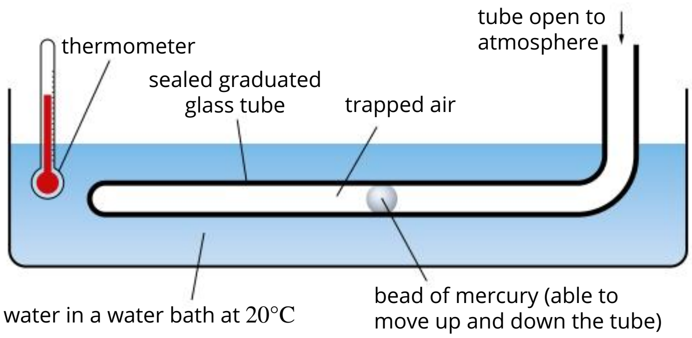
7 Using the kinetic theory of matter, explain why washing hanging on a clothes line will dry quicker on a hot day than on a cold day.

8 **a** Use the assumptions of the kinetic theory of gases to explain why raising the temperature of a constant mass of gas, at constant pressure, will result in an increase in the volume occupied by the gas.

b State the conditions under which real gases are *least* likely to obey the gas laws.

c The equipment used to measure the relationship between the volume of a gas and temperature is shown below.

Describe how the pressure would be kept constant in the equipment set-up.



d Sketch a graph that shows the relationship between pressure (vertical axis) and temperature in degrees Celsius (horizontal axis) for a particular amount of gas at a constant volume.

9 A sample of N_2 gas collected at 28.0°C and 740.0 mmHg pressure occupies 240.0 mL. Calculate the volume it will occupy at STP.

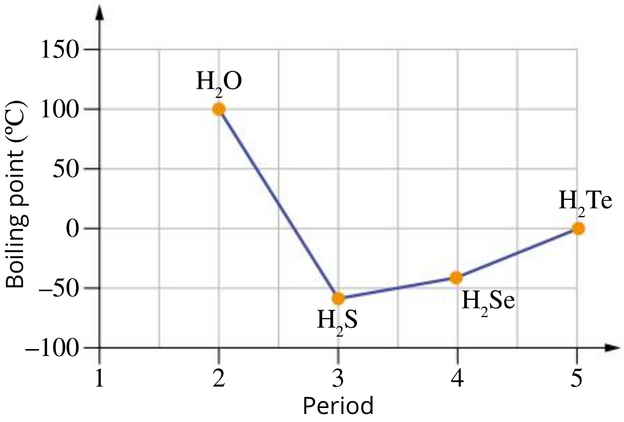
1 atm = 760 mm Hg = 101.325 kPa

10 The first step in the industrial manufacture of urea, a widely used fertiliser, is the reaction of ammonia with carbon dioxide at a high temperature and pressure, to form ammonium carbamate:

$$2NH_3(g) + CO_2(g) \rightarrow NH_2COONH_4$$

The ammonium carbamate then decomposes to form urea. If 1.00 kL of ammonia at 210°C and 240 atm reacts with a slight excess of carbon dioxide to form 177 kg of ammonium carbamate, calculate the percentage yield of this reaction.

11 The following graph shows the boiling points of four group 16 hydrides.



a Explain why the boiling point of H_2S is so much lower than that of H_2O.

b Explain why the boiling point of H_2S is also lower than that of H_2Se.

12 Ammonia (NH_3) is a constituent of many cleaning products for bathrooms.

a Draw an electron dot formula of an ammonia molecule, including non-bonding electron pairs.

b Draw a structural formula for two ammonia molecules. Clearly show, and give the name of, the shape of these molecules. On your diagram, label the type of bonds that exist between the:

- atoms within each ammonia molecule
- two ammonia molecules

c Draw a structural formula for a molecule of:

i nitrogen gas

ii carbon dioxide gas

d Explain the following statements.

i The bonds between nitrogen molecules and those between molecules of carbon dioxide are of the same type even though the bonds inside these molecules differ in strength and polarity.

ii The bonds between ammonia molecules are different from those between nitrogen molecules or carbon dioxide molecules.

13 **a** Describe the VSEPR hypothesis.

Consider the following compounds: H_2S, PF_3, $CHCl_3$, CS_2, BF_3.

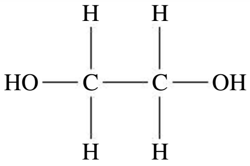
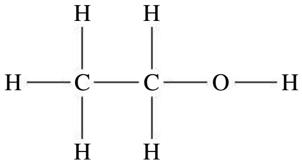
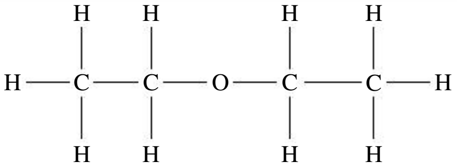
b Explain how the VSEPR hypothesis can be used to determine the shape of each molecule.

c Draw the structure of each compound and include any non-bonding pairs of electrons.

d Identify the shape of each molecule and calculate an approximate value of the bond angle at the central atom.

e Predict if each molecule is polar or non-polar and justify your prediction.

14 Some physical properties and structural formulas of three compounds are provided in the table below.

Compound	Ethane-1,2-diol	Ethanol	Diethyl ether
boiling point at 1.0 atm (°C)	197	78	35
vapour pressure at 20°C (mmHg)	0.6	5.9	440
molar mass ($g\,mol^{-1}$)	62	46	74
structural formula			

a Vapour pressure is defined as the pressure exerted by a vapour that is in equilibrium with its liquid in a closed container. Describe on the molecular level how this equilibrium is established.

b Explain the relationship between vapour pressure and boiling point.

c Explain the relative order of the following in terms of the molecular structure and the nature of intermolecular forces.

- **i** boiling points
- **ii** vapour pressure of the three compounds

15 Both hexanol, $CH_3CH_2CH_2CH_2$ CH_2CH_2OH, and CH_2OH methanol, CH_3OH, have a polar O–H bond capable of forming hydrogen bonds with water. Explain why methanol is soluble in water, while hexanol is insoluble.

16 A chemist analysed an alloy that contained magnesium, zinc and a small amount of non-metallic impurity. She took a 2.00 g sample of the alloy and divided it into two portions of equal mass.

One of these portions she reacted completely with dilute hydrochloric acid and obtained 0.420 L of hydrogen gas at 22.0°C and 100.0 kPa pressure.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$
$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

The other portion she reacted as completely as possible with sodium hydroxide solution and obtained 0.308 L of hydrogen gas at STP. In this part of the experiment, only the amphoteric metal, the zinc, reacts:

$$Zn(s) + 2OH^-(aq) + 2H_2O(l) \rightarrow Zn(OH)_4{}^{2-}(aq) + H_2(g)$$

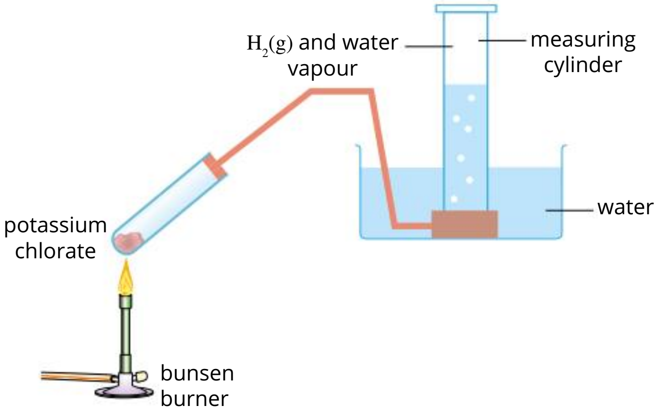
a Use the data from the second experiment to calculate the percentage by mass of zinc in the alloy.

b Use the information from the first experiment to determine the percentage, by mass, of magnesium in the alloy.

c Determine the percentage, by mass, of impurities in the alloy. Assume that the impurities did not react with either hydrochloric acid or sodium hydroxide.

17 The apparatus shown below was used to determine the molar volume of oxygen. The oxygen was generated in the test-tube by heating potassium chlorate and collected by the displacement of water in the measuring cylinder.

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$



The measurements made during the experiment are provided in the following table.

mass of test-tube + contents before heating	38.550 g
mass of test-tube + contents after heating	38.225 g
volume of gas collected	242.0 mL
atmospheric pressure	99.00 kPa
room temperature	20.0°C
vapour pressure of water at room temperature	2.34 kPa

a Calculate the pressure of the oxygen gas in the measuring cylinder.

b Determine the number of moles of oxygen produced.

c Calculate the molar volume of the oxygen at the experimental conditions.

d Determine the molar volume of oxygen at STP.

e Calculate the experimental error if the accepted value of the molar volume of a gas is 22.7 $L\,mol^{-1}$ at STP.

UNIT 2 • REVIEW

Topic 2: Aqueous solutions and acidity

Multiple-choice questions

1 State which type of bonds hold molecules of water in a crystalline lattice in ice.

A polar covalent bonds only
B hydrogen bonds and dispersion forces only
C polar covalent bonds and hydrogen bonds only
D polar covalent bonds, hydrogen bonds and dispersion forces

2 Cadmium is a metal used in photography, in nickel–cadmium batteries and solar cells, and for metal plating. Drinking water contaminated with cadmium can have harmful health effects. Determine the concentration of cadmium, in ppb, in a reservoir containing 500 ML (1 ML = 10^6 L) of water if 1.75 kg of cadmium is discharged into it.

A 3.5 ppb
B 875 ppb
C 1750 ppb
D 3000 ppb

3 Identify which of the following statements about solutions are true.

- **i** When sodium chloride is dissolved in water, the water is the solute.
- **ii** An unsaturated solution contains more solute than is needed to saturate it.
- **iii** A supersaturated solution contains less dissolved solid per unit volume than a saturated solution.
- **iv** In general, the solubility of a gas in a liquid decreases with increasing temperature.
- **v** When sodium nitrate dissolves in water, dissociation occurs.

A ii, iii and iv
B i, ii and iv
C ii and v
D iv and v

4 Identify which of the following does NOT change when a solution is diluted by the addition of more solvent.

A volume of solvent
B number of moles of solute
C concentration of solution
D mass of solution

5 Select the ionic equation for the reaction that occurs when solutions of barium chloride and sodium sulfate are mixed.

A $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
B $2Na^+(aq) + SO_4^{2-}(aq) \rightarrow Na_2SO_4(s)$
C $BaCl_2(aq) + NaSO_4(aq) \rightarrow NaCl(aq) + BaSO_4(s)$
D $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow 2NaCl(aq) + BaSO_4(s)$

6 Identify which of the following pairs of solutions would NOT form a precipitate when mixed together.

A magnesium chloride solution and potassium sulfide solution
B zinc chloride solution and potassium phosphate solution
C calcium chloride solution and copper(II) nitrate solution
D lead nitrate solution and iron(III) iodide solution

7 A solution of sodium hydroxide when added to a solution of aluminium chloride produces a gel-like precipitate. The precipitate reaction could be represented as:

A $Na^+(aq) + Cl^-(aq) \rightarrow NaCl(s)$
B $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$
C $Al^{3+}(aq) + 3Cl^-(aq) \rightarrow AlCl_3(s)$
D $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

8 You have been provided with three unlabelled beakers containing aqueous solutions of zinc chloride, sodium carbonate and barium nitrate. Select the reagent that could be used to identify each of the solutions.

A dilute nitric acid
B dilute sulfuric acid
C dilute hydrochloric acid
D sodium chloride solution

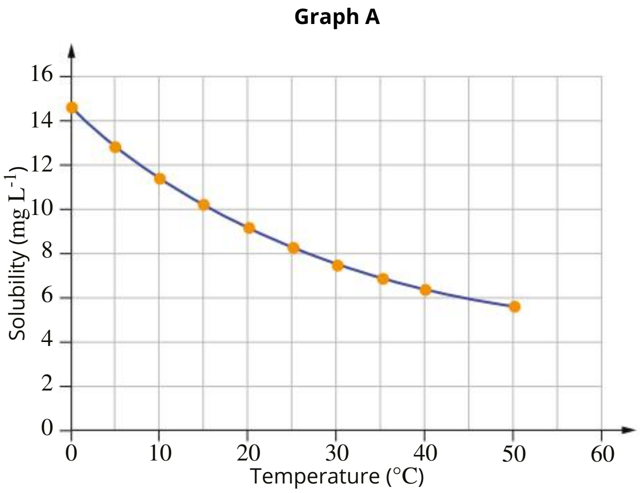
9 State which one of the following describes the types of bonds broken in the solute and formed with water when hydrogen chloride dissolves.

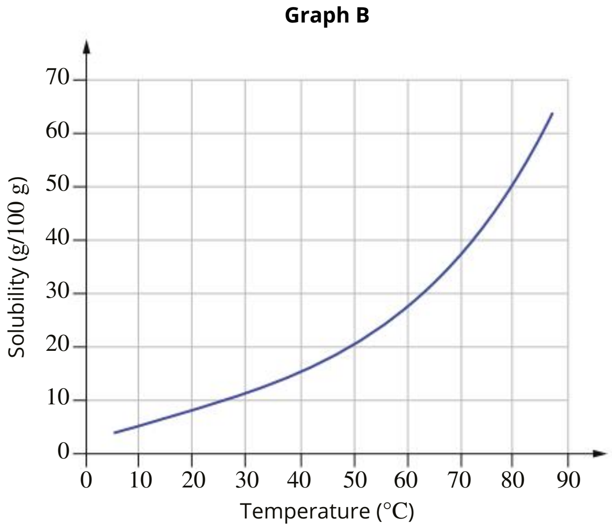
	Bonds broken	Bonds formed
A	covalent	hydrogen and dipole–dipole
B	dipole–dipole	covalent and ion–dipole
C	dipole–dipole	hydrogen and dipole–dipole
D	covalent	covalent and ion–dipole

10 Use the pH value of the solutions listed below to identify the substance that has the highest concentration of H_3O^+ ions.

A soap solution pH 9.0
B sea water pH 8.1
C rainwater pH 5.6
D wine pH 3.6

11 Consider the shape of the following solubility curves.





Select which one of the following alternatives correctly attributes the most likely shape of the solubility curve for the ionic solid potassium chlorate and the gas oxygen.

	Potassium chlorate	Oxygen
A	graph A	graph A
B	graph A	graph B
C	graph B	graph A
D	graph B	graph B

12 Beakers A and B both contain nitric acid. The pH of the acid in beaker A is 3 whereas the pH of the acid in beaker B is 1. From this information, select the correct deduction about the concentration of hydrogen ions in beaker A.

A It is three times that in beaker B.
B It is one-third that in beaker B.
C It is a hundred times that in beaker B.
D It is one-hundredth of that in beaker B.

13 Predict the products of the reaction between iron(II) carbonate and dilute hydrochloric acid.

A H_2O and $FeCl_2$
B H_2 and $FeCl_2$
C H_2, Fe, Cl_2 and CO_2
D CO_2, H_2O and $FeCl_2$

14 Identify the equation that represents the reaction between solid magnesium hydroxide and dilute nitric acid.

A $Mg(s) + 2H_3O^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l) + H_2(g)$
B $Mg^{2+}(aq) + 2NO_3^-(aq) \rightarrow Mg(NO_3)_2(s)$
C $Mg(OH)_2(s) + 2H_3O^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
D $Mg(OH)_2(s) + 2NO_3^-(aq) \rightarrow Mg(NO_3)_2(s) + 2OH^-(aq)$

Short-answer questions

1 After dissolving 5.00 g of table salt (NaCl) in water, a solution is made up to 200 mL with water.

a State a chemical equation to represent this process.
b Draw a labelled diagram to show the arrangement of water molecules around the dissolved sodium ions and chloride ions.
c Calculate the molar concentration of the solution.
d Explain why ice floats on water.
e Explain the term 'surface tension' and describe how it arises in water.

2 A student has three solutions of potassium chlorate ($KClO_3$). One is unsaturated, one is saturated and one is supersaturated.

a Define the term 'saturated'.
b Devise and describe a simple method that could be used to determine which solution is which. Describe the results you would expect for each solution.
c A solution of potassium chlorate is made by dissolving 5.67 g of the solid in distilled water to make up 250.0 mL of solution. The molar mass of $KClO_3$ is 122.6 g mol^{-1}.
    i Calculate the concentration of this solution in g L^{-1}.
    ii Calculate the concentration of this solution in mol L^{-1}.
    iii If this solution is added to a beaker containing 350 mL of water, determine the concentration of the diluted solution in mol L^{-1}.

3 Calculate the volume of water, in mL, that must be added to 20.0 mL of 0.5 mol L^{-1} NaCl to change its concentration to 0.20 mol L^{-1}.

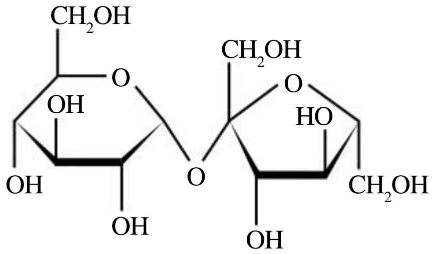
4 Some excess potassium chloride solution is added to a solution containing 3.15 g of lead(II) nitrate. A white precipitate forms.

- **a** Write a full chemical equation for the formation of the precipitate.
- **b** Write an ionic equation for the formation of the precipitate.
- **c** For this reaction, give the name of the:
  - **i** precipitate
  - **ii** spectator ions
- **d** Assume there was sufficient potassium chloride to react with all of the lead(II) nitrate.
  - **i** Calculate the number of moles of lead(II) nitrate that reacted.
  - **ii** Calculate the mass of precipitate that would form.

5 You are given five solutions. They are not labelled but are known to be sodium carbonate, sulfuric acid, potassium nitrate, magnesium nitrate and copper(II) nitrate. You have been set the task to identify the solutions.

- **a** Predict what you would observe if a sample of each solution was added to a sample of each of the other four solutions. (If no observable reaction is predicted, write 'no change'.)
  - **i** sodium carbonate and sulfuric acid
  - **ii** sodium carbonate and potassium nitrate
  - **iii** sodium carbonate and magnesium nitrate
  - **iv** sodium carbonate and copper(II) nitrate
  - **v** sulfuric acid and potassium nitrate
  - **vi** sulfuric acid and magnesium nitrate
  - **vii** sulfuric acid and copper(II) nitrate
  - **viii** potassium nitrate and magnesium nitrate
  - **ix** potassium nitrate and copper(II) nitrate
  - **x** magnesium nitrate and copper(II) nitrate
- **b** Write a balanced ionic equation for each of the reactions predicted in part **a**.
- **c** Describe how the solutions can be identified by mixing each solution with each of the other four solutions.

6 Common salt (sodium chloride, NaCl) is an ionic compound. Table sugar (sucrose, $C_{12}H_{22}O_{11}$) is a covalent molecular compound. The structure of sucrose is provided.



Both of these substances are highly soluble in water yet the solution processes for these substances are different. Give a brief description of the solution process for the following items.

- **a** common salt
- **b** sugar

7 When dilute sulfuric acid is added to substance A (a solid), there is a vigorous effervescence and the gas explodes when ignited in air. When the remaining solution is concentrated, colourless crystals separate. A confused student identifies these crystals as magnesium oxide, zinc sulfate, copper sulfate or calcium oxide.

- **a** Determine the identity of the crystals. Explain your reasoning.
- **b** Identify the metal and explain your answer.
- **c** Write an ionic equation to show the reaction of A with dilute sulfuric acid.

8 A 0.20 mol sample of hydrogen chloride gas is bubbled into 300.0 mL of distilled water to produce an acidic solution. The 300.0 mL of solution is divided evenly into three 200 mL flasks.

- **a** Write the full equation and an ionic equation for the reaction that occurs when a short length of magnesium is added to the first flask.
- **b** An 80.0 mL volume of 1.0 M silver nitrate solution ($AgNO_3$) is added to the second flask. A white solid forms.
  - **i** Write an ionic equation to explain the appearance of the white solid.
  - **ii** Following the addition of the silver nitrate solution, predict if any hydrogen chloride will remain in the mixture. Explain your answer with the aid of relevant calculations.
- **c** A small amount of copper(II) carbonate is added to the third flask.
  - **i** Describe the expected observations.
  - **ii** Write full and ionic equations to account for these observations.

9 The health authorities in Australia recommend the safe level of lead in drinking water to be 0.01 ppm. Despite lead and its compounds being very toxic, products used every day, such as car batteries, still contain lead. Lead and many of its compounds are insoluble in water, so it is assumed that very little lead enters the drinking water supply by the dissolution of lead substances.

Lead(II) oxide, a compound regarded as insoluble in water, is used to make crystal glassware and as a pottery glaze. The solubility of lead oxide at 25°C is 0.002 g per 100 g of water. If some water is stored in a bowl of crystal glass, determine whether this water could possibly dissolve sufficient lead oxide to raise the level of lead in the water over the recommended safe level.

10 The solubilities of three different substances at 20°C and 80°C are given in the table below.

Substance	Solubility in water (g per 100 g water)	
	20°C	80°C
potassium sulfate	12.0	21.0
sodium chloride	36	37
graphite	insoluble	insoluble

a If 16.0 g of sodium chloride and 50 g of water were mixed, determine what mass of sodium chloride would dissolve at 20°C.

b If 9.0 g of potassium sulfate and 50 g of water were mixed, determine what mass of potassium sulfate would dissolve at:

i 20°C

ii 80°C

c If 1.5 g of graphite and 50 g of water were mixed, determine what mass of graphite would dissolve at 20°C.

d Determine the solubility of potassium sulfate at 20°C expressed in mol L^{-1}.The density of water at 20°C is 1.00 g mL^{-1}. Assume that the volume of the solution is equal to the volume of water.

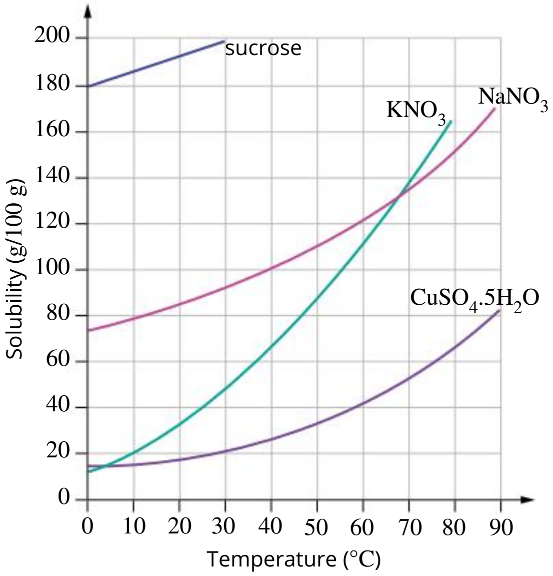
e You are given a mixture that contains 16.0 g of sodium chloride, 9.0 g of potassium sulfate and 1.5 g of graphite.

i Describe how you could obtain a pure sample of potassium sulfate and a pure sample of graphite from this mixture.

ii Determine the maximum mass of potassium sulfate that could be obtained by this method.

f Use the data provided in the table above to explain why it would be difficult to obtain a pure sample of sodium chloride from the mixture.

11 Consider the following solubility curves.



a **i** Determine the minimum mass of NaNO$_3$ that must be dissolved in 100 g of water to produce a saturated solution at 50°C.

ii Predict what would be observed if the solution in part **i** were cooled to 20°C.

iii Determine what mass of water must be added to 110 g of NaNO$_3$ to make a saturated solution at 10°C.

b Determine the maximum amount of table sugar (sucrose) that can dissolve in a cup containing 200 g of water at 30°C.

c Three different solutions of KNO$_3$ are prepared by adding various masses of KNO$_3$ to 100 g of water. All three solutions are maintained at 50°C.

The masses of KNO$_3$ contained in the solutions are 30 g, 80 g and 87 g.

i Use this information to clearly explain the differences between a saturated solution, an unsaturated solution and a supersaturated solution.

ii Describe a simple test that you could use to identify which solution is which.

iii Predict what would happen to each solution if it was cooled to 30°C.

d Determine what mass of water must be added to 35 g of CuSO$_2$.5H$_2$O to make a saturated solution at 20°C.

12 In 1887, the Swedish scientist Svante Arrhenius developed a model that explains some of the properties of acids and bases.

a Describe how acids and bases are defined in the Arrhenius model.

b The results of a student investigation into some properties of strong and weak acids are summarised in the table below.

Explain, in terms of the Arrhenius model, the difference in pH, and reactivity between the two acids.

Properties of strong and weak acids

Property	Hydrochloric acid, HCl	Ethanoic acid, CH_3COOH
volume of acid solution	20.0 mL	20.0 mL
concentration of acid	0.10 mol L^{-1}	0.10 mol L^{-1}
pH	1.0	3.0
electrical conductivity	high	low
reactivity with sodium carbonate solution	fast	slow

c In a follow-up experiment, the students found that 10.0 mL of 0.20 mol L^{-1} sodium hydroxide solution was required to neutralise the hydrochloric acid solution. Predict if you would expect the volume of sodium hydroxide required to neutralise the ethanoic acid solution to be less than or equal to 10.0 mL. The concentration of base used in the two neutralisation reactions is the same. Explain your answer.

Topic 3: Rates of chemical reactions

Multiple-choice questions

1 State which of the following will NOT increase the rate of the reaction between 0.2 g of $CaCO_3$ and 100 mL of 0.1 M hydrochloric acid.

A crushing the $CaCO_3$

B using a larger volume of the acid

C increasing the temperature of the acid

D using more concentrated acid

2 Consider the statement: 'The rates of all chemical reactions increase with a rise in temperature.'

Identify which of the following is correct.

A This statement is true, since the activation energies for all reactions increase with a rise in temperature.

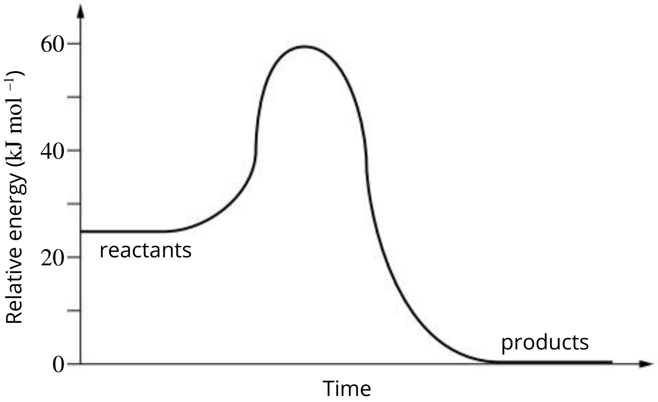
B This statement is false, since the rate of an exothermic reaction decreases with a rise in temperature.

C This statement is false, since the rate of an endothermic reaction decreases with a rise in temperature.

D This statement is true, since a greater proportion of molecular collisions will result in reaction.

The following information refers to Questions **3** and **4**.

The graph shows the potential energy for a chemical reaction.



3 Determine the correct values for the enthalpy change, Δ*H*, and the activation energy, E_a, in kJ mol^{-1} for the reaction.

A $\Delta H = -25$ kJ mol^{-1} and $E_a = 35$ kJ mol^{-1}

B $\Delta H = +25$ kJ mol^{-1} and $E_a = 35$ kJ mol^{-1}

C $\Delta H = 35$ kJ mol^{-1} and $E_a = 60$ kJ mol^{-1}

D $\Delta H = -60$ kJ mol^{-1} and $E_a = 35$ kJ mol^{-1}

4 Predict what happens if a catalyst is added to the system.

A Both Δ*H* and E_a decrease.

B Δ*H* remains constant but E_a decreases.

C Δ*H* decreases and E_a remains constant.

D Both Δ*H* and E_a remain constant.

5 Select the alternative that explains why the rate of a chemical reaction decreases as the reactions proceeds.

A The reactant concentrations decrease with time.

B A catalyst is needed to maintain a constant rate of reaction.

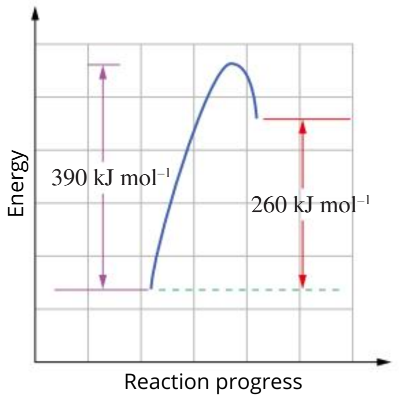
C The proportion of reactant molecules with energies in excess of the activation energy decreases as the reaction proceeds.

D Heat energy must be added to maintain the reaction rate.

The following information refers to Questions **6** and **7**.

The energy profile diagram shown is for the following reaction.

$$CO_2(g) + NO(g) \rightarrow CO(g) + NO_2(g)$$



6 Determine the ΔH of the forward reaction, in kJ mol^{-1}.

A −130

B +130

C +230

D +260

7 Determine the activation energy of the reverse reaction, in kJ mol^{-1}.

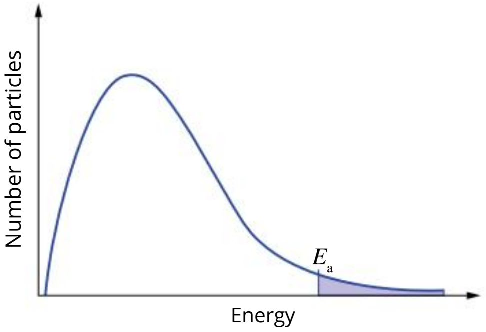
A 40

B 130

C 230

D 360

8 The diagram below shows the typical Maxwell–Boltzmann distribution of kinetic energies of the particles in a reaction mixture.



If E_a represents the activation energy for the reaction, identify what the shaded area represents.

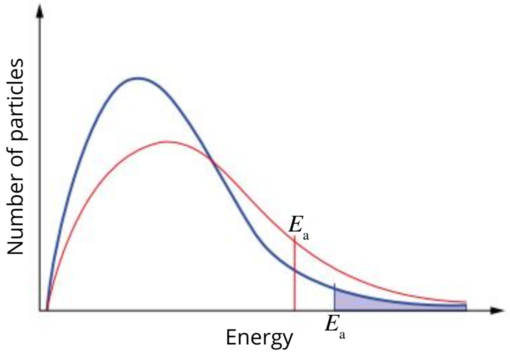
A the total amount of energy that will be released when the reaction reaches equilibrium

B the proportion of reactant particles that are moving fast enough at a given instant for a collision to result in the formation of products

C the energy that must be added to the reactant mixture to initiate the reaction

D the proportion of reactant particles that will be converted to products when the reaction reaches equilibrium

9 Two changes are made to the reaction mixture from Question **8**, affecting both the distribution curve and the activation energy, as shown in red in the diagram below.



State which of the following gives changes in conditions that are consistent with these effects.

	Change in distribution curve	Change in activation energy
A	higher temperature	catalyst added
B	higher temperature	lower pressure
C	higher pressure	lower temperature
D	higher pressure	catalyst added

10 Identify which of the following statements about a catalyst is NOT true.

A A catalyst is not consumed in the course of a reaction.

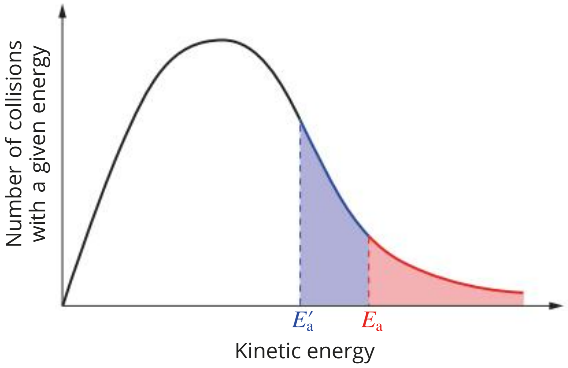
B A catalyst alters the reaction pathway between reactants and products.

C A catalyst reduces the energy released or absorbed by a reaction.

D The proportion of molecules with sufficient energy to react is increased by a catalyst.

11 In the following Maxwell–Boltzman distribution curve:

- E_a represents the activation energy for an uncatalysed reaction
- E_a' represents the activation energy for the catalysed reaction.



Select the coloured region under the graph that represents the number of particles with sufficient energy to lead to the successful formation of products in a catalysed chemical reaction.

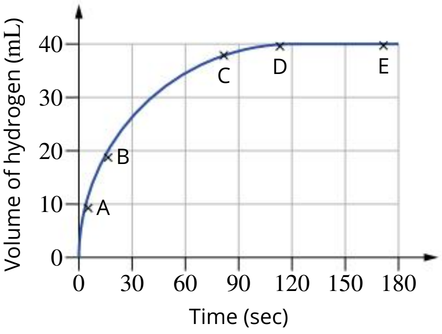
A blue only

B white only

C blue and pink only

D white and blue only

12 Refer to the following graph, which shows the volume of hydrogen produced against time during a rate of reaction experiment between zinc and dilute hydrochloric acid.



Determine the point in the graph at which the reaction rate is the greatest.

Short-answer questions

1 State one possible parameter for increasing the rate of each of the following reactions. Choose a *different procedure* in each case.

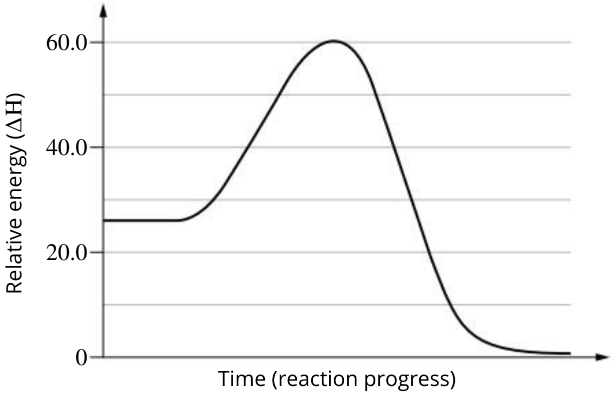
a A 1 cm cube of zinc is added to 4.0 mol L^{-1} HCl.

b Hydrogen peroxide decomposes to water and oxygen.

c A large crystal of washing soda (hydrated sodium carbonate) is reacted with 1.0 mol L^{-1} nitric acid.

d Solutions of 0.1 mol L^{-1} of potassium iodide and lead (II) nitrate are mixed and a yellow precipitate of lead(II) iodide forms.

2 Examine the potential energy diagram for a chemical reaction.



a Determine ΔH and the activation energy for the forward reaction.

b State if the reverse reaction is exothermic or endothermic.

c When the forward reaction occurs, determine if the temperature of the surroundings increases or decreases.

d Predict how ΔH and the activation energy will be affected if a catalyst is added to the system.

e On the same axes, sketch the approximate potential energy diagram for the catalysed reaction.

3 **a** Explain this statement:

> 'Many solid catalysts are used in pellet form or as wire gauze or as nanoscale particles, rather than as a large sheet.'

b A reaction occurs according to the following equation.

$$X(g) + 3Y(g) \rightarrow 2Z(g)$$

i The rate of reaction of Y at a certain temperature is 2.4×10^{-4} mol L^{-1} s^{-1}. Calculate the rate of reaction of X at the same temperature.

ii The energy of activation for the forward reaction to produce Z as shown in the equation is 60 kJ mol^{-1}. The ΔH for the reaction as shown is –90 kJ mol^{-1}. Calculate the energy of activation for the reverse reaction.

4 The activation energy for the decomposition of hydrogen peroxide to oxygen was measured under two different conditions.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

I The activation energy for the decomposition reaction when an enzyme was added was found to be 36.4 kJ mol^{-1}. The temperature of the reaction mixture increased.

II When platinum was added to another sample of the hydrogen peroxide solution, the activation energy was 49.0 kJ mol^{-1}. The temperature of the reaction mixture increased.

a State the function of the enzyme and platinum in each of these reactions.

b Sketch, on the same set of axes, the energy profiles for the decomposition of hydrogen peroxide with the enzyme and the decomposition using the platinum.

c Deduce which reaction system, I or II, would be faster. Explain your answer.

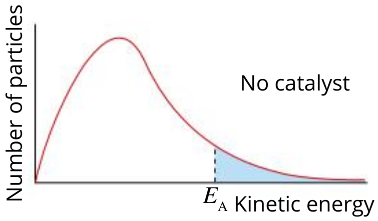
5 Hydrogen iodide decomposes as follows.

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g) \quad \Delta H = -26\ kJ\,mol^{-1}$$

The activation energy for the uncatalysed reaction is 186 kJ mol^{-1} and for a catalysed reaction it is 100 kJ mol^{-1}.

a On the graph below, sketch appropriately labelled energy profiles for both the catalysed and uncatalysed reactions.

b **i** The graph below illustrates the number of particles that decompose in a reaction system with no catalyst. Draw a graph to illustrate the reaction system if a catalyst were used.



ii Explain why a catalyst increases the rate of a reaction.

6 The fizzy sensation from sherbet lolly powder is caused by a gas produced in a reaction between two substances. The sherbet powder consists of the two dry reactants citric acid, $C_6H_8O_7$, and sodium hydrogencarbonate, and a sweetener such as icing sugar. The gas-producing chemical reaction does not begin until the dry ingredients become wet. As soon as they dissolve into the saliva in your mouth, the reaction begins. In this reaction the citric acid reacts with the hydrogencarbonate ions to form carbon dioxide (the fizz), water and citrate ions, $C_6H_5O_7^{3-}$.

a Explain in terms of the collision theory why a reaction occurs much quicker when the two reactants are 'wet'.

b Predict if you would expect the rate of 'fizz production' to change if some sherbet powder were to be poured into hot water. Explain your answer in terms of the collision theory.

7 Hydrogen gas is produced when a magnesium strip is added to dilute hydrochloric acid. The rate of reaction may change in any of these ways.

A The rate of reaction is decreased.
B The rate of reaction is increased.
C The rate of reaction is unchanged.

Choose the effect (A, B or C) that is most appropriate to each of the following situations. Justify your selection in terms of collision theory.

a The concentration of acid used is decreased.
b The reaction mixture is stirred.
c Magnesium in powdered form is used instead of a magnesium strip.
d The temperature is raised from 20°C to 30°C.
e A catalyst is added to the reaction mixture.

8 Ethanoic acid, CH_3COOH, is industrially manufactured from methanol according to the following reaction:

$$CH_3OH(l) + CO(g) \rightarrow CH_3COOH(l); \qquad \Delta H = -137\ kJ\,mol^{-1}$$

The most modern chemical plants now use rhodium, Rh, as a catalyst for this reaction.

a **i** Explain what is meant by the term '$\Delta H = -137$ kJ mol^{-1}' written after the chemical equation.

ii Calculate the amount of heat released or absorbed when 100 g of ethanoic acid is formed in the reaction given above.

b **i** Explain the role of a catalyst in a reaction.

ii In terms of the collision theory, explain how a catalyst affects the rate of a reaction. Include an energy profile of the reaction in your answer.

9 When oxygen gas is reacted with hydrogen bromide, bromine gas is produced and heat is evolved according to the equation:

$$4HBr(g) + O_2(g) \rightarrow Br_2(g) + 2H_2O(g)$$

a Draw an energy profile for the reaction, clearly labelling the enthalpy of reactants and products, the ΔH for the reaction and the energy of activation. Indicate the parts on your diagram that represent the transition state, and the regions where energy is absorbed and released.

b **i** Define the term 'activation energy'.

ii Indicate how activation energy is related to the strength and number of bonds.

iii Describe the relationship between the rate of a reaction and activation energy

c Explain what is meant by the transitional state.

d The use of a catalyst does not change the value for ΔH for this reaction. Explain why.

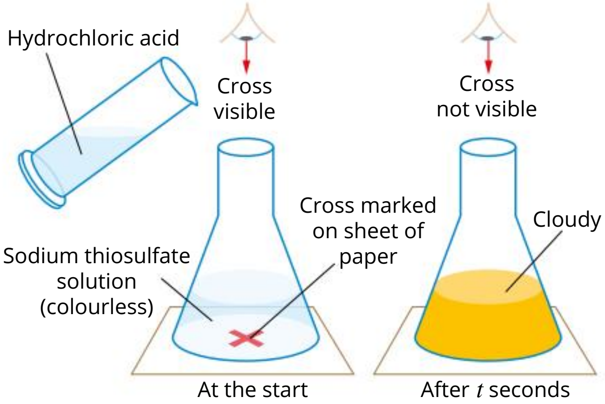
e Express the rate of the reaction given above in terms of the rate of change in the concentration of reactants and products.

UNIT 2 • REVIEW

10 A student investigated the factors affecting the rate of reaction between a solution of sodium thiosulfate and hydrochloric acid. The equation for the reaction is:

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$$

The reaction was carried out in a conical flask placed on top of a piece of white paper with a dark cross marked on it. The rate of reaction was determined by measuring the time taken for the cross to be masked by the suspension of sulfur formed during the reaction, as shown.



The rate was determined for different concentrations of sodium thiosulfate and for different temperatures. The volume of each solution and the concentration of hydrochloric acid was kept constant. The results are summarised in the table below.

Experiment number	[$Na_2S_2O_3$(aq)] (M)	Temperature (°C)	Time taken for the cross to be masked (s)
1	0.1	20	36
2	0.2	20	20
3	0.1	25	28

- **a** Explain, in terms of collision theory, why the rate in experiment 2 is higher than the rate in experiment 1.
- **b** Explain, in terms of collision theory, why the rate in experiment 3 is higher than the rate in experiment 1.
- **c** Predict what factors, other than the two investigated in this experiment, may affect the rate of a reaction.

11 A group of students was asked to measure the rate of the reaction between 1.0 mol L^{-1} hydrochloric acid and marble chips (calcium carbonate).

- **a** Write an equation for the reaction that occurs when marble chips are added to dilute hydrochloric acid.

One student investigated the rate of the reaction by measuring, at 10 second intervals, the volume of carbon dioxide collected in a gas syringe.

- **b** Draw a labelled diagram indicating how you would assemble the apparatus.

The students' results are shown in the following table.

Time (sec)	Volume (mL)
0	0.0
10	27
20	44
30	54
40	60
50	65
60	68
70	71
80	73
90	74
100	76
110	77
120	78
130	79
140	80
150	80
160	80

- **c** Draw a graph showing how the volume of carbon dioxide varies with time.
  - **i** Describe how you would determine the initial rate of reaction.
  - **ii** Determine the instantaneous rate at 10 seconds.
  - **iii** Calculate the average rate of production of carbon dioxide between 30 and 60 seconds.
  - **iv** Calculate the average rate of consumption of hydrochloric acid between 30 and 60 seconds.
- **d** Another student decided to determine the rate of the same reaction by recording the loss in mass in the reaction system at 10-second intervals.
  - **i** Explain why this would be a valid method for investigating the rate of the reaction.
  - **ii** Draw a labelled diagram indicating how the required apparatus could be assembled.
  - **iii** Draw a graph showing how the mass of the reaction system changes over time.
  - **iv** On the same graph, sketch the change in mass of the reactions system when the concentration of the hydrochloric acid used is 2.0 mol L^{-1} instead of 1.0 mol L^{-1}.

APPENDIX A Symbols, units and fundamental constants

TABLE 1 Units and symbols based on the SI system*

Quantity	Symbol for physical quantity	Corresponding SI unit	Symbol for SI unit	Definition of SI unit
Mechanics				
length	l	metre	m	fundamental unit
area	A	square metre	m^{2}	
volume	V	cubic metre	m^{3}	
mass	m	kilogram	kg	fundamental unit
density	d	–	kg m^{-3}	
time	t	second	s	fundamental unit
force	F	newton	N	kg m s^{-2}
pressure	P	pascal	Pa	N m^{-2}
energy	E	joule	J	N m
Electricity				
electric current	I	ampere	A	fundamental unit
electric charge	Q	coulomb	C	A s
electric potential difference	V	volt	V	J A^{-1} s^{-1}
Nuclear and chemical quantities				
atomic number	Z	–	–	–
neutron number	N	–	–	–
mass number	A	–	–	$Z + N$
amount of substance	n	mole	mol	fundamental unit
relative atomic mass	A_r	–	–	–
relative molecular mass	M_r	–	–	–
molar mass	M	–	–	kg mol^{-1}
molar volume	V_m	–	–	m^{3} mol^{-1}
concentration	c	–	–	mol dm^{-3} or mol L^{-1}
Thermal quantities				
temperature	T	kelvin	K	fundamental unit
specific heat capacity	c	–	–	J kg^{-1} K^{-1}

*Units listed in red are the arbitrarily defined fundamental units of the SI system.

TABLE 2 SI prefixes, their symbols and values

SI prefix	Symbol	Value
pico	p	10^{-12}
nano	n	10^{-9}
micro	µ	10^{-6}
milli	m	10^{-3}
centi	c	10^{-2}
deci	d	10^{-1}
kilo	k	10^{3}
mega	M	10^{6}
giga	G	10^{9}
tera	T	10^{12}

TABLE 3 Some physical constants

Description	Symbol	Value
Avogadro's constant	N_A	6.02 × 10^{23} mol^{-1}
mass of electron	m_e	9.109 × 10^{-31} kg
mass of proton	m_p	1.673 × 10^{-27} kg
mass of neutron	m_n	1.675 × 10^{-27} kg
gas constant	R	8.31 J mol^{-1} K^{-1}
ionic product for water	K_w	1.0 × 10^{-14} mol^{2} L^{-2} at 298 K
molar volume of an ideal gas	V_m	
at 273 K, 100 kPa		22.7 L mol^{-1}
specific heat capacity of water	c	4.18 J g^{-1} K^{-1}
density of water	d	1.00 g mL^{-1} at 298 K
Temperature in kelvin	T_K	T°C + 273
Pressure	KPa	100 kPa
Atmospheric pressure	KPa	1atm = 101.325 kPa

APPENDIX B Table of relative atomic masses

TABLE 1 Table of relative atomic masses*

Element name	Symbol	Atomic number	Relative atomic mass	Element name	Symbol	Atomic number	Relative atomic mass	Element name	Symbol	Atomic number	Relative atomic mass
actinium	Ac	89	–	hafnium	Hf	72	178.49	praseodymium	Pr	59	140.98
aluminium	Al	13	26.98	hassium	Hs	108	–	promethium	Pm	61	–
americium	Am	95	–	helium	He	2	4.00	protactinium	Pa	91	231
antimony	Sb	51	121.76	holmium	Ho	67	164.93	radium	Ra	88	–
argon	Ar	18	39.95	hydrogen	H	1	1.01	radon	Rn	86	–
arsenic	As	33	74.92	indium	In	49	114.82	rhenium	Re	75	186.21
astatine	At	85	–	iodine	I	53	126.90	rhodium	Rh	45	102.91
barium	Ba	56	137.33	iridium	Ir	77	192.22	roentgenium	Rg	111	–
berkelium	Bk	97	–	iron	Fe	26	55.85	rubidium	Rb	37	85.47
beryllium	Be	4	9.01	krypton	Kr	36	83.80	ruthenium	Ru	44	101.07
bismuth	Bi	83	208.98	lanthanum	La	57	138.91	rutherfordium	Rf	104	–
bohrium	Bh	107	–	lawrencium	Lr	103	–	samarium	Sm	62	150.4
boron	B	5	10.81	lead	Pb	82	207.22	scandium	Sc	21	44.96
bromine	Br	35	79.90	lithium	Li	3	6.94	seaborgium	Sg	106	–
cadmium	Cd	48	112.41	livermorium	Lv	116	–	selenium	Se	34	78.97
caesium	Cs	55	132.91	lutetium	Lu	71	174.97	silicon	Si	14	28.09
calcium	Ca	20	40.08	magnesium	Mg	12	24.31	silver	Ag	47	107.87
californium	Cf	98	–	manganese	Mn	25	54.94	sodium	Na	11	22.99
carbon	C	6	12.01	meitnerium	Mt	109	–	strontium	Sr	38	87.62
cerium	Ce	58	140.12	mendelevium	Md	101	–	sulfur	S	16	32.06
chlorine	Cl	17	35.45	mercury	Hg	80	200.59	tantalum	Ta	73	180.95
chromium	Cr	24	52.00	molybdenum	Mo	42	95.95	technetium	Tc	43	–
cobalt	Co	27	58.93	moscovium	Mc	115	–	tellurium	Te	52	127.60
copernicium	Cn	112	–	neodymium	Nd	60	144.24	tennessine	Ts	117	–
copper	Cu	29	63.55	neon	Ne	10	20.18	terbium	Tb	65	158.93
curium	Cm	96	–	neptunium	Np	93	–	thallium	Tl	81	204.38
darmstadtium	Ds	110	–	nickel	Ni	28	58.69	thorium	Th	90	232.04
dubnium	Db	105	–	nihonium	Nh	113	–	thulium	Tm	69	168.93
dysprosium	Dy	66	162.50	niobium	Nb	41	92.91	tin	Sn	50	118.71
einsteinium	Es	99	–	nitrogen	N	7	14.01	titanium	Ti	22	47.87
erbium	Er	68	167.26	nobelium	No	102	–	tungsten	W	74	183.84
europium	Eu	63	151.96	oganesson	Og	118	–	uranium	U	92	238.03
fermium	Fm	100	–	osmium	Os	76	190.2	vanadium	V	23	50.942
flerovium	Fl	114	–	oxygen	O	8	16.00	xenon	Xe	54	131.29
fluorine	F	9	19.00	palladium	Pd	46	106.42	ytterbium	Yb	70	173.05
francium	Fr	87	–	phosphorus	P	15	30.97	yttrium	Y	39	88.91
gadolinium	Gd	64	157.25	platinum	Pt	78	195.08	zinc	Zn	30	65.38
gallium	Ga	31	69.72	plutonium	Pu	94	–	zirconium	Zr	40	91.22
germanium	Ge	32	72.63	polonium	Po	84	–				
gold	Au	79	196.97	potassium	K	19	39.10				

*Based on the atomic mass of $^{12}C = 12$.
The values for relative atomic masses given in the table apply to elements as they exist in nature (without artificial alteration of their isotopic composition) and to natural mixtures that do not include isotopes of radiogenic origin.

APPENDIX C

Electronegativities and first ionisation energy of selected elements

Atomic Number

Symbol

electronegativity

first ionisation energy (kJ mol^{-1})

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 2.2 1310																	2 **He** 2372
3 **Li** 1.0 520	4 **Be** 1.6 900											5 **B** 2.0 801	6 **C** 2.6 1087	7 **N** 3.0 1402	8 **O** 3.4 1314	9 **F** 4.0 1681	10 **Ne** 2081
11 **Na** 0.9 496	12 **Mg** 1.3 738											13 **Al** 1.6 578	14 **Si** 1.9 787	15 **P** 2.2 1012	16 **S** 2.6 1000	17 **Cl** 3.2 1251	18 **Ar** 1521
19 **K** 0.8 419	20 **Ca** 1.0 590	21 **Sc** 1.4 633	22 **Ti** 1.5 659	23 **V** 1.6 651	24 **Cr** 1.7 653	25 **Mn** 1.6 717	26 **Fe** 1.8 763	27 **Co** 1.9 760	28 **Ni** 1.9 737	29 **Cu** 1.9 746	30 **Zn** 1.7 906	31 **Ga** 1.8 579	32 **Ge** 2.0 762	33 **As** 2.0 945	34 **Se** 2.6 941	35 **Br** 3.0 1140	36 **Kr** 1351
37 **Rb** 0.8 403	38 **Sr** 1.0 550	39 **Y** 1.2 600	40 **Zr** 1.3 640	41 **Nb** 1.6 652	42 **Mo** 2.2 684	43 **Tc** 2.1 702	44 **Ru** 2.2 710	45 **Rh** 2.3 720	46 **Pd** 2.2 804	47 **Ag** 1.9 731	48 **Cd** 1.7 868	49 **In** 1.8 558	50 **Sn** 2.0 709	51 **Sb** 2.0 831	52 **Te** 2.1 869	53 **I** 2.7 1008	54 **Xe** 2.6 1170
55 **Cs** 0.8 376	56 **Ba** 0.9 503																

Groups in accordance with IUPAC numbering.

Atomic and ionic radii of selected elements

Atomic radius (×10^{-12} m)
Symbol
Ionic radius (charge of ion)

32 **H**																	37 **He**
130 **Li** 76 (1+)	99 **Be** 45 (2+)											84 **B** 27 (3+)	75 **C** 16 (4+)	71 **N** 146 (3–)	64 **O** 140 (2–)	60 **F** 133 (1–)	62 **Ne**
160 **Na** 102 (1+)	140 **Mg** 72 (2+)											124 **Al** 54 (3+)	114 **Si** 40 (4+)	109 **P** 38 (5+)	104 **S** 184 (2–)	100 **Cl** 181 (1–)	101 **Ar**
200 **K** 138 (1+)	174 **Ca** 100 (2+)	159 **Sc** 75 (3+)	148 **Ti** 86 (2+) 61 (4+)	144 **V** 79 (2+) 54 (5+)	130 **Cr** 62 (3+) 44 (6+)	129 **Mn** 83 (2+) 54 (4+)	124 **Fe** 61 (2+) 55 (3+)	118 **Co** 65 (2+) 55 (3+)	117 **Ni** 69 (2+)	122 **Cu** 77 (1+) 73 (2+)	120 **Zn** 74 (2+)	123 **Ga** 62 (3+)	120 **Ge** 53 (4+) 272 (4–)	120 **As** 58 (3+) 46 (5+)	118 **Se** 196 (2–)	117 **Br** 196 (1–)	116 **Kr**
215 **Rb** 152 (1+)	190 **Sr** 118 (2+)	176 **Y** 90 (3+)	164 **Zr** 72 (4+)	156 **Nb** 72 (3+) 64 (5+)	146 **Mo** 65 (4+)	138 **Tc** 65 (4+)	136 **Ru** 68 (3+) 62 (4+)	134 **Rh** 67 (3+) 60 (4+)	130 **Pd** 86 (2+) 62 (4+)	136 **Ag** 115 (1+)	140 **Cd** 95 (2+)	142 **In** 80 (3+)	140 **Sn** 118 (2+) 69 (4+)	140 **Sb** 76 (3+)	137 **Te** 221 (2–)	136 **I** 220 (2–)	136 **Xe**
238 **Cs** 167 (1+)	206 **Ba** 135 (2+)																

APPENDIX E The electrochemical series

TABLE 1 The electrochemical series. The strongest oxidising agents are at the top left of the table and the strongest reducing agents are at the bottom right of the table

	Oxidising agents		Reducing agents		E° (V)
↑ Increasing oxidising strength	$F_2(g) + 2e^-$	$\rightleftharpoons$	$2F^-(aq)$	Increasing reducing strength ↓	+2.87
	$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$2H_2O(l)$		+1.77
	$Au^+(aq) + e^-$	$\rightleftharpoons$	$Au(s)$		+1.68
	$Cl_2(g) + 2e^-$	$\rightleftharpoons$	$2Cl^-(aq)$		+1.36
	$O_2(g) + 4H^+(aq) + 4e^-$	$\rightleftharpoons$	$2H_2O(l)$		+1.23
	$Br_2(l) + 2e^-$	$\rightleftharpoons$	$2Br^-(aq)$		+1.09
	$Ag^+(aq) + e^-$	$\rightleftharpoons$	$Ag(s)$		+0.80
	$Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$		+0.77
	$O_2(g) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2O_2(aq)$		+0.68
	$I_2(s) + 2e^-$	$\rightleftharpoons$	$2I^-(aq)$		+0.54
	$O_2(g) + 2H_2O(l) + 4e^-$	$\rightleftharpoons$	$4OH^-(aq)$		+0.40
	$Cu^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Cu(s)$		+0.34
	$Sn^{4+}(aq) + 2e^-$	$\rightleftharpoons$	$Sn^{2+}(aq)$		+0.15
	$S(s) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2S(g)$		+0.14
	$2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2(g)$		0.00
	$Pb^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Pb(s)$		–0.13
	$Sn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Sn(s)$		–0.14
	$Ni^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Ni(s)$		–0.23
	$Co^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Co(s)$		–0.28
	$Fe^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Fe(s)$		–0.44
	$Zn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Zn(s)$		–0.76
	$2H_2O(l) + 2e^-$	$\rightleftharpoons$	$H_2(g) + 2OH^-(aq)$		–0.83
	$Mn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Mn(s)$		–1.03
	$Al^{3+}(aq) + 3e^-$	$\rightleftharpoons$	$Al(s)$		–1.67
	$Mg^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Mg(s)$		–2.34
	$Na^+(aq) + e^-$	$\rightleftharpoons$	$Na(s)$		–2.71
	$Ca^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Ca(s)$		–2.87
	$K^+(aq) + e^-$	$\rightleftharpoons$	$K(s)$		–2.93
	$Li^+(aq) + e^-$	$\rightleftharpoons$	$Li(s)$		–3.02

APPENDIX F Common ions and solubilities of ionic compounds

TABLE 1 Names and formulas of some common positive and negative ions

Positive ions (cations)						Negative ions (anions)			
+1		+2		+3		−1		−2	
caesium	Cs^{+}	barium	Ba^{2+}	aluminium	Al^{3+}	acetate (ethanoate)	CH_3COO^{-}	carbonate	CO_3^{2-}
copper(I)	Cu^{+}	cadmium(II)	Cd^{2+}	chromium(III)	Cr^{3+}	bromide	Br^{-}	chromate	CrO_4^{2-}
gold(I)	Au^{+}	calcium	Ca^{2+}	gold(III)	Au^{3+}	chloride	Cl^{-}	dichromate	$Cr_2O_7^{2-}$
lithium	Li^{+}	cobalt(II)	Co^{2+}	iron(III)	Fe^{3+}	cyanide	CN^{-}	hydrogen phosphate	HPO_4^{2-}
potassium	K^{+}	copper(II)	Cu^{2+}			dihydrogen phosphate	$H_2PO_4^{-}$	oxalate	$C_2O_4^{2-}$
rubidium	Rb^{+}	iron(II)	Fe^{2+}	+4		fluoride	F^{-}	oxide	O^{2-}
silver	Ag^{+}	lead(II)	Pb^{2+}	lead(IV)	Pb^{4+}	hydrogen carbonate	HCO_3^{-}	sulfate	SO_4^{2-}
sodium	Na^{+}	magnesium	Mg^{2+}	tin(IV)	Sn^{4+}	hydrogen sulfate	HSO_4^{-}	sulfide	S^{2-}
		manganese(II)	Mn^{2+}			hydrogen sulfide	HS^{-}	sulfite	SO_3^{2-}
		mercury(II)	Hg^{2+}			hydrogen sulfite	HSO_3^{-}		
		nickel	Ni^{2+}			hydroxide	OH^{-}	−3	
		strontium	Sr^{2+}			iodide	I^{-}	nitride	N^{3-}
		tin(II)	Sn^{2+}			nitrate	NO_3^{-}	phosphate	PO_4^{3-}
		zinc	Zn^{2+}			nitrite	NO_2^{-}	phosphide	P^{3-}
						permanganate	MnO_4^{-}		

TABLE 2 Solubility of common ionic compounds in water

Soluble ionic compounds		
Soluble in water (>0.1 mol dissolves per L at 25°C)	**Exceptions: insoluble (<0.01 mol dissolves per L at 25°C)**	**Exceptions: slightly soluble (0.01–0.1 mol dissolves per L at 25°C)**
most chlorides (Cl^{-}), bromides (Br^{-}) and iodides (I^{-})	$AgCl$, $AgBr$, AgI, PbI_2	$PbCl_2$, $PbBr_2$
all nitrates (NO_3^{-})	no exceptions	no exceptions
all ammonium (NH_4^{+}) salts	no exceptions	no exceptions
all sodium (Na^{+}) and potassium (K^{+}) salts	no exceptions	no exceptions
all ethanoates (CH_3COO^{-})	no exceptions	no exceptions
most sulfates (SO_4^{2-})	$SrSO_4$, $BaSO_4$, $PbSO_4$	$CaSO_4$, Ag_2SO_4
Insoluble ionic compounds		
Insoluble in water	**Exceptions: soluble**	**Exceptions: slightly soluble**
most hydroxides (OH^{-})	$NaOH$, KOH, $Ba(OH)_2$, NH_4OH*, $AgOH$**	$Ca(OH)_2$, $Sr(OH)_2$
most carbonates (CO_3^{2-})	Na_2CO_3,K_2CO_3, $(NH_4)_2CO_3$	no exceptions
most phosphates (PO_4^{3-})	Na_3PO_4, K_3PO_4, $(NH_4)_3PO_4$	no exceptions
most sulfides (S^{2-})	Na_2S, K_2S, $(NH_4)_2S$	no exceptions

*NH_4OH does not exist in significant amounts in an ammonia solution. Ammonium and hydroxide ions readily combine to form ammonia and water.

**AgOH readily decomposes to form a precipitate of silver oxide and water.

APPENDIX G Average bond enthalpies at 298 K

TABLE 1 Bond energies and average bond energies

Single bonds (kJ mol^{-1})							
C–H	413	N–H	391	O–H	463	F–F	159
C–C	346	N–C	286	O–O	144		
C–N	286	N–N	158	O–F	191	Cl–F	255
C–O	358	N–O	214	O–Cl	206	Cl–Cl	242
C–F	492	N–F	278	O–I	201		
C–Cl	324	N–Cl	192			Br–F	249
C–Br	285	N–Br	243	S–H	364	Br–Cl	219
C–I	228			S–F	327	Br–Br	193
C–S	289	H–H	436	S–Cl	271		
		H–F	567	S–Br	218	I–Cl	211
Si–H	323	H–Cl	431	S–S	266	I–Br	178
Si–Si	226	H–Br	366			I–I	151
Si–C	307	H–I	298				
Si–O	466						
Multiple bonds Δ H (kJ mol^{-1})							
C=C	602	N=N	418	O=O	494		
C≡C	835	N≡N	942				
C=N	615			S=O	523		
C≡N	887			S=S	429		
C=O	799						

TABLE 2 Standard enthalpies of formation and standard entropies

Substance	Formula	ΔH°_f (kJ mol^{-1})
ammonia	$NH_3(g)$	–45.9
butane	$C_4H_{10}(g)$	–126
carbon dioxide	$CO_2(g)$	–393.5
carbon monoxide	$CO(g)$	–110.5
ethane	$C_2H_6(g)$	–84.0
ethanol	$C_2H_5OH(l)$	–277.7
ethene	$C_2H_4(g)$	+52.0
ethyne (acetylene)	$C_2H_2(g)$	+228
glucose	$C_6H_{12}O_6(s)$	–1271
hydrogen	$H_2(g)$	0
hydrogen bromide	$HBr(g)$	–36.3
hydrogen chloride	$HCl(g)$	–92.3
hydrogen fluoride	$HF(g)$	–273.3
hydrogen iodide	$HI(g)$	+26.5
methane	$CH_4(g)$	–74.8
nitrogen	$N_2(g)$	0
nitrogen dioxide	$NO_2(g)$	+33.1
nitrogen monoxide	$NO(g)$	+90.2
octane	$C_8H_{18}(l)$	–250.3
oxygen	$O_2(g)$	0
propane	$C_3H_8(g)$	–104.6
water(g)	$H_2O(g)$	–241.8
water(l)	$H_2O(l)$	–285.8

TABLE 3 Enthalpies of combustion*

Substance	Formula	ΔH°_c (kJ mol^{-1})
butane	$C_4H_{10}(g)$	–2886
carbon (graphite)	$C(s)$	–394
ethane	$C_2H_6(g)$	–1560
ethene	$C_2H_4(g)$	–1411
ethanol	C_2H_5OH	–1367
glucose	$C_6H_{12}O_6(s)$	–2803
hydrogen	$H_2(g)$	–286
methane	$CH_4(g)$	–890
methanol	$CH_3OH(l)$	–725
octane	$C_8H_{18}(l)$	–5450
propane	$C_3H_8(g)$	–2220

*Values are measured at standard state conditions (298 K, 1 bar).

Answers

Chapter 1 Chemistry Skills and Assessment toolkit

1.1 Chemical science

1.1 Key questions

Retrieval

1 Chemistry is a branch of science that studies the composition, structure, properties and change of matter. This includes the properties of individual atoms, how atoms form chemical bonds to create chemical compounds, the interactions between substances through intermolecular forces that are closely related to the general properties, and the chemical reactions that form different substances.

Comprehension

2 Student answers will vary.

Part A Working scientifically

1.2 Orders of magnitude

SkillBuilder 1.2.1 121 000 = 1.21×10^5

TY 1.2.1 4.14×10^6 chloride ions

TY 1.2.2 **a** 0.000 000 000 144 m
b 602 000 000 000 000 000 000 000 molecules

TY 1.2.3 0.09 L h^{-1}

1.2 Key questions

Retrieval

1 Writing numbers in scientific notation keeps the number in a consistent format, while writing numbers out in long form would be very tedious for very large and very small numbers.
2 between 1 (inclusive) and 10 (exclusive)
3 Tg is teragrams, so 10^{12} g

Comprehension

4 **a** 6.37×10^{-9} m **b** 3.19×10^2 s
5 **a** 0.000 000 000 11 m **b** 7 031 000 ms^{-2}
6 **a** 10 **b** –3
7

Number	Scientific notation
0.002	2×10^{-3}
2000	2×10^3
0.1234	1.234×10^{-1}
12.34	1.234×10^1
123.4	1.234×10^2

Analysis

8 **a** 1.56×10^3 μm
b 8.35×10^{-7} cm
9 human hair: 50×10^{-6} m (or 5×10^{-5} m)
virus: 2×10^{-8} m
So the ratio of the width of the hair to the virus
$= \frac{5\times10^{-5}}{2\times10^{-8}}$
$= 2500$
The hair is 2500 times wider.

10

Shape	Surface area (cm^2)	Volume (cm^3)	Surface area : volume (cm^{-1})
cube (3 cm)	54	27	2
sphere (radius 1.5 cm)	28.3	14.1	2
cylinder with closed ends (radius 1.2 cm; height 2 cm)	24.1	9.0	2.7

1.3 Mathematical basics for chemistry

TY 1.3.1 **a** 1.4×10^{-31} m^3 **b** 1.3×10^{-20} m^2

TY 1.3.2 D

1.3 Key questions

Retrieval

1 gradient $= \frac{\text{change in } y}{\text{change in } x}$

Comprehension

2 **a** $75 \times 0.789 = 59.175 = 59$ g **b** $\frac{24.5}{1.798} = 31.05$ mL

Analysis

3 **a** average change is the gradient from 0 min to 10 min

$$\text{gradient} = \frac{\text{change in temp(min10 – min 0)}}{10 - 0}$$
$$= \frac{78 - 22}{10}$$
$$= 5.6°\text{C min}^{-1}$$

b You can plot a graph of the results and see if the points form a straight line. You can then calculate the gradient for each time interval, and if the gradient is the same for each, then the temperature increase is constant. Alternatively, work out the expected temperature based on an increase of 5.6°C per min. Remember though that due to the precision of the thermometer, the number will not be exact.

Time (min)	Temperature (°C)	Temperature if 5.6°C per min
0	22	22
2	34	33.2
4	43	44.4
5	50	50
8	64	66.8
10	78	78

As you can see, the predicted temperature based on a constant rise is not the same as the measured temperature, so the increase in temperature is not constant.

1.4 Units

TY 1.4.1 3.0×10^{27} carbon atoms

1.4 Key questions

Retrieval

1

	Quantity	Standard unit of measurement	Symbol
a	time	second	s
b	thermodynamic temperature	kelvin	K
c	length	metre	m
d	mass	kilogram	kg

2 nano, micro, milli, centi, deca, kilo, mega

Comprehension

3 Measurements are made by general members of society, such as doctors, cooks, motorists, pilots and workers in the financial sector. It is important that everybody who measures the same amount agrees on a system of units so that its value is agreed upon and can be transferred globally.

Analysis

4 0.03000 L (3.000×10^{-2} L)
5 **a** 23 mL g^{-1} h^{-1} **b** 79 kL s^{-1}
c It is correct as written but could be also written as 4 L.
d 1.56 g
6 656 nm = 656×10^{-9} m = 6.56×10^{-7} m
7 8×10^{23} atoms

1.5 Uncertainties in measurement and error

TY 1.5.1 52 ± 5.5 s **TY 1.5.2** 11.4 ± 5% **TY 1.5.3** C

TY 1.5.4 **a** 23°C, 2 s.f. **b** 98.6°C, 3 s.f.

TY 1.5.5 **a** 8.5 ± 0.5 cm **b** 85 ± 0.5 mL **c** 58 ± 1 g

1.5 Key questions

Retrieval

1 Uncertainties and errors are typically written as an absolute value or percentage.
2 **a** ± 1.0 cm **b** ± 1.0° **c** ± 0.001 g **d** ± 5 volts
(The limit of reading is 10 volts but the absolute uncertainty is exactly half of this because there is only one reading and that is where the needle is pointing.)

Comprehension

3 The accuracy of a measurement is how close it is to an accepted value, while the precision of a measurement is how many significant figures can be accurately written down describing the measurement.
4 The limit of reading of a stopwatch is the smallest division, which is 0.01 s, but the user of the stopwatch could not possibly be that precise in reaction time when timing a given event. So the figure of ±0.005 s is far too small to be used as the absolute uncertainty.

Analysis

5 **a** average = 11.2
uncertainty = 0.6
Overall, this is written as 11.2 ± 0.6.
b average = 25.4
Uncertainty = 3.2
Overall, this is written as 25.4 ± 3.2.
6 **a** ±0.1
b ±0.05 mL
c ±0.001 g

1.6 Tables and graphing

1.6 Key questions

Retrieval

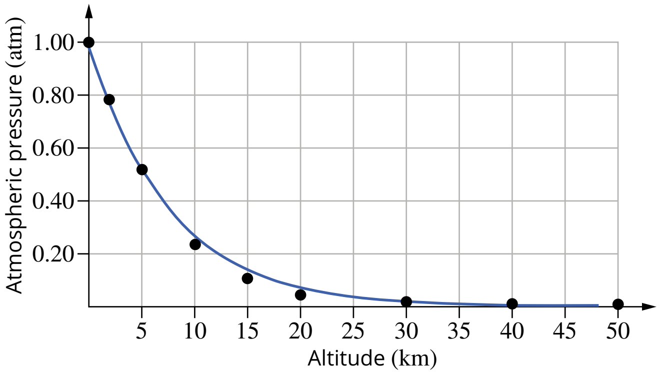
1 A trend line (line of best fit) is used to show the overall trend in the data, and can be used to predict values between the data points.

Comprehension

2 An outlier is a data point that does not fit the trend. While an outlier may be caused by an error, it is not always the case. It just represents data that is not consistent with the rest of the data from the set.
3 A sloping linear graph shows that there is a constant change in the dependent variable as the independent variable changes.
4 an inverse relationship
5 A line of best fit can be drawn through the points.
6 a positive relationship

Analysis

7 Antonia should check her data, especially the uncertainties as they are probably too small. Increasing the uncertainties increases the size of the error bars, which will make the straight line of best fit easier to draw.
8 The best fit line should either be passing directly through the origin, or the maximum and minimum lines pass below and above the origin respectively.
9 If the data is continuous and a relationship exists between the variables on the *x* and *y* axes, then a best fit line should be drawn. Dot-to-dot graphs may only be used with discrete data where there is no meaning of the values between the points, e.g. ionisation energy versus atomic number graphs.
10 **A graph of atmospheric pressure against altitude**



As the altitude increases, the atmospheric pressure decreases exponentially.

1.7 Statistics

1.7 Key questions

Retrieval

1 The mean is the average of a data set.
The median is the 'middle' value in an ordered list of values.
The mode is the value that occurs most often in a data set.

Comprehension

2 The standard deviation is a measure of the extent to which the individual data points within a data set differ from the mean value.

Analysis

3 **a** 9 (28 – 19) **b** 21 **c** 19
d 22 (22.29 but due to the precision of the values only 22 can be presented)
e 3
4 mean

Part B Student experiment

1.8 Research and planning

1.8 Key questions

Retrieval

1 **a** Variables are the factors that change during an experiment and which can influence results.

b

Independent variable	Controlled variables	Dependent variable
The experimental variable—the factor that is being changed.	Factors that you want to keep the same for every trial to ensure a fair test.	The factor that you are measuring to get results; it is dependent on your independent variable.

Comprehension

2 Accuracy is the ability to obtain the correct measurement, whereas precision is the ability to consistently obtain the same measurement.

3 Quantitative data relates to numeric variables that can be counted or measured (quantified) and recorded with units (e.g. centimetres, grams, kilometres, seconds). Qualitative data relates to categorical variables and can be counted or described but not measured (e.g. dog breed, plant genus or taste of fruit). Observations of categorical variables can be recorded as descriptions or images.

Analysis

4 **a** pH of the solution
  b mass of the mussel shells
  c weigh the mussel shells
  d volume of the solution, time that the mussel shells were left in the solution, temperature of the solution

5 Measure the time needed to produce 25 mL of hydrogen gas (dependent variable) with five different concentrations of KOH (0.2 M, 0.4 M, 0.6M, 0.8 M and 1.0 M) (independent variable). Each concentration will be tested three times.

1.9 Conducting and experimenting

1.9 Key questions

Retrieval

1 The choice of equipment and instruments will influence the reliability of the experiment.

Comprehension

2 There is less chance of error when using higher-precision instruments. This lowers the uncertainty of the measurements, which makes scientific conclusions more accurate.

Analysis

3 Some of the data that was collected was not relevant. The independent variable is the temperature and the dependent variable is the time to collect 25 mL of hydrogen. The other variables need to be constant; therefore experiments 3 and 4, where the concentration of hydrochloric acid is changed, and experiment number 5 where the starting mass of zinc is different from experiments 1 and 2. This means that there are only two valid experiments that show the variation of time to produce 25 mL of hydrogen gas and change in temperature. This is not a sufficient number of experiments to produce a trend.

In addition, each point should be repeated at least three times and an average calculated for those points; however, in the data only a single experiment is conducted at each temperature.

1.10 Results

1.10 Key questions

Retrieval

1 raw and processed data
2 a constant bias in measurement
3 errors that follows no regular pattern
4 The response must demonstrate the:
- correct and relevant processing of data
- thorough identification of relevant trends, patterns or relationships
- thorough and appropriate identification of the uncertainty and limitations of the evidence
- collection of sufficient and relevant raw data.

Comprehension

5 Processing data is the manipulation of the numbers or observations to understand and ascertain true values, uncertainties, errors, anomalies and relationships. Interpreting data is placing this understanding into words and providing an in-depth explanation of what the numbers mean.

6 **a** mistake **b** random **c** systematic

Analysis

7 It cannot be concluded that the volume of hydrogen gas produced between the reaction of zinc and ethanoic acid is lower. Even though the mean value with ethanoic acid is lower than the mean value with the reaction with hydrochloric acid, the uncertainty of ±2 indicates that both values are within each other's range, i.e. 87 and 86 mL are within the ranges of the uncertainties with the different acids. Hence, there is no significant evidence to suggest that any less hydrogen gas is produced through the reaction of zinc with ethanoic acid.

1.11 Communicating and writing a scientific report

1.11 Key questions

Retrieval

1 **a** The method outlines the exact details for other scientists to repeat the experiment including specific details about instruments, equipment models and precision, techniques employed and all information required for others to repeat the same results.
  It is not common to differentiate the materials (equipment and instruments used) from the procedure that uses them.
  b The conclusion provides an explanation of the results including quality of experimentation (accuracy, precision, validity and reliability of the methodology) and relates the results to current scientific understanding (theory).
  The strength of the relationship between the experimental evidence (data, observations or results) is to be stated.
  No new data is to be included in this section.

2 Results

Comprehension

3 A bibliography is a list of all the sources used during the research to develop understanding (including information in the journal), even if the information was not used directly or explicitly in the scientific report. A reference list only lists the sources used and cited (or in-text referenced) in the scientific report.

Analysis

4 **a** A **b** D **c** D
  d There is no relationship between the dependent variable and the independent variable as the graph is straight and horizontal. This means that the dependent variable is the same value regardless of the value of the independent variable.

5 time constraints, limited resources such as equipment availability or equipment range

Part C Research investigation

1.12 Developing the research question from a claim

1.12 Key questions

Retrieval

1 the dependent and independent variables

Comprehension

2 Knowledge and understanding about the variables is developed and therefore a more specific or appropriate question can be written.

3 The ISMG requires a rationale for the development of the research question from the claim, and this is achieved by applying your understanding of scientific concepts to the development which is from research.

Analysis

4 **a** Are biofuels a greener and a better alternative to oil-based fuels, based on their mode of production from renewable sources and comparable energy content of the fuel?

b Does the more effective method of drug delivery offered by nanocups mean that analgesics can be delivered safely, cheaply and with fewer side effects to the source of pain and infection?

1.13 Finding and choosing suitable resources

1.13 Key questions

Retrieval

1 **a** information created by the person or persons directly involved in a study, investigation or experiment or observing an event

b data collected by a person or group other than the person or group using the data

Comprehension

2 **a** secondary

b primary

c primary

d secondary

Analysis

3 The NASA website would be the best resource to use. The article 'Batteries and fuel cells in space' has been published in a well-respected peer-reviewed journal and so the information will be accurate and related to the topic that you are learning about. However, technology is advancing rapidly, so the fact that it is a print resource from the twentieth century means that it is likely to be out of date. As a print resource cannot be changed once published, the information will be that from last century. The book on fuel cells is more recent but like the journal article, it is a print resource and so the information contained within has not been updated from the early twenty-first century. Future editions of the book may provide more current information. In addition, the book is not limited in its focus to applications related to space exploration and so may not contain information specific to your area of learning.

The NASA website is run by the leading institute into space exploration and so the information contained will be both highly accurate and directly related to the use of hydrogen fuel cells in space travel. In addition, although the original article was published in 2009, since it is a website, updated information or additional links may be available from the link.

1.14 Research: taking and organising notes

TY 1.14.1 Students' own answers but the following is a guide.

Thinking	Working
Can swap the order of information, the reason (low pressure system) with the object (cooler days and not so cool nights).	Cooler days and not so cool nights will be brought by a low pressure system.
Can change numerous words such as adjectives, verbs, adverbs and nouns.	Due to a low pressure system, following thereafter will be colder days and nights not as cool as usual.
Can elaborate on the information to explicitly outline explanations.	As a low pressure system passes (with atmospheric pressure below 1000 mbar), it will draw the air up into the atmosphere that has been warmed by the sun making the days cooler, and will draw heat from the land during the nights making them not so cool.

1.14 Key questions

Retrieval

1 summarised in your own words, using a variety of means, such as diagrams and written text, with full details, noting the entry date and using a title, noting the source used

Comprehension

2 A scientific journal is a comprehensive record of everything that you have done with respect to an investigation. It includes your methodology in such detail that it can be reproduced by you or someone else including the amount and nature of chemicals that you used. All your observations and thought processes that are relevant to the observations and results should be recorded in the journal. This would typically include explanations for conducting the experiments, back information and references. It is particularly important to record experiments even if they do not work as this will prevent you repeating the experiment and wasting time. Finally, it is important to understand that at a research laboratory, a laboratory journal is a legal document and dates can be critical in determining who discovered something first or issuing a patent or investigations in relation to fraud.

Analysis

3 Students' own answers

1.15 Writing a report for the research investigation

1.15 Key questions

Retrieval

1 There is no fixed structure to a research investigation report but it must:

- outline the purpose or focus of the research investigation
- identify concepts relevant to the research question
- discuss prior research that has been conducted relevant to the concepts or directly addressing the research question—this should include a critical analysis of the methodology, results and conclusions of the past studies
- evaluate the studies and any other information pertinent to the research question
- draw a justified conclusion about the research question, including sources of uncertainty and future direction that could provide greater clarity.

Comprehension

2 A research investigation critically analyses the information relevant to the research question and outlines the significance of the research question or area of investigation.

Analysis

3 Identify strengths and weaknesses in the methodology, samples and results by considering their validity and reliability and statistical uses. Establish what their results and statistics mean. Did the results in the previous research studies support one another or did findings from one study contradict the results from another? Finally, identify any gaps and critically evaluate the conflicting evidence in the studies that are related to the research question. For instance, were the differences in the methodology (including conditions such as temperature or measurement technique/equipment) or chemicals used the reason for the differences in the studies and results?

Chapter 2 Elements, compounds and mixtures

2.1 Characterising matter

2.1 Key questions

Retrieval

1 anything that occupies space, has mass and can be observed with our senses

2 **a** heterogeneous mixtures and homogeneous mixtures
b elements and compounds

3 A physical change in matter is a process where the form of matter may be changed without changing its chemical identity or its chemical composition. No new substances are formed during physical changes.

4 A pure substance (or simply a substance) is matter that has a definite and distinct set of physical and chemical properties that do not vary in composition from sample to sample. Pure substances can be classified as either elements or compounds.

5 **a** iron **b** potassium **c** silver **d** lead **e** mercury

6 A compound is composed of two or more elements chemically bonded in a fixed ratio. The two major classes of compounds are ionic compounds and molecular compounds.

7 molecular compounds; ionic compounds

8 Answers include but are not limited to:
a boiling point, viscosity, density, freezing point, refractive index
b melting point, density, colour, hardness, electrical conductivity, thermal conductivity
c boiling point, colour, density, freezing point

Comprehension

9 **a** freezing **b** condensation **c** sublimation **d** evaporation

10 silver-grey coloured metal = physical property; melts at 420°C = physical property; is placed in dilute sulfuric acid, hydrogen is given off and the metal dissolves = chemical property; density of 7.13 g cm^{-3} = physical property; reacts slowly with oxygen to form a metal oxide = chemical property

11 **a** heated in a reaction vessel to produce mercury metal and oxygen gas = chemical change
b splint burst into flame = chemical change

12 Elements and compounds are examples of pure substances whose composition does not vary from sample to sample.
Neither can be broken down into simpler substances by simple physical processes. Compounds can be broken down to simpler substances by chemical processes; elements cannot be broken down to simpler substances by chemical means. A mixture is a physical combination of pure substances whose composition can vary from sample to sample.
Mixtures can be separated into their constituent components by physical processes.

Analysis

13 **a** chemical **b** chemical **c** physical **d** physical

14 **a** physical change **b** chemical change **c** physical change
d chemical change

15 **a** element **b** mixture **c** compound **d** compound
e mixture **f** compound **g** element **h** mixture

16 approximately −2°C

17 **a** metal, Mg **b** metal, Mn **c** metal, Ag **d** metal, Hg
e non-metal, Ne **f** metalloid, As **g** non-metal, S
h metalloid, Si

18 Ionic compounds are composed of metals chemically bonded to non-metals, while molecular compounds are composed of non-metals chemically bonded to other non-metals.
a ionic **b** molecular **c** molecular **d** ionic

19 Elements in the same group (i.e. vertical column) of the periodic table tend to share similar physical and chemical properties. Potassium is a group 1 metal and therefore lithium (Li), sodium (Na), rubidium (Rb) and caesium (Cs) are all group 1 metals that would share similar physical and chemical properties with potassium. Note: even though hydrogen is a group 1 element, it is not a metal and it exists as a molecule in its elemental form. It does not share the same physical properties.

2.2 Homogeneous and heterogeneous mixtures

2.2 Key questions

Retrieval

1 **a** A phase is a region of matter that is physically and chemically uniform in terms of composition and properties. It is physically distinct from other regions of matter and is typically mechanically separable from other phases.
Heterogeneous mixtures contain two or more phases and, as such, are not uniform throughout. For homogeneous mixtures, there is only one visually distinguishable phase and the mixture has uniform properties throughout.
b A solution is a homogeneous mixture and therefore has only one phase. All solutions have only one phase.

2 A solution is a homogeneous mixture of two or more substances. Solutions are uniform throughout and a small sub-sample of a solution will have chemical and physical properties that are representative of the whole sample. Solutions are composed of a solvent and solutes; solutes are said to be dissolved in the solvent. The solvent is the substance present in the greatest amount in a homogeneous mixture; all other substances are solutes.

3 **a** heterogeneous mixture
b homogeneous mixture, element, compound
c heterogeneous mixture, different homogeneous mixtures can be composed of different matter but it is uniform within a sample

4 A homogeneous mixture involves one or more pure substances being dissolved in another substance; for example, sea water.
A heterogeneous mixture consists of two or more substances that have distinguishable phases; for example, granite and salad dressing.

Comprehension

5 Both types of mixtures are physical combinations of two or more pure substances. A heterogeneous mixture has two or more visually distinguishable phases with differing chemical and physical properties and is not uniform throughout—different sub-samples of a heterogeneous mixture will have different compositions. A homogeneous mixture (also known as a solution) has only one single phase and is uniform in composition throughout—different sub-samples of a homogeneous mixture will have identical compositions.

Analysis

6 **a** heterogeneous **b** heterogeneous **c** heterogeneous
d homogeneous **e** heterogeneous **f** homogeneous
g heterogeneous **h** homogeneous

7 The state of matter refers to matter existing in one of five possible physical states, i.e. solid, liquid or gas being the three most commonly encountered, with Plasma and Bose–Einstein

Condensate being the other two. A phase is a region of matter within a larger sample of matter that is physically distinct and has different physical and/or chemical properties from other parts of the sample. Phases are typically mechanically separable from other phases. We can refer to the solid phase or liquid phase of a heterogeneous mixture, e.g. in a sand–water mixture the solid sand would be called the solid phase and the liquid water would be called the liquid phase. It is also possible to have different phases of a mixture in the same state, e.g. the two liquid phases of an oil–water salad dressing. It is even possible to have different phases within the same state of a single pure substance, e.g. ice has up to 16 different phases depending on temperature and pressure, with different packing geometries of the water molecules imparting different physical properties on the different phases.

8 **a** Composition cannot vary. An element is an example of a pure substance and has a definite and distinct set of physical and chemical properties that does not vary in composition from sample to sample.
b Composition cannot vary. A compound is an example of a mixture of pure substances with a fixed ratio of components. It has a definite and distinct set of physical and chemical properties that does not vary in composition from sample to sample.
c Composition can vary. Even though any one single sample of a homogeneous mixture is uniform throughout, the relative proportions of components that make up the mixture can vary from sample to sample.
d Composition can vary. The composition of a heterogeneous mixture is not uniform and can vary within a single sample as well as from sample to sample.

9 A heterogeneous mixture has different phases with differing chemical and physical properties. These mixtures can vary in composition both within a sample (from one part of a sample to another) and can also vary in composition between different samples. Homogeneous mixtures (e.g. solutions) have a single phase and while they have a uniform composition within a particular sample they can vary in composition from sample to sample. For example consider a salt water solution made by dissolving 1 gram of sodium chloride in 10 litres of water compared to a solution prepared by dissolving 10 grams of sodium chloride in 10 litres of water. Both solutions have uniform composition within each sample but when compared to each other, their compositions vary—and so do their chemical and physical properties.

10 **a** heterogeneous mixture of ink solvent and solid carbon black particles
b an element with a solid phase and a gaseous phase in the same sample of matter
c heterogeneous mixture with a salt solution as the liquid phase and the crystals as the solid phase
d homogeneous mixture having a single blue liquid-solution phase
e heterogeneous mixture of at least two solid phases of silica crystals and calcite crystals

11 **a** element **b** element **c** compound
d homogeneous mixture (or solution)
e heterogeneous mixture of two types of solid crystals

12 **a** mixture (of copper and zinc) **b** mixture (of silver and copper)
c mixture (of copper and tin and other trace metals)
d element **e** element **f** element

2.3 Separating mixtures

2.3 Key questions

Retrieval

1 **a** boiling point **b** melting point **c** boiling point

2 **a** a heterogeneous mixture of solids where the particles of the different phases are of different sizes (e.g. separating rubbish from beach sand)
b a heterogeneous mixture of undissolved solids in a liquid (e.g. separating silt particles from a muddy water sample)
c heterogeneous mixtures where the different phases have different densities (e.g. separating sawdust from sand by adding water)
d homogeneous mixture of two substances with different boiling points
e either a homogeneous or heterogeneous mixture of two substances with different boiling points (e.g. recovering the solid particles from muddy water or the dissolved salt from sea water by evaporating the water in the mixture)

3 when you have a heterogeneous mixture of relatively low density liquid sitting atop another more dense liquid phase or sitting atop a solid phase

4 decantation and flotation

Comprehension

5 **a** take advantage of the differences in magnetic properties of the phases by using a magnet
b take advantage of the differences in boiling points of the two substances by using evaporation (of the water) or distillation if recovery of both the salt and water is required
c take advantage of the differences in boiling points of the two substances by distillation if recovery of both the water and ethanol was required

6 when recovery of both the solid salt and liquid water components is required

7 **a** and v **b** and iii **c** and **ii** **d** and iv
e and vi **f** and i **g** and vii

8 Distillation is typically used for separating simple solutions containing, for example, two components which have relatively large differences in boiling points. Fractional distillation is best used for complex mixtures of liquid substances that have relatively small differences in boiling points. In terms of the distillation of crude oil, the main difference between conventional distillation and fractional distillation is the presence of a large fractionation column (or tower) above the heated mixture, which has several condensers coming off at different heights. As the mixture is vaporised it rises up the column and cools down with increasing height. Different components of the crude oil will condense at different temperatures, and therefore at different heights, allowing the different fractions to be collected.

Analysis

9 **a** sieving **b** decanting
c filtration through a colander
d filtration using a laboratory filtration apparatus
e fractional distillation **f** distillation

10 The evaporation technique involves heating the solution with the purpose of boiling off the liquid solvent to leave behind (or recover) the dissolved solid salt. In this method the liquid solvent (i.e. water) would be lost. In distillation the aim is also to boil off the liquid solvent to isolate the solid salt. At the same time, however, the vaporised solvent is recondensed to the liquid state to allow recovery of both the solvent and the solute. Distillation would be used if it was necessary to separate and recover both the dissolved solid and liquid components. The evaporation technique would only be used if it was necessary to recover only the dissolved salt.

Chapter 2 Review

Retrieval

1 D

2 **a** physical property **b** physical property
c chemical property **d** chemical property
e physical property

3 chemical change; physical change

4 Vertical columns in the periodic table are called groups. Elements are arranged in groups according to similarities in physical and chemical properties. Elements that belong to the same group share similar physical and chemical properties because they have similar outer-shell electron configurations.

Comprehension

5 c and **d**

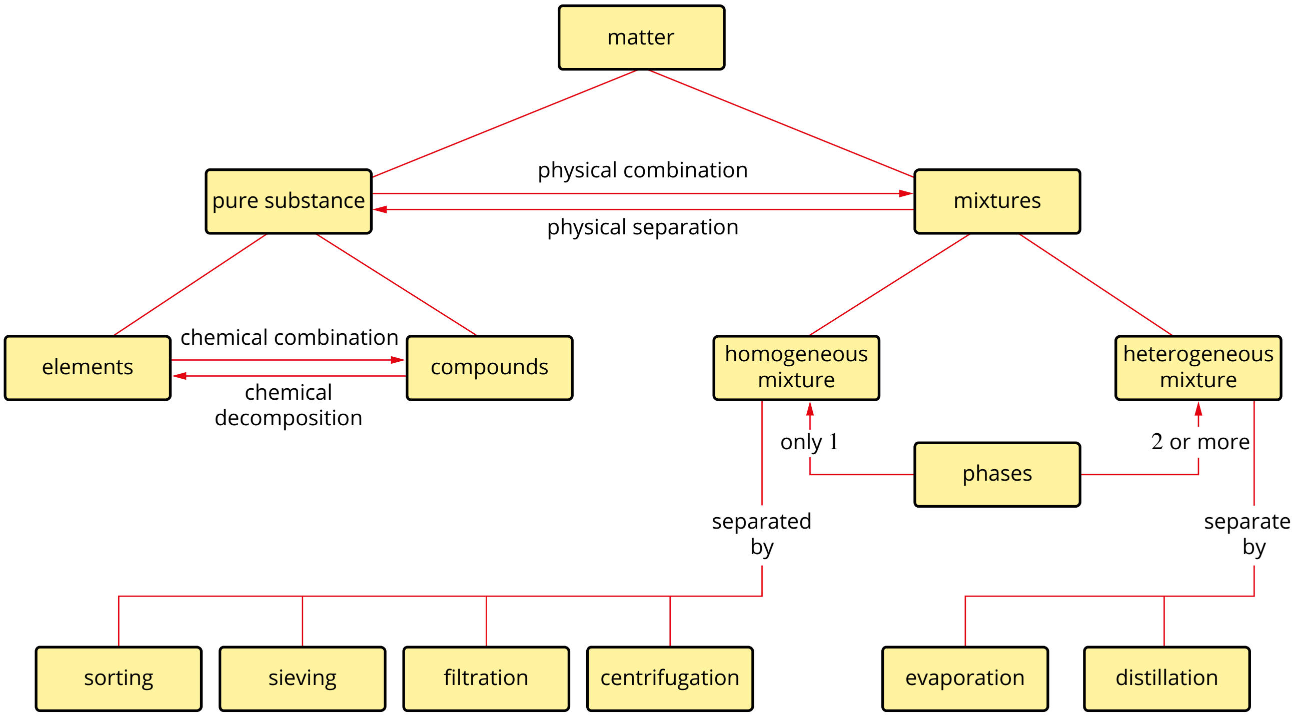
6 **a** a heterogeneous mixture likely comprising iron oxides, iron sulfides and other minerals
- **b** a pure substance (although even the purest sample of copper wire is likely to contain trace levels of impurities introduced during the refining process)
- **c** a heterogeneous mixture of water and sand particles—the latter comprising of quartz, calcite and other minerals
- **d** a pure substance (although, again, even the purest sample may contain trace impurities introduced during the distillation process)

7 **a** pure substance; liquid alcohol phase, gaseous alcohol phase
- **b** heterogeneous mixture; liquid solution and solid pigment particles
- **c** pure substance; solid copper phase and liquid copper phase
- **d** heterogeneous mixture; solid quartz phase and solid calcite phase

8 Centrifugation and filtration would be two simple physical processes that could be used to rapidly separate undissolved solids from a liquid phase. Depending on the type of centrifuge employed (e.g. how rapidly it spins and the g-forces produced) the centrifugation method may leave a small amount of suspended particles in the water layer. Filtration using a well-defined membrane filter will remove all particles greater than the defined pore size. For example, a 0.45 μm membrane filter is generally used to separate and measure the amount of total suspended solids (TSS) in a water sample. However, colloidal particles less than 0.45 μm in size may still be retained in the water phase.
As with any separation process, the choice of separation technique comes down to the purpose of performing the separation in the first place.

9 Distillation would be used if you wanted to separate and recover both the dissolved solid and liquid components. The evaporation technique would lose the liquid component and leave behind only the solid residue. Distillation would recover both components.

10



Analysis

11 **a** compound **b** compound
- **c** insufficient information to classify as an element or compound
- **d** insufficient information to classify as an element or compound

12 **a** physical property **b** physical property
- **c** chemical property **d** chemical property
- **e** physical property **f** physical property
- **g** chemical property

13 **a** physical change **b** chemical change
- **c** physical change **d** chemical change

14 **a** chemical change **b** physical change
- **c** physical change **d** chemical change
- **e** chemical change **f** chemical change

15 Both are types of pure substances and have uniform physical and chemical properties that do not vary from sample to sample. Neither can be broken down to simpler substances by simple physical processes. Compounds can be broken down to simpler substances via chemical reactions but elements cannot. Elements are the simplest form of matter that exist and chemically combine (via chemical reactions) to produce compounds. Elements consist of a single type of atom; compounds are composed of two or more elements combined in a fixed ratio and therefore contain different types of atoms.

16 **a** hydrogen (symbol: H, atomic number: 1) and oxygen (symbol: O, atomic number: 8)
- **b** nitrogen (symbol: N, atomic number: 7) and hydrogen (symbol: H, atomic number: 1)
- **c** carbon (symbol: C, atomic number: 6) and hydrogen (symbol: H, atomic number: 1)
- **d** nitrogen (symbol: N, atomic number: 7) and oxygen (symbol: O, atomic number: 8)
- **e** sulfur (symbol: S, atomic number: 16) and fluorine (symbol: F, atomic number: 9)

17 All of the elements are non-metals. The compounds listed in Question 16 are molecular compounds as they are non-metals bonded to other non-metals.

18 **a** sodium (symbol: Na, atomic number: 11) and chlorine (symbol: Cl, atomic number: 17)

- **b** calcium (symbol: Ca, atomic number: 20) and fluorine (symbol: F, atomic number: 9)
- **c** aluminium (symbol: Al, atomic number: 13) and oxygen (symbol: O, atomic number: 8)
- **d** copper (symbol: Cu, atomic number: 29), sulfur (symbol: S, atomic number 16) and oxygen (symbol: O, atomic number: 8)
- **e** iron (symbol: Fe, atomic number: 26), carbon (symbol: C, atomic number: 6) and oxygen (symbol: O, atomic number: 8)

19 The elements in each compound are composed of metals bonded to non-metals.

20 **a** chemical change **b** chemical change **c** physical change **d** physical change **e** physical change

21 **a** element
- **b** heterogeneous mixture, homogeneous mixture
- **c** compound

22 **a** heterogeneous mixture **b** heterogeneous mixture **c** heterogeneous mixture **d** homogeneous mixture
- **e** pure substance (i.e. an element)
- **f** pure substance (i.e. a compound)
- **g** homogeneous mixture (otherwise known as a solution)

23 **a** The solvent is nitrogen gas; solutes include oxygen gas and a range of trace gases including carbon dioxide, water vapour and argon.
- **b** The solvent is water; solutes include dissolved salts such as sodium chloride (and trace smaller amounts of other salts) plus dissolved gases including oxygen and nitrogen.
- **c** The solvent is water; the main solute is ethanoic acid.
- **d** The solvent is water; the main solute is ethanol.
- **e** The solvent is water, the main solutes are dissolved gases such as oxygen and nitrogen as well as dissolved salts.

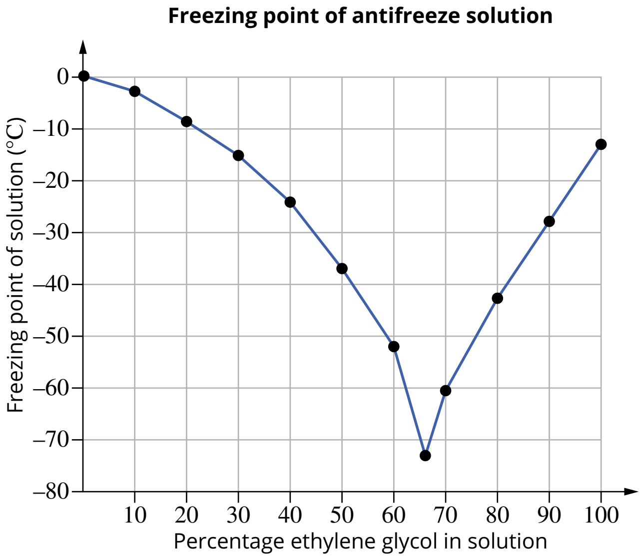
24 Sand. When sand is mixed with water a heterogeneous mixture will be formed with the sand settling at the bottom of the water sample. This mixture could easily be separated using decantation by pouring off the overlaying liquid layer.
Filtration could also be used. All of the other solids will dissolve in water to form a homogeneous solution. This would involve more complex separation techniques such as evaporation to recover the solid component or distillation to recover both the solid and liquid components.

Knowledge utilisation

25 Pure water will have different physical properties to a mixture containing trace impurities of dissolved salts. Some physical properties that will be different (and easily measured) include electrical conductivity, boiling point and freezing point. Even trace amounts of dissolved salts will greatly increase the electrical conductivity of water.
Freezing point depression and boiling point elevation will also be observed for solutions containing dissolved salts so measuring the boiling point and/or freezing point would be another approach that could be used.

26 Sterling silver jewellery is a homogeneous mixture of silver and copper metals. The mixture will have a different melting point compared to a sample of pure silver. Measuring the melting points of both samples can therefore be used to tell which is which.

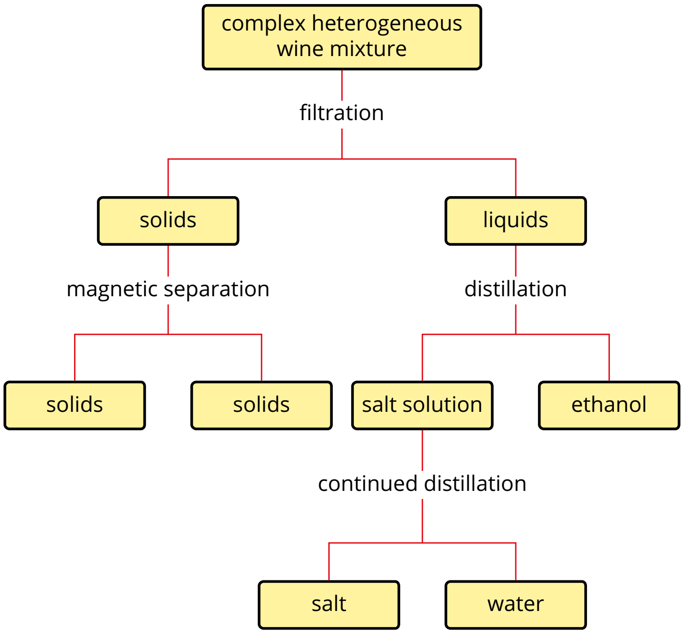
27 **a** approximately 40% ethylene glycol solution
- **b** approximately 20% ethylene glycol solution



28 **a** This is a heterogeneous mixture of two immiscible liquids. The less dense oil can be decanted from the top of the more dense water layer. A separating funnel could also be used to drain off the bottom water layer.
- **b** This is a heterogeneous mixture of solid particles with differing physical characteristics, most notably size. Sieving the mixture would be the easiest and most appropriate method for separating this mixture.

29 Sugar is readily dissolved in water, whereas sand is not. The mixture could be mixed with water to dissolve the sugar, leaving a heterogeneous mixture of sand within a sugar solution. The sand could then be filtered, leaving the sugar solution remaining. The sugar solution could then be separated using evaporation to boil off the liquid water, leaving the solid sugar remaining. In this way both solids will be recovered.

30



Chapter 3 The atomic world

3.1 Nanomaterials and nanoparticles

TY 3.1.1 2.43×10^6

3.1 Science as a Human Endeavour

Nanomaterials: Australian Institute for Biotechnology and Nanotechnology

1 **a** A drug could be adsorbed to the outside of a particle OR it could be absorbed to the interior of the particle (for example, to the interior of a carbon nanotube). In both situations, the drug is associated with the surface, but not bonded to it because the carbon nanotube is inert.

b Viruses have a variety of shapes. Tobacco mosaic virus, and the influenza virus (and others) have a cylindrical shape with a diameter of about 18 nm. This is a similar size and shape to a carbon nanotube, and so cell surface receptors may take up the nanotube in the same way as the virus particle is internalised.

2 Carbon nanoparticles are flexible, strong, conductive of electricity, durable and inert. Because they are inert (will not react chemically) they are unlikely to damage body tissues; also, their size (up to 100 nm) is much smaller than body cells (a red blood cell is 6000 nm diameter). This means that they are unlikely to impede blood flow.

However, the fact that these particles are very durable means that they are unlikely to be metabolised (broken down) by the body and many accumulate in cells or tissues when used over the long term. This may physically stress cells and cause cell or tissue death.

3.1 Key questions

Retrieval

1 8.35×10^7 nm

2 **a** 1.35×10^7 nm **b** 4.2×10^6 nm

Comprehension

3 250 sheets **4** 2500 times larger

5 The height of the figure is approximately seven times the width of the nanotube. Nanotubes are about 1 nm wide. So, the figure is about 7 nm tall, which is 7×10^{-9} metres.

6 This statement is incorrect because when objects are very small in size a greater proportion of the material making up the volume is found on the surface—there is less inner mass. Nanoparticles have a high surface area to volume ratio.

7 There are two reasons why their size makes nanoparticles useful for transporting medicine. First, their small size makes it easy for the nanoparticles to get into the body by inhalation, injection or absorption through the skin. Second, their small size gives nanoparticles a large surface area compared to their volume. This means that many molecules can be adsorbed onto the surface of a relatively small volume of nanoparticles.

Analysis

8 The size of the flea is 2 mm long and 1 mm wide. This means that it could not hide behind any of the structures in Figure 3.1.1.

9 **a**

Shape	Surface area	Volume	Surface area to volume ratio
cube 2 cm × 2 cm × 2 cm	24 cm^2	8 cm^3	3 : 1
sphere radius 1.38 cm	23.9 cm^2	11 cm^3	2.17 : 1
tube (cylinder) radius 1 cm height 2 cm	19 cm^2	6.3 cm^3	3.02 : 1

b The tube (cylinder) has the greatest with a SA : V ratio of 3.02.

c The sphere would hold the most volume.

d A large SA : V ratio would be good for rapid cooling, while a small SA : V ratio would be good for conservation of heat.

10 Absorbing water means that the molecules enter the nanoparticle, whereas adsorbed water would be attached to the surface of the substance or nanoparticle. A change in colour is an indication that a chemical change or reaction has occurred as a result of water entering the particle. Adsorbed water molecules are attached on the surface and do not have the same opportunity to react with the substance or nanoparticle.

3.2 Inside the atom

TY 3.2.1 2,8,3

TY 3.2.2 2,7 **TY 3.2.3** 5

3.2 Key questions

Retrieval

1 10 000–100 000 times larger

2 Protons and neutrons make up most of the mass of an atom and they are found in the nucleus.

3 The negative electrons are attracted to the positive protons through electrostatic attraction, which pulls the electrons in towards the nucleus.

4 12 neutrons

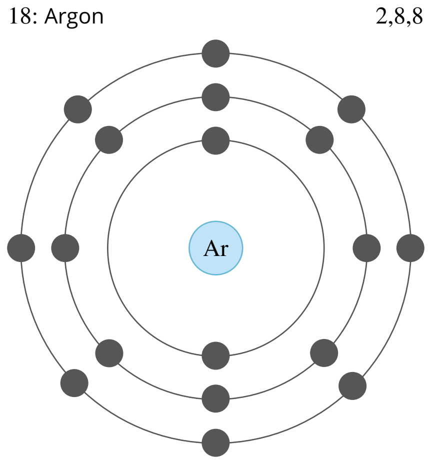
5 **a** 2 electrons in the first shell, 3 electrons in the second shell
b 2 electrons in the first shell, 8 electrons in the second shell and 2 electrons in the third shell
c 2 electrons in the first shell, 8 electrons in the second shell and 10 electrons in the third shell
d 2 electrons in the first shell, 8 electrons in the second shell, 18 electrons in the third shell and 7 electrons in the fourth shell

6 **a** helium, He **b** fluorine, F **c** aluminium, Al
d nitrogen, N **e** chlorine, Cl

7 **a** O **b** Mg **c** Ar d Be

Comprehension

8



9 The atom has two electrons in the first shell, eight electrons in the second shell and eight electrons in the third shell. Therefore, the number of valence electrons is 8.

10 **a** 2,2 **b** 2,8,6 **c** 2,8,8 **d** 2,8,2 **e** 2,8,18,3

11 **a** He **b** F **c** Al **d** N **e** Cl

12 **a** 2,8,10 **b** 2,8,18,10 **c** 2,8,1 **d** 2,8,7

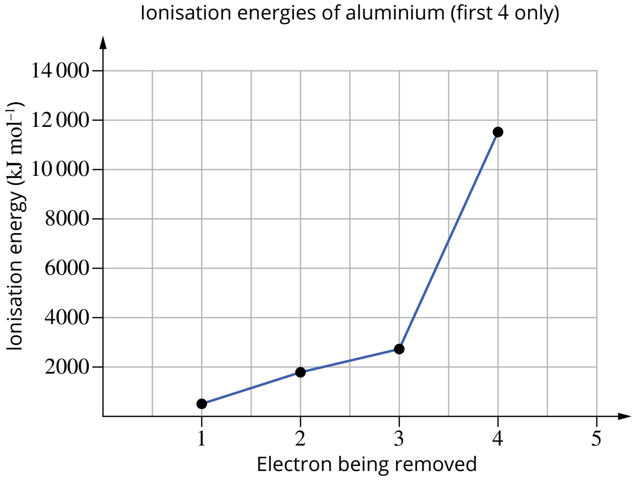
Analysis

13 Be—two electrons in the first shell, two electrons in the second shell
Mg—two electrons in the first shell, eight electrons in the second shell and two electrons in the third shell.
Another similarity is that these elements are in the same group of the periodic table (group 2) and have similar chemical properties.

14 72 electrons

15 The atom is sulfur because it has 16 electrons. This electron arrangement is unexpected because the second shell is not filled and electrons have been placed in the third shell. The rules for electron configuration are that the lower energy shells are filled first; however, the second shell (lower energy) has not been filled. A possible reason for this is that this atom has been excited by an energy source. Two electrons have 'jumped' from the second shell to the third shell. It is expected that, moments after this, the two electrons will return to the second shell and, at the same time, energy in the form of heat or light will be released from the atom.

16 **a**



b Aluminium has 13 electrons and its electron configuration is 2,8,3. The three electrons in the outer shell, $n = 3$, will be more loosely held than the inner electrons in $n = 2$ and $n = 1$ shells. The data provided shows this as there is an increase of about 1000 kJ mol^{-1} for the first three ionisation energies. However, there is a greater difference between the third and the fourth ionisation energies (8832 kJ mol^{-1}), indicating that the fourth electron (in shell $n = 2$) is much more tightly held.

3.3 Using notation to describe atoms

TY 3.3.1 92 protons, 143 neutrons, 92 electrons **TY 3.3.2** 5

3.3 Science as a Human Endeavour

Radioisotopes: nuclear medicine

1 The radioisotope should have a low energy (such as beta radiation) and should have a short half-life of hours or days. This means that damage to body cells is minimised.

2 Various answers are possible.

3 Various answers are possible.

3.3 Key questions

Retrieval

1 Z = 15, protons = 15, neutrons = 31 – 15 = 16, electrons = 15

2 nitrogen, ${}^{14}_{7}\mathrm{N}$

3 **a** potassium, ${}^{39}_{19}\mathrm{K}$ **b** nickel, ${}^{59}_{28}\mathrm{Ni}$ **c** boron, ${}^{11}_{5}\mathrm{B}$
d lithium, ${}^{7}_{3}\mathrm{Li}$

4 **a** Al, 14 neutrons **b** F, 10 neutrons **c** Mn, 30 neutrons
d He, 2 neutrons

Comprehension

5 ${}^{17}_{8}\mathrm{O}$ ${}^{18}_{8}\mathrm{O}$

6 Isotopes of the same element have the same atomic number and therefore the same number of protons and electrons. The number of neutrons is different between isotopes of an element; therefore, they have different mass numbers.

7

Element	Nuclear symbol notation	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
hydrogen	${}^{1}_{1}\mathrm{H}$	1	1	1	0	1
hydrogen	${}^{2}_{1}\mathrm{H}$	1	2	1	1	1
fluorine	${}^{19}_{9}\mathrm{F}$	9	19	9	10	9
sulfur	${}^{33}_{16}\mathrm{S}$	16	33	16	17	16
sulfur	${}^{34}_{16}\mathrm{S}$	16	34	16	18	16
copper	${}^{65}_{29}\mathrm{Cu}$	29	65	29	36	29
oxygen	${}^{16}_{8}\mathrm{O}$	8	16	8	8	8
copper	${}^{63}_{29}\mathrm{Cu}$	29	63	29	34	29

Analysis

8 There are six isotopes in the table in Question **7**:

${}^{1}_{1}\mathrm{H}$ and ${}^{2}_{1}\mathrm{H}$; ${}^{33}_{16}\mathrm{S}$ and ${}^{34}_{16}\mathrm{S}$; ${}^{65}_{29}\mathrm{Cu}$ and ${}^{63}_{29}\mathrm{Cu}$.

These are identified as isotopes because each pair of elements has the same A value (number of protons) but a different Z value (mass number), therefore they must have a different number of neutrons.

9 **a** 91 **b** 40 **c** 31 **d** 40
Both calcium and argon have the same mass number, Z = 40. This means that the total number of protons and number of neutrons is 40 for both atoms. However, this does not mean that their atoms are identical because their atomic numbers show that they have a different number of protons (argon has 18 protons and calcium has 20 protons). The reason that their mass numbers are the same is that Ca has 20 neutrons and Ar has 22 neutrons.

10 Boron and carbon have the same number of neutrons—they both have six neutrons. Also, fluorine and neon have the same number of neutrons—10 neutrons. This is possible because the number of neutrons for an element is the number required to stabilise the nucleus and shield the protons from each other. This can sometimes be the same number for different elements. However, different elements will never have the same number of protons because this number (atomic number) is the characteristic that defines the element.

3.4 Electronic structure of the atom

TY 3.4.1 The aluminium atom has 13 electrons. It has two electrons in the 1*s*-subshell, two electrons in the 2*s*-subshell, six electrons in the 2*p*-subshell, two electrons in the 3*s*-subshell and one electron in the 3*p*-subshell.

TY 3.4.2 $1s^22s^22p^63s^23p^63d^34s^2$

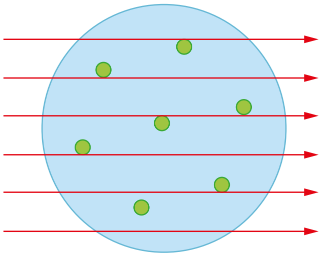
TY 3.4.3 $1s^22s^22p^63s^23p^63d^{10}4s^24p^3$

TY 3.4.4 $[\mathrm{Ar}]4s^2$

3.4 Science as a Human Endeavour

Models of the atom: development of atomic theory

1 The results of the experiment would show that the alpha particles would travel straight through the gold and none would be deflected. This is because Thomson proposed that the particles (both positive and negative) were spread throughout the particle, and there would not be a mass heavy enough to deflect the alpha beam.



2 Millikan's oil drop experiment found that all negative charges were multiples of 1.5924×10^{-19} coulombs. Thus, the electron has 1.5924×10^{-19} C of electric charge.

3 The cathode ray tube is a tube with a metal cathode at one end and a metal anode at the other end. It contains a gas at low pressure. When an electric potential is applied, a phosphorescent glow is seen. Thomson used different metals as electrodes and different gases in the tubes. In every test, he observed the same results and from this he deduced that the rays were a stream of particles which came from the negative electrode and were negatively charged. He suggested that these particles were subatomic particles, and were in all matter. He also deduced that atoms must contain positive subatomic particles since atoms have no overall charge.

4 Any two of the following:

Scientist	Law	Theory
Dalton	Gaseous elements are combined in fixed ratios.	Atoms are particles that combine in fixed ratios to make compounds.
Thomson	Very light, negatively charged particles are emitted from the cathode, travel to the anode.	The atom contains separate negative and positive particles throughout.
Rutherford	Alpha particles are deflected by gold foil.	The mass of the atom is in a very dense, small positively charged area called the nucleus.
Bohr	Elements emit characteristic spectra.	Electrons exist in distinct energy levels and energy is released/absorbed when they move between these.
Chadwick	Alpha particles 'push' a neutral particle out of a metalloid; these are able to remove protons from another substance.	There is a heavy neutral particle in the nucleus.

3.4 Key questions

Retrieval

1 *s*—2, *p*—6, *d*—10, *f*—14

2 1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p* 5*s*

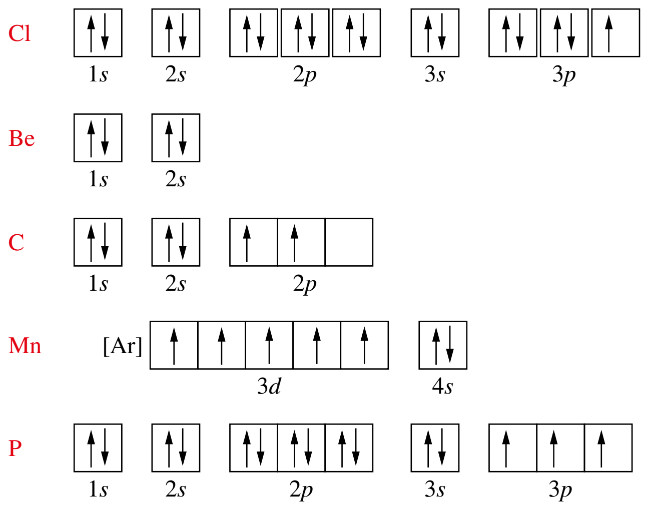
3 In the quantum mechanical model there is no 1*p*, 1*d* or 2*f* subshell.

4

Element (atomic number)	Electron configuration using the shell model	Electron configuration using the subshell model
boron (5)	2,3	$1s^22s^22p^1$
lithium (3)	2,1	$1s^22s^1$
chlorine (17)	2,8,7	$1s^22s^22p^63s^23p^5$
sodium (11)	2,8,1	$1s^22s^22p^63s^1$
neon (10)	2,8	$1s^22s^22p^6$
potassium (19)	2,8,8,1	$1s^22s^22p^63s^23p^64s^1$
scandium (21)	2,8,9,2	$1s^22s^22p^63s^23p^63d^14s^2$
iron (26)	2,8,14,2	$1s^22s^22p^63s^23p^63d^64s^2$
bromine (35)	2,8,18,7	$1s^22s^22p^63s^23p^63d^{10}4s^24p^5$

Comprehension

5



The elements in order of greatest number of unpaired electrons to least number of unpaired electrons is: Mn, P, C, Cl, Be.

6 $1s^22s^22p^63s^23p^63d^64s^2$

7 **a** $[Ar]4s^2$
b $[Ne]3s^1$
c $[Ne]3s^23p^5$

8 The subshell model is a refinement of the shell model. The shell model proposed that all electrons in the one shell were of equal energy. The subshell model describes different electron energy levels (called subshells) within a shell.

Analysis

9 **a** Ga **b** Br **c** Cu **d** Mn
e Sn **f** Sb **g** Xe **h** K

10 $1s^22s^22p^63s^13p^5$

Chapter 3 Review

Retrieval

1 **D** 2 B & C 3 C

4 **a** 5×10^7 nm **b** 1.2×10^7 nm **c** 2×10^{12} nm

5 in the nucleus

6 The protons and neutrons are found tightly packed in the nucleus in the centre of the atom and the electrons are found in the area around the nucleus. Each electron has a precise position and energy level. It has a unique shell, subshell and orbital.

7 **a** atomic number is 24, mass number is 52
b electron number is 24, proton number is 24, neutron number is 28

8 **a** $^{31}_{15}$P **b** $^{80}_{35}$Br **c** $^{9}_{4}$Be **d** $^{70}_{31}$Ga

9 **a** fluorine, F **b** copper, Cu
c phosphorus, P **d** rhodium, Rh

10 **a** $1s^2$
b $1s^2 2s^2 2p^2$
c $1s^2 2s^2 2p^5$
d $1s^2 2s^2 2p^6 3s^2 3p^1$
e $1s^2 2s^2 2p^6 3s^2 3p^6$
f $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$
g $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$

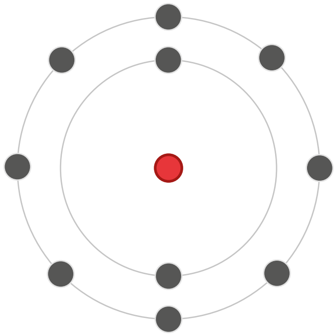
11 **a** copper **b** scandium **c** calcium **d** chlorine

12 **a** sodium **b** aluminium **c** vanadium **d** gallium

Comprehension

13 Zinc oxide nanoparticles are much smaller than zinc oxide powder and sunscreens containing such nanoparticles are clear rather than opaque white. A clear sunscreen is more appealing to users than white sunscreen.

14 The atom has two electrons in the first shell and eight electrons in the second shell. Therefore, the number of valence electrons is eight.



15 The atom is 10 000 (10^4) times larger than the nucleus. Therefore the radius of an atom with a nucleus the size of an orange would be 10 000 × 10 cm. This is 100 000 cm or 1000 m or 1 km.

16 The maximum number of electrons in any shell is given by $2n^2$, where n is the shell number.

17 The atomic number is the number of protons and is the same number for all atoms of each element. The mass number is not a good way to define an element because the mass number is different for isotopes of the same element. The atomic number, however, does not change for alternative isotopes of the same element.

18 **a** tin, 69 neutrons **b** sodium, 12 neutrons
c molybdenum, 54 neutrons **d** aluminium, 14 neutrons

19 In a neutral atom, the number of positive charges equals the number of negative charges. Therefore, the number of protons is the same as the number of electrons. When these two numbers are not equal, the atom will be charged and is called an ion (see Chapter 7).

20 Chemical properties such as bonding and reactivity are related to the number of electrons in the outer shell. Isotopes of an element have a different number of neutrons but the same number of electrons and protons. Therefore, it is expected that chemical properties for isotopes will be the same. However, physical properties that are related to mass are likely to be different because isotopes have different masses.

21 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

Analysis

22 Protons and neutrons have the same mass, whereas electrons are much lighter ($\frac{1}{1800}$ of that of a proton). Protons have a full positive charge, electrons have a negative charge and neutrons have no charge.

23 The two atoms described are not isotopes of the same element because they have a different number of protons. One atom is fluorine (atomic number 9) and one atom is neon (atomic number 10). However, fluorine and neon have the same number of neutrons (10 neutrons).

24 **a** G and H, D and F **b** B, C, D, I
c C **d** A **e** 7

25 The terms 'shell', 'subshell' and 'orbital' describe aspects of the quantum mechanical model of electron arrangement in an atom.

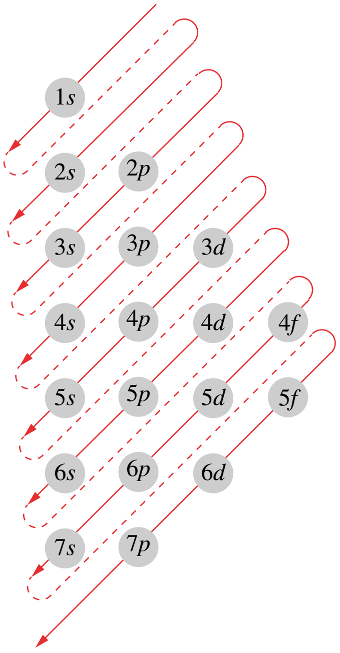
A shell is an energy level given by n, where lower numbers are lower energy levels and are closer to the nucleus. The maximum number of electrons found in any shell is given by $2n^2$. The fluorine atom contains nine electrons, and so will have two electrons in the first shell and seven electrons in the second shell (2,7).

Within a shell there are further divisions in energy level called subshells (s,p,d,f). The s, p, d and f subshells hold a maximum of 2, 6, 10, and 14 electrons respectively. The subshell electron configuration of fluorine is $1s^2 2s^2 2p^5$.

An orbital is a further grouping of electrons—its maximum number of electrons is two. When the orbital holds two electrons they have opposite spin. In the fluorine atom, there are four orbitals containing two electrons and one orbital holding a single electron.

The key difference between the terms 'shell', 'subshell' and 'orbital' is the number of electrons each will hold. Another difference is that shells and subshells describe the position of an electron in terms of energy level, whereas orbitals describe a different characteristic—the spin of the electron.

26 The quantum mechanical model states that lower energy subshells are filled before higher energy subshells (Aufbau principle). The 4s subshell is lower energy than the 3d subshell and therefore will be filled first. This means that the 4s subshell will be filled after the 3p subshell but before the 3d subshell. This order of filling is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ etc.
The following diagram can be used to remember this order.



27 In chromium and copper, the 3d-subshell is being filled, but the 4s-subshell is not full. Their electron configurations are:
Cr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
Cu $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
Most elements have a full 4s-subshell when the 3d orbital is being filled. This is because the 4s-subshell is lower energy than the 3d-subshell. However, a 3d-subshell which has a single electron in each of its five orbitals (according to Hund's rule) is a stable electron arrangement. This means that in Cr, $4s^1 3d^5$ is more stable than $4s^2 3d^4$. The same is true for the Cu atom, which has a stable 3d10 electron arrangement, with all 3d orbitals containing a pair of electrons. This gives Cu a half-filled 4s-subshell and a full 3d-subshell ($4s^1 3d^{10}$).

Knowledge utilisation

28 Carbon nanotubes are a rolled version of graphite that has a diameter of the order of nanometres. They have high tensile strength, high electrical conductivity, high ductility, high surface area and are mostly inert. Some of the projects using these characteristics are shown in the table below.

Project	Nanotube characteristic used	Explanation
bone repair	high tensile strength and low weight	A scaffold of nanotubes is used as a strong basis for new bone to form.
drug delivery	high surface area	Nanotubes 'loaded' with anticancer drug (e.g. doxorubicin) are taken up by cells via cell surface receptors. Infrared radiation subsequently triggers release of the drug.
body and vehicle armour	high tensile strength and low weight	Woven fibres containing centimetre-long pieces of nanotube have exceptional strength and can repel bullets. Flexibility of the fabric is maintained.
solar cells	electrical conductivity	Nanotubes show strong UV–visible absorption and electrical conductivity—transforming solar radiation to electrical current.

29 In 2012, the Large Hadron Collider was used to break apart two high-speed colliding protons. In addition to other subatomic particles previously described as part of the Standard Model, a particle of mass $\frac{125-127\,\text{GeV}}{c^2}$ was described. This particle is often referred to as the Higgs field, and so, as a type of force, it relates to physics concepts. However, its ability to confer mass makes the idea integral to the chemical basis of matter. Gluons, quarks and leptons (see Standard Model) confer mass to protons and neutrons, and the Higgs boson confers mass to the latter two.

Chapter 4 The periodic table and properties of the elements

4.1 The periodic table

4.1 Key questions

Retrieval

1 Periods are the horizontal rows in the periodic table. Periods are numbered 1–7.
Groups are the vertical columns in the periodic table and are labelled 1–18.
The periodic table has four main blocks. The elements in each block have the same type of subshell (*s*, *p*, *d* or *f*) as their highest energy subshell.

2 **a** alkali metals **b** alkaline earth metals
c halogens **d** noble gases

3 **a** 1 **b** 15 – 10 = 5 **c** 17 – 10 = 7 **d** 2

Comprehension

4 Atoms are listed in the periodic table in order of atomic number, which represents the number of protons in the nucleus. The elements are subsequently organised into vertical groups in which all elements have the same number of valence electrons and horizontally into periods based on the number of highest occupied shell. In addition, the elements are classified by blocks in terms of the element's highest energy subshell.

Analysis

5 **a** silicon, Si, 2,8,4 or $1s^22s^22p^63s^23p^2$
b beryllium, Be, 2,2 or $1s^22s^2$
c argon, Ar, 2,8,8 or $1s^22s^22p^63s^23p^6$

6 **a** 4 **b** 2 **c** 1 **d** 1 **e** 7 **f** 3

4.2 Periodic trends

TY 4.2.1 +7

4.2 Key questions

Retrieval

1 The effective nuclear charge of an atom is a measure of the attractive force felt by the valence electrons towards the nucleus. The effective nuclear charge of an atom of carbon is 6 – 2 = +4

2 The first ionisation energy is the energy required to remove one electron from an atom of an element in the gas phase and is represented by the equation $\text{M(g)} + \text{ energy } \rightarrow \text{M}^+\text{(g)} + \text{e}^-$

Comprehension

3 **a** $1s^22s^22p^63s^23p^64s^2$
b $1s^22s^22p^6$
c $1s^22s^22p^6$
d $1s^22s^22p^63s^23p^3$

4 The repulsion between the valence electrons and the inner-shell electrons counteracts the attraction caused by the positive nuclear charge. The inner electrons reduce the effective nuclear charge experienced by the valence electrons from the nucleus. The more electron shells there are, the greater the shielding effect.

5 Electronegativity is the ability of an atom to attract electrons towards itself. Therefore, as the effective nuclear charge increases, the electronegativity increases.

6 **a** **i** F **ii** Fr
b **i** group 17 **ii** group 1
c Electronegativity is the ability of an atom to attract electrons towards itself. As group 18 elements have a full outer shell, they do not have the ability to attract electrons to their outer shell.

7 Across a period, the number of occupied shells in the atoms remains constant but effective nuclear charge increases. The valence electrons become more strongly attracted to the nucleus, so more energy is required to remove an electron from an atom. Therefore, the first ionisation energy increases across a period.

8 When a metal forms a cation, it loses the electrons from the valence shell to attain the electron configuration of the noble gases. Therefore, aluminium loses the three outer valence electrons to attain the electron configuration of neon. Since the electrons come from the outermost shell, its loss will make the Al^{3+} cation much smaller than the Al atom from which it was formed.

Analysis

9

	Trend in atomic radius	Explanation	Example
down a group	increases	Effective nuclear charge stays constant and the number of shells increases as you move down a group. As a result, atomic radius increases.	Li < Na < K; F < Cl < Br, I
left to right across a period	decreases	As you move across a period, the number of occupied shells in the atoms remains constant but the effective nuclear charge increases. The valence electrons become more strongly attracted to the nucleus, so atomic radius decreases across a period.	F < O < N < C < B < Be < Li

	Trend in ionic radius	Explanation	Example
down a group	increases	Effective nuclear charge stays constant and the number of shells increases as you move down a group. As a result, ionic radius increases.	$Li^+ < Na^+ < K^+$ $F^- < Cl^- < Br^-$, I^-
left to right across a period	decreases for species of the same charge	As you move across a period, the number of occupied shells in the ions remains constant for species of the same charge but the effective nuclear charge increases. The valence electrons become more strongly attracted to the nucleus, so ionic radius decreases across a period for the species of the same charge. Anions are larger than cations in the same period.	$Al^{3-} < Mg^{2+} < Na^+ < Cl^- < S^{2-} < P^{3-}$

10 N, B, Cl, Al, Ga

11 K, Na, Al, Ca, S, P, He

12 The two ions are isoelectronic.

	Full electron configuration	Effective nuclear charge
Mg^{2+}	$1s^22s^22p^6$	12 – 10 = 2
F^-	$1s^22s^22p^6$	9 – 10 = –1

Mg has a greater effective nuclear charge; therefore, the attractive force on the valence electrons is much higher for the Mg^{2+} ion and so the valence electrons are pulled closer to the nucleus. This means that the Mg^{2+} ion is actually smaller than the F^-.

13 As you move from left to right across groups 1, 2 and 13–17, the charge on the nucleus increases. Each time the atomic number increases by one, the electrons are attracted to an increasingly more positive nucleus. Within a period, the outer electrons are in the same shell—that is, they have the same number of inner-shell electrons shielding them from the nucleus. Therefore, the additional nuclear charge attracts the electrons more strongly, drawing them closer to the nucleus and so decreasing the size of the atom.

4.3 Impact of the periodic trends on atomic properties

4.3 Key questions

Retrieval

1 The energy released when an atom gains an electron is known as electron affinity as shown in the equation:
$X(g) + e^- \rightarrow X^-(g) + \text{energy}$

2 **a** metal **b** non-metal **c** metalloid
d metal **e** metal **f** non-metal

Comprehension

3 **a** acidic **b** basic **c** basic **d** acidic

4 The reactivity of the alkali metals with water is related to the ease with which the metal ion is ionised. This is based on the first ionisation energy of the elements. Down the group the effective nuclear charge stays constant; however, the number of shells increases. Therefore, the valence electrons are less attracted to the nucleus as they are further from the nucleus. As a result, the energy required to overcome the attraction between the nucleus and the valence electron i s less, and the first ionisation energy decreases down a group. This means that less energy is required to ionise caesium than lithium and so caesium is more reactive than lithium.

5 The electron affinity is the energy released when an electron is transferred to the atom to form an anion. This is related to the ability of an atom to attract electrons towards itself. Generally, as the effective nuclear charge increases, the electron affinity increases.

Analysis

6 Sodium will react more violently with chlorine than iodine to form the sodium halide. This is because the energy required to generate the sodium ion in each case will be the same. However, since the electron affinity of chlorine is greater than the electron affinity of iodine, more energy will be released when the electron is transferred to the chlorine atom.

7 Ar < O < S < Cl

8 Moving across period 2 from lithium to fluorine:
  a the radius of the atoms decreases as the effective nuclear charge increases
  b there is a trend from metals (lithium, beryllium) to non-metals (boron, carbon, nitrogen, oxygen and fluorine)
  c electronegativity increases as the effective nuclear charge increases and size of the atoms decreases

9 **a** fluorine
  b Aluminium and magnesium are metals. Fluorine has the highest effective nuclear charge (it is located furthest right on the periodic table) while also containing the least number of electron shells. This means that fluorine attracts an electron more strongly than other elements and is therefore the most reactive non-metal.

4.4 Introduction to bonding

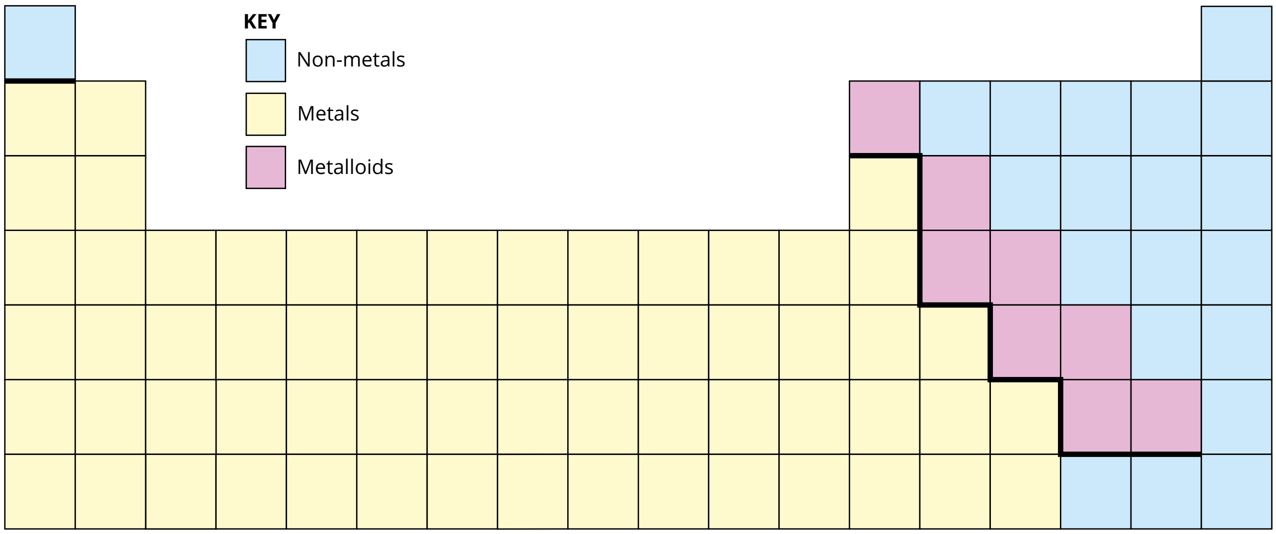
4.4 Key questions

Retrieval

1 Metalloids are elements that exhibit both metallic and non-metallic properties.

Comprehension

2



Analysis

3 **a** ionic **b** metallic **c** covalent **d** ionic
e ionic **f** covalent **g** covalent

4 **a** electronegativity value for carbon = 2.6
electronegativity value for nitrogen = 3.0
electronegativity difference = 3.0 – 2.6 = 0.4 < 0.5 → covalent
b electronegativity value for aluminium = 1.6
electronegativity value for oxygen = 3.4
electronegativity difference = 3.4 – 1.6 = 1.8 > 1.7 → ionic
c electronegativity value for silicon = 1.9
electronegativity value for iodine = 2.7
electronegativity difference = 2.7 – 1.9 = 0.8 → from 0.5 – 1.7 → polar covalent

5 Electronegativity is the ability of an atom to attract electrons towards itself. The greater the electronegativity, the more strongly the valence electrons of an atom are attracted to the nucleus.
If two atoms bonded together have a very large difference in electronegativity the electrons in the bond are attracted to one atom to such a degree that an electron is transferred to the atom to create two oppositely charged ions. This is an ionic bond and is typically formed between a metal and non-metal.
If the difference in electronegativity is less than 1.7, then an electron is not transferred but the electrons in the bond are shared by the two atoms. This is a covalent bond and is formed between two non-metals.
If metal atoms are bonded together, they form a lattice structure and the valence electrons create a mobile sea of electrons that are not localised on the atoms but able to flow across the lattice structure. This is metallic bonding.

Chapter 4 Review

Retrieval

1 B **2** C **3** B **4** C
5 **a** period 1, *s*-block **b** period 2, *p*-block **c** period 3, *p*-block
d period 4, *d*-block **e** period 7, *f*-block
6 Fe **7** potassium **8** 14 – 10 = 4
9 7 – 2 = 5
10 The metallic character decreases from left to right and increases down the group.
11 **a** amphoteric **b** basic c acidic
12 **a** increases **b** decreases **c** increase **d** increases

Comprehension

13 **a** Period is the row in which the element is in (number of the highest occupied electron shell) and the group is the column (related to the number of valence electrons).
b Electron configuration of nitrogen is $1s^22s^22p^3$, the highest occupied shell is 2, and there are five valence electrons—two *s* electrons and three *p* electrons.

14 **a** period 2, group 2 **b** period 3, group 14
c period 4, group 15

15 $Al(g) + energy \rightarrow Al^+(g) + e^-$

16 **a** Chlorine is on the right-hand side of the periodic table and sodium is on the left. Atomic radius decreases across a period because the increasing core charge pulls the outer-shell electrons more tightly to the nucleus, causing the volume of the atom to decrease.
b Fluorine is further to the right on the periodic table than lithium and core charge increases from left to right across the periodic table. As core charge increases, the electrons are held more tightly to the nucleus and more energy is required to remove the first one.
c Ba and Be are in the same group with Be higher than Ba. Going down a group the atom size is increasing, meaning the outer-shell electrons are further from the nucleus. The outer electrons of Be are, therefore, held more tightly and are less readily released.
d The *s*-block elements have an *s*-subshell as their outer occupied electron subshell. The *s*-subshell can take one or two electrons, so the block is only two groups wide.

17 The cations formed when metals in the same period lose electrons are all isoelectronic; however, as you cross the period, the effective nuclear charge of the atom increases. Since the number and shell of the valence electrons is the same, the electrostatic attraction between the nucleus and the valence electrons increases and pulls the electrons closer to the nucleus.
Similarly, when the non-metals gain electrons to form anions they are all isoelectronic. The higher effective nuclear charge across the period attracts the electrons more strongly, thus reducing the ionic size of the anions as you cross the period.

18 **a** Electronegativity decreases down the group. Electronegativity is the ability to attract valence electrons towards the nucleus. The greater the effective nuclear charge, the stronger the attraction; however, the effective nuclear charge is the same for elements in the same group so this does not affect the electronegativity going down the group. Down the group the atoms increase in size. Since the electrostatic attraction between the nucleus and valence electrons is inversely proportional to the distance between them, as the atom gets bigger, the attraction decreases. This reduces the electronegativity of the element. In addition, down the group the number of occupied shells between the valence electrons and the nucleus increases. This increases the shielding effect, further reducing the electrostatic attraction and hence the electronegativity decreases more.

b Atomic radius decreases across a period. Across a period the number of occupied shells between the nucleus and the valence electrons is the same so the shielding effect does not affect the atomic radius. The valence electrons are in the same shell and so about the same distance away; however, as you cross the period the effective nuclear charge increases. This means that the electrostatic attraction on the valence electrons increases and they are pulled closer to the nucleus.

c The ionic radius of an anion increases compared to its parent ion. The effective nuclear charge and shielding effect are the same for each species. However, the number of valence electrons increases. Therefore, since the effective nuclear charge is now being shared by more electrons the electrostatic attraction decreases. This means that the valence electrons are less strongly attracted to the nucleus and so the ionic radius of the anion is larger than that of the parent atom.

d First ionisation energy decreases down a group. Ionisation energy is the energy required to remove the least firmly held electron. The greater the effective nuclear charge, the stronger the attraction; however, the effective nuclear charge is the same for elements in the same group so this does not affect the first ionisation energy going down the group. Down the group the atoms increase in size. Since the electrostatic attraction between the nucleus and valence electrons is inversely proportional to the distance between them, as the atom gets bigger, the attraction decreases. Since the valence electrons are held less strongly, this reduces the first ionisation energy of the element. Therefore, it becomes easier to remove the electron. In addition, down the group the number of occupied shells between the valence electrons and the nucleus increases. This increases the shielding effect, reducing the electrostatic attraction, and hence the first ionisation decreases further.

19 **a** Li **b** Na **c** Ca **d** Al

Analysis

20 **a** Ca **b** Ar **c** C **d** Na or Mg **e** Li **f** N **g** F

21 Ordered elements in groups

Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
Li, K, Cs, Fr	Mg, Ca, Sr, Ba	B, Al, Ga	C, Ge	As, Sb	S, Te	Br, I, Ts	He, Kr, Rn

22 **a** $1s^22s^22p^63s^23p^3$
b $1s^22s^22p^63s^2$
c $1s^22s^22p^63s^23p^63d^{10}4s^24p^2$
d $1s^22s^22p^4$

23 **a** group 13 **b** group 17 **c** group 1
d group 18 **e** group 14 **f** group 14

24 **a** More energy is required to remove the electron from phosphorus.
(Note: Students need to discuss competing effects. Depending on the argument, the conclusion reached by the students could vary from the answer above.)
Phosphorus is larger as the valence electrons are in the third shell and so has increased shielding from the occupied inner shell compared to magnesium in which the valence electrons are in the second shell.
However, magnesium is in group 2 compared to phosphorus, which is in group 15. Therefore, phosphorus has a much larger effective nuclear charge (5 compared to 2). This means that the valence electrons are attracted much more strongly towards the nucleus in phosphorus than they are in magnesium.
The effective nuclear charge is a much greater impact on the amount of energy required to remove the electron and so the ionisation energy of phosphorus is much higher than that of magnesium.

b First ionisation energy decreases down a group; therefore, iodine will require less energy to remove the valence electron. The greater the effective nuclear charge, the stronger the attraction. However, since both elements are in group 17, the effective nuclear charge is the same for each. In fluorine, the valence electron is in the second shell, whereas in iodine it is in the fifth shell, so iodine's valence electrons are further from the nucleus. The electrostatic attraction between the nucleus and valence electrons is inversely proportional to the distance between them, thus larger atoms have a smaller electrostatic attraction. Since the valence electrons are held less strongly, it becomes easier to remove the electron. In addition, down the group the number of occupied shells between the valence electrons and the nucleus increases. This increases the shielding effect, reducing the electrostatic attraction. Hence, the first ionisation decreases further. Thus less energy is required to remove a valence electron from iodine than fluorine.

25 Mg < Na < Ca < K

26 $Ca^{2+} < K^+ < Cl^- < S^{2-}$

27 K < Na < Ca < Mg < F

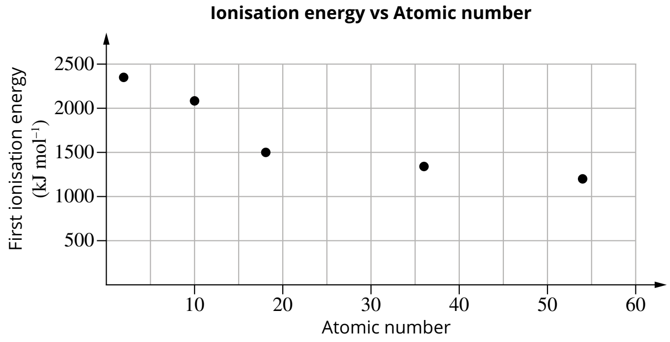
28 Al < K < S < Br

29 **a** nitrogen **b** calcium **c** chlorine

30 The metals and non-metals react to form ionic compounds. The metals lose electrons to form cations and the non-metals gain electrons to form anions. Potassium has a lower ionisation energy than lithium because down the group the effective nuclear charge stays constant and the number of shells increases. Therefore, the valence electrons are less attracted to the nucleus as they are further from the nucleus. As a result, the energy required to overcome the attraction between the nucleus and the valence electron is less, and the first ionisation energy decreases down a group.
Atoms of non-metallic elements undergo chemical reactions to gain electrons and form a stable octet arrangement. Energy is released when an atom gains an electron. The more easily a non-metal can attract or share electrons, the more reactive that non-metal is. Non-metallic elements that have fewer electron shells and higher effective nuclear charges will have a greater attractive force between the nucleus and valence electrons and thus be more reactive. Therefore, chlorine is more reactive than bromine.
For these reasons, K and Cl_2 are the most reactive pair of elements.

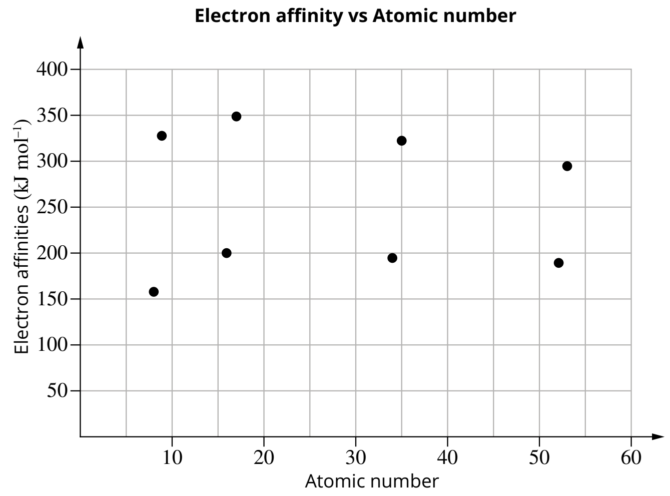
31 **a** lithium < sodium < potassium < rubidium
b The reactivity of metals increases down a group. This is because the number of electron shells increases down a group, so the valence electrons are further from the nucleus and more easily lost. Since metals lose electrons in their reactions, those that lose electrons most easily will be most reactive.

32



Therefore, the best estimate from the data is between about 700 kJ mol^{-1} and 1300 kJ mol^{-1}, based on the two trends, with the value likely to be closer to 1200 kJ mol^{-1}.

33



The electron affinity of fluorine does not follow the trend, due to it small size, and is less than chlorine. If this pattern is repeated with oxygen then its electron affinity would be expected to be approximately 158 kJ mol^{-1}.

34 **a** A and G, D and F
b A and B, D and E, C and G, F and I
c A **d** B **e** F **f** I **g** A, G and I **h** E

Knowledge utilisation

35 **a** As you go across the period, the electron affinity generally increases, from Li (520 kJ mol^{-}1) to B (801 kJ mol^{-1}) to F (1681 kJ mol^{-1}), because of the increased effective nuclear charge.
b N and Ne are exceptions to this trend as the electron affinity is lower than the preceding element. N has a half-filled *p*-subshell ($1s^22s^22p^3$) and Ne has a filled shell ($1s^22s^22p^6$); therefore, the energy that is released when an electron is added to nitrogen and neon is smaller than expected due to the disruption of the additional stability associated with a half-filled and filled subshell.

36 The metallic character is related to the ease with which the element loses electrons. This is related to ionisation energy. Effective nuclear charge increases, and the number of occupied shells remains constant across a period. As a result, the valence electrons become more strongly attracted to the nucleus, and more energy is required to remove an electron. Therefore, first ionisation energy increases across a period. Therefore, the metallic character decreases across a period.
Moving down the group, the effective nuclear charge stays constant and the number of shells increases. Therefore, the valence electrons are less attracted to the nucleus as they are further from the nucleus. As a result, the energy required to overcome the attraction between the nucleus and the valence electron is less, and the first ionisation energy decreases down a group. As a result, the metallic character of the elements increases.

Chapter 5 Analytical techniques

5.1 Mass spectrometry of atoms

TY 5.1.1 A_r(Be) = 10.81

TY 5.1.2 %abundance of ^{206}Pb = 28%; % abundance of ^{207}Pb = 22%; % abundance of ^{208}Pb = 50%

TY 5.1.3 %abundance of 62.95 isotope = 70.5%; %abundance of 64.95 isotope = 29.5%

5.1 Key questions

Retrieval

1 **a** The atomic number is the number of protons of an element.
b The nature of the element is determined by the number of protons in the nucleus; therefore, all atoms of the same element have the same atomic number.

2 Isotopes are atoms of the same element (so have the same number of protons) that have different numbers of neutrons in their nucleus. Therefore, isotopes have the same atomic number but a different mass number.

3

Isotope name	Isotope symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
potassium-37	$^{37}_{19}\text{K}$	19	37	19	18	19
oxygen-17	$^{17}_{8}\text{O}$	8	17	8	9	8
uranium-235	$^{235}_{92}\text{U}$	92	235	92	143	92
uranium-238	$^{238}_{92}\text{U}$	92	238	92	146	92
boron-10	$^{10}_{5}\text{B}$	5	10	5	5	5
boron-11	$^{11}_{5}\text{B}$	5	11	5	6	5

Comprehension

4 The relative atomic mass is related to both the mass of the isotope and the isotopic abundance of each isotope.

5 A mass spectrometer separates the individual isotopes in a sample of an element and determines the mass of each isotope, relative to the carbon-12 isotope, and the relative abundances of the isotopes. It does this by ionising the isotopes and detecting the ions based on their mass.

Analysis

6 Atomic number of chlorine is 17, therefore there are 17 protons and 17 electrons since the mass number is the sum of the number of protons and neutrons. If ^{37}Cl has 17 protons, it must have 20 neutrons.
number of protons = 17; number of neutrons = 20; number of electrons = 17

7 $A_r(M) = 10.81$

8 **a** Cu^{2+}: 29 protons; 34 neutrons; 27 electrons
b P^{3-}: 15 protons; 16 neutrons; 18 electrons

9 Boron-10 will travel further because it is the lighter isotope.

10 **a** $^{39}_{19}\text{K}$ **b** $^{15}_{7}\text{N}$ **c** $^{80}_{35}\text{Br}$ **d** $^{201}_{80}\text{Hg}$ **e** $^{42}_{20}\text{Ca}$
f $^{60}_{27}\text{Co}$ **g** $^{75}_{53}\text{I}$ **h** $^{36}_{16}\text{S}$ **i** $^{238}_{92}\text{U}$ **j** $^{65}_{29}\text{Cu}$

11 A_r(I) = 126.86

12 gallium

13 **a** 85.47 **b** 121.76 **c** 87.62

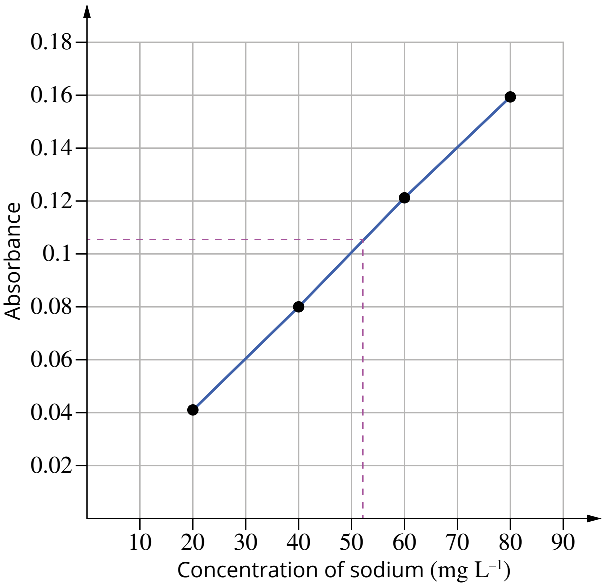
14 The relative atomic mass of argon (39.947) is greater than that of potassium (39.098). Therefore, if the periodic table were organised based on relative atomic masses then potassium would be listed before argon; however, because the grouping of the elements in the periodic table is related to the atomic number (i.e. the number of protons in the nucleus), potassium is placed after argon.

15 ^{69}Ga = 60.13%; ^{71}Ga = 39.87%

16 % abundance of ^{79}Br is 50.70%
% abundance of ^{81}Br is 49.30%

5.2 Atomic absorption spectroscopy

TY 5.2.1



The sample's absorption is 0.104.
The concentration of sodium for the sample can be read off the graph as 52 mg L^{-1}

5.2 Key questions

Retrieval

1 The visible region of light corresponds to light with wavelengths between 400 and 700 nm on the electromagnetic spectrum.

2 When a sample containing copper is heated in the flame of a Bunsen burner, the flame turns a green colour. This is because the *electrons* in the copper atoms absorb energy and move to *higher* energy levels and then *emit* light that corresponds to a green colour as they return to *lower* energy levels.

3 Lyman (ultraviolet), Balmer (visible), Paschen (infrared)

Comprehension

4 The electromagnetic spectrum can be described as a stream of photons, each travelling in a wave-like pattern, carrying energy and moving at the speed of light. It is classified by the range of frequencies or wavelengths of electromagnetic radiation. This includes visible light, UV and IR radiation.

Analysis

5 Strontium should have a scarlet flame instead of the yellow colour observed in the experiment. This could be because there was an impurity on the nichrome wire from a previous experiment. Since sodium ions often dominate the emission from other ions, it is likely that this was the solution used before the strontium flame test. Therefore, the student should wash the wire in distilled water, thoroughly clean the wire by dipping it in concentrated hydrochloric acid and then hold the wire in the tip of the blue part of the flame until no colour is seen before repeating the experiment.

6 **a** green

b It is necessary to heat the wire strongly before performing the flame test to vaporise traces of other substances that could lead to a false result.

c Copper wire would give the flame a green colour, and so interfere with the test. The wire must be made of an element that does not change the colour of the Bunsen burner flame.

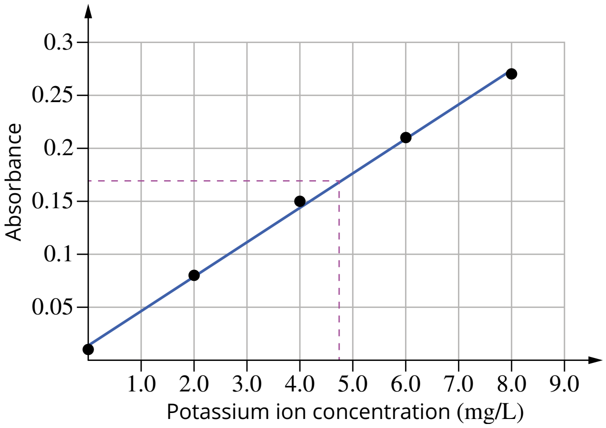
d Flame tests are rarely used for analysis by today's chemists because relatively few elements may be analysed in this way; the presence of one element (sodium, for example) may mask the presence of another, and colours emitted by some elements are similar and difficult to distinguish; and a relatively large amount of sample is needed to give a clear colour.

7 The Lyman and Balmer series are both spectral lines that are formed when an excited electron falls back to a lower energy shell and in the process a photon of light is emitted. In the Lyman series the ground state to which the electron returns is the $n = 1$ shell; whereas the ground state for the Balmer series is the $n = 2$ shell. The spectral lines become progressively closer together as the number of the excited shell increases until a continuum is formed. The wavelength of the light emitted is in the ultraviolet region for the Lyman series but in the visible region for the Balmer series.

8 The higher the shell number (n), the higher the energy of the electron in the shell. Therefore, the electron in the $n = 5$ shell has more energy than when it is in the $n = 2$ shell. This means the electron loses energy when it falls from $n = 5$ to $n = 2$. The energy is released as light so as the electron falls from the $n = 5$ to $n = 2$ a photon of light is emitted.

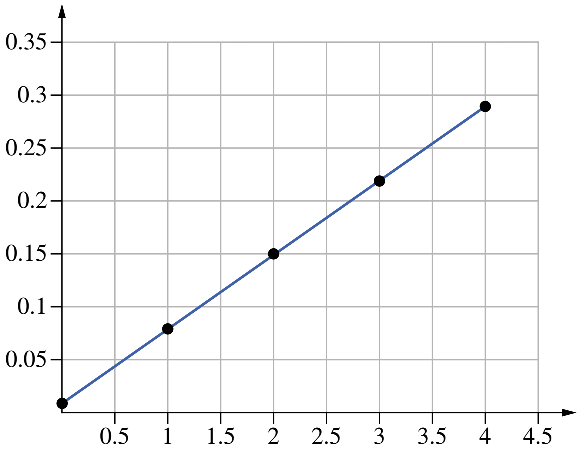
9 Atomic emission spectroscopy (AES) is regarded as superior to flame tests because, in AES, more elements produce emission spectra in the hotter flame, and a much smaller quantity of an element needs to be present in order for it to be detected. Using AES, the amount of the element present in a sample can be accurately determined, even in the presence of larger amounts of other elements that would mask the flame colour of the element to the naked eye.

10 **a, b**

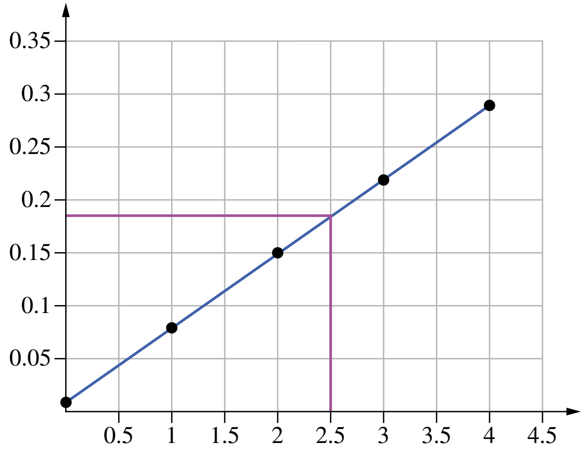


From the calibration curve, the potassium ion concentration is about 4.8 mg/L.

11 **a**



b 2.57 ppm



c The 5 mL sample was diluted to 50 mL. This is a factor of 10.
[Fe] in the undiluted Milo = 25.7 ppm

d Step 1: Rewrite the concentration in ppm.
25.7 ppm means 25.7 g Fe per 10^6 g Milo
Step 2: Assume the density of Milo is 1 g mL^{-1} and calculate the mass of Fe in 250 g Milo.
25.7 g Fe per 10^6 mL Milo = ? g Fe per 250 mL Milo.

$$m(\text{Fe}) = 25.7 \times \frac{250}{10^6}\ \text{g}$$
$$= 6.4 \times 10^{-3}\ \text{g (2 significant figures)}$$

e $\%\text{Fe} = \frac{6.4}{18} \times 100 = 36\%$

f The water used to prepare the standards may have contained traces of iron.

g If it had not been diluted, the absorbance reading for the Milo would have been outside the range of the calibration curve. The liquid might also have been too thick to spray into the flame.

5.3 How analytical chemistry has expanded our understanding of the universe

5.3 Science as a Human Endeavour

Distribution of elements in the universe

1 Student answers will vary.
2 Student answers will vary.

5.3 Key questions

Retrieval

1 Spectroscopy is an analytical technique that provides information based on the interaction between matter and electromagnetic radiation.

2 Once a tree dies, the ratio of carbon-14 to carbon-12 isotopes begins to decrease as carbon-14 decays to carbon-12. Conventional carbon dating counts the radioemissions from the sample to estimate the length of time since the tree died. An accelerator mass spectrometer measures the relative isotopic abundances of the carbon-12 and carbon-14 isotopes.

Comprehension

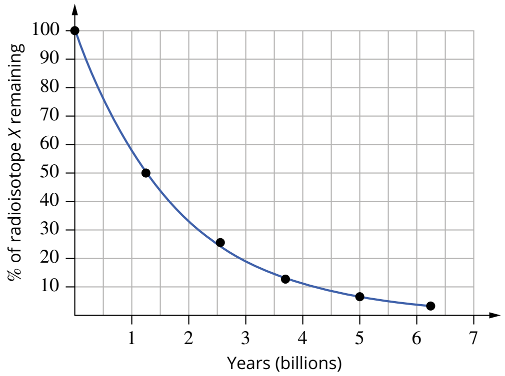
3 Atoms have particular electron shells, which have particular energy levels. This means that the atoms only absorb light of certain frequencies associated with the energy gaps between these shells. By measuring the light emitted by the stars, astronomers can identify the absorption lines in the spectrum and match the position of lines observed to colours emitted by different elements. Comparison of the spectrum obtained for a star with the spectra of known elements allows astronomers to identify the elements in the star.

4 The Doppler effect can make the radiation we receive from objects display a different wavelength from what we expect. By measuring the amount of this ‘shift’ in the wavelength, astronomers can determine whether an object is moving towards or away from us. The greater the shift appears, the greater the relative speed. Therefore, by measuring the Doppler shift, the relative velocity can be determined.

Analysis

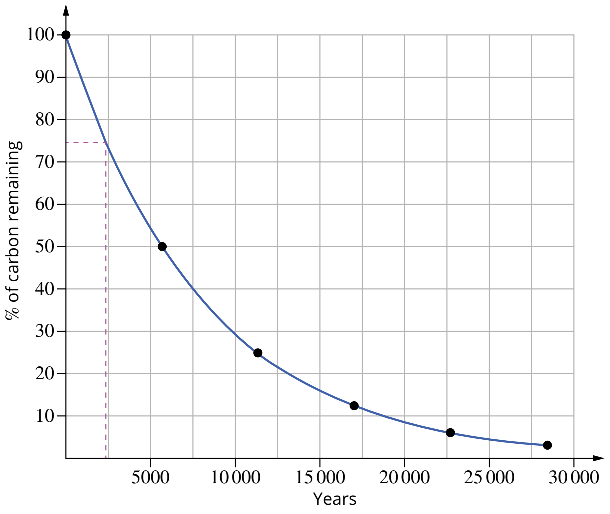
5 The spectral lines for star A are shifted towards the red end of the spectrum. This redshift indicates that the star is moving away from Earth.
The spectral lines for star B are shifted towards the blue end of the spectrum. This blueshift indicates that the star is moving towards Earth.

6 **a** The half-life is approximately 1.25 billion years.
b The time taken for the radioisotope to decay from 100% to 50% is the half-life of the isotope. As the time taken for the remaining 50% to decay to 25% also represents the time for half of the isotope to decay (equivalent to the half-life), the time is the same.



c Based on the graph, the rock is about 1.7 billion years old.
d From the half-lives shown in the table, the radioisotope that is closest in half-life to the decay shown in the graph is potassium-40.

7 **a** A skull with 50% of its original carbon-14 remaining is equivalent to one half-life; therefore, the skull is about 5700 years old.
b After one half-life there is 50% of the original carbon-14. Then after an additional half-life, there is 25% of the original carbon-14 left. After an additional half-life, half of the 25% of the carbon-14 has decayed, meaning that 12.5% of the original carbon-14 remains. This means that the shark’s tooth is about three half-lives old. This is 17 100 years.
c Based on the decay rate shown in the graph, the Dead Sea Scrolls are about 2300 years old.



Chapter 5 Review

Retrieval

1 **a** carbon

b the mass number of the isotope, which is the sum of the protons and neutrons in the nucleus of the isotope

c the atomic number, which is the number of protons in the nucleus. The identity of the element is determined by the number of protons in the atom. Since all the isotopes have the same atomic number, they all are atoms of carbon.

d

	Number of protons	Number of neutrons	Number of electrons
$^{12}_{6}C$	6	6	6
$^{13}_{6}C$	6	7	6
$^{14}_{6}C$	6	8	6

2 The relative isotopic abundance is the percentage abundance of an isotope in the natural environment.

3

Component	Description
flame	where the sample is sprayed and light is absorbed
hollow source cathode lamp	produces light with wavelengths that are absorbed by the metal being analysed
computer	converts the amount of light detected into the amount of light absorbed by the sample
monochromator	selects a specific wavelength of light
detector	measures the amount of light

4 D

5 C

Comprehension

6 atomic number 45; mass number 103; symbol $^{103}_{45}Rh$

7 The number of peaks in the mass spectrometer indicates the number of isotopes. There are five peaks; therefore, zirconium has five isotopes.

8 **a** lilac

b The flame would be yellow. Since the emission from sodium ions is very strong, it would overwhelm the lilac colour from the potassium and so the lilac colour may not be visible.

c The energy required to promote an electron in a chlorine atom is much higher than for the potassium atom. As a consequence, the flame of a Bunsen burner excites relatively few chlorine atoms. Furthermore, energies emitted as electrons in the excited atoms return to their lowest energy levels are mainly outside the energies of visible light.

9 When atoms absorb energy, electrons are excited to different higher energy levels. When the electrons return to the ground state from these excited states, they release the energy as light by undergoing a number of transitions of different energy, as shown in Figure 5.2.9 on page 119. Each transition results in a line of specific energy in the emission spectrum.

10 Fingerprints are unique and can be used to determine the identity of a person. Similarly, the absorption spectrum of an atom is specific to the composition and temperature of the element. Therefore, since stars all contain a unique composition of elements the absorption spectrum of each star will be unique.

Analysis

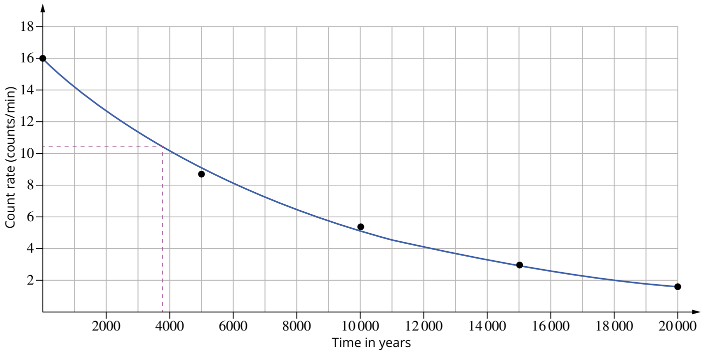
11 Atoms of the same element all have the same number of protons and electrons; however, the number of neutrons can vary. This is the reason for isotopes. The mass and isotopic abundance of atoms of the same element can be identified by a mass spectrometer.

12 A_r(Cu) = 63.60

13 As a consequence of the presence of different numbers of protons in the nuclei of barium and calcium atoms, the energies of electrons in the shells of the two atoms are not the same. Different amounts of energy are released when an electron moves from a higher energy level to a lower one in a barium atom compared with when the same process occurs in a calcium atom. These different energies are seen as light of different colours.

14 The unknown mixture contains gas A and gas D.

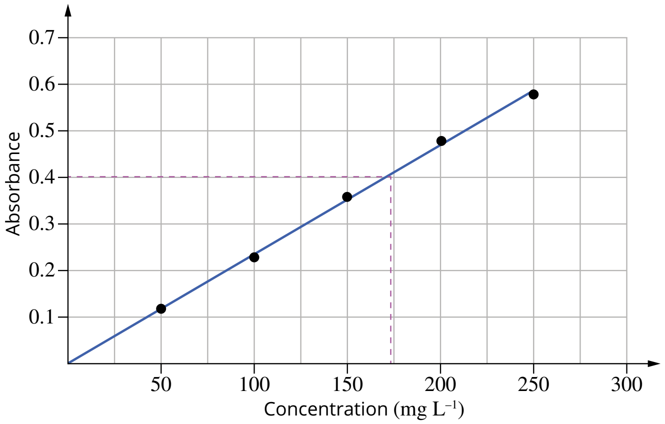
15



From the graph, the count rate corresponds to 3700 years. This means that the tree died approximately 3700 years ago, dating the eruption to around 1700 BCE.

16 The percentage of carbon-14 atoms still present is 92%. If 16 counts per minute is 100% then 92% would be 14.7 counts per minute. From the graph, 14.7 counts per minute equates to about 740 years. Therefore, the shroud was made around 1250 CE. This is very approximate due to the accuracy of the graph and the ability to mark the respective values.

17



The concentration of the copper ions is about 180 mg L^{-1}.

18 From the graph, the concentration of lead in the solution prepared from the oyster sample is 9.0 ppm, which is equivalent to 9.0 μg mL^{-1}. So, in the 10 mL solution there is $9.0 \times 10^1 = 90$ μg of lead. Since 90 μg lead was present in the 1.50 g portion of oyster:

concentration in the oyster = $\frac{90\ \mu g}{1.50\ g} = 60\ \mu g\ g^{-1}$, or 60 ppm

Since the recommended limit of lead in oysters for human consumption is 2.5 ppm, it would be wise not to eat the oysters that were the subject of this test.

19 A_r(A) = 18.67

20 **a** A_r(O) =16.00 **b** A_r(Ag) = 107.86 **c** A_r(H) = 1.01

21 **a** 2.45 ng mL^{-1}

b Step 1

Calculate the mass of lead in the total volume of milk, using the answer from part **a**. This is the same mass of Pb as is present in 2.5 g of milk powder.

2.45 ng per 1 mL milk = 2.45×50 ng per 50 mL milk
= 122.5 ng per 50 mL milk

Step 2

Convert into concentration of lead in ng g^{-1}.

$$[\text{Pb}]=\frac{122.5}{2.5}$$
$$=49 \text{ ng g}^{-1}$$

c Convert concentration to ppm.

$$49 \text{ ng per 1 g of milk powder} = 49\times10^{-9} \text{ g per 1 g of milk powder}$$
$$= 49\times10^{-3} \text{ g per } 10^{6} \text{ g of milk powder}$$
$$= 0.049 \text{ ppm}$$

d
$$\%\text{Pb}=\frac{\text{mass of Pb}}{\text{mass of milk powder}}\times100$$
$$=\frac{0.049}{10^{6}}\times100\%$$
$$=4.9\times10^{-6}\%$$

e The distilled water still contains minute traces of contaminants from the air, glassware or any other material the water has come into contact with. This arises because of the extreme sensitivity of the technique.

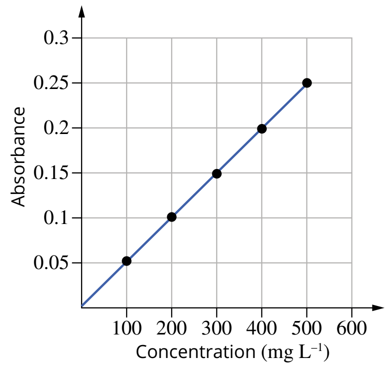
22 8%

23 A mass spectrum shows peaks of *m/z* ratio for each isotope present. Since the charge of the isotope is generally +1, the ratio allows scientists to determine the mass of each isotope. The height of each peak enables the determination of the relative isotopic abundance. The identity of an element can be identified by calculating its relative atomic mass from the relative abundance and mass of each isotope.

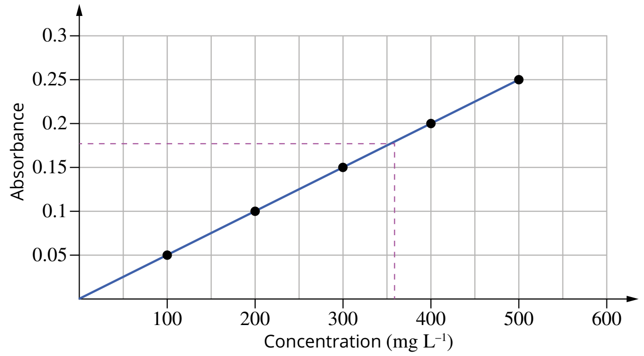
24 C

Knowledge utilisation

25 a



b



Reading from the graph, the concentration of NaCl is about 360 mg L^{-1}.

c The sample was diluted 40 times so the undiluted sauce is 40 times more concentrated.

$$= 360\times40$$
$$=14\,400 \text{ mg}$$
$$=14.4 \text{ g L}^{-1}$$

d All the Na^+ in the sauce is present as NaCl or NaCl was the only source of Na^+ ions.

e The concentration of sodium ions (g per L) in the sauce needs to be converted to mass of NaCl per litre.

$$[\text{Na}^+]=\frac{14.4}{23}$$
$$=0.626 \text{ moles}$$

Therefore, number of moles of NaCl = 0.626

$$\text{mass of NaCl per litre}=0.626\times(23+35.5)$$
$$=36.6 \text{ g}$$
$$\text{amount of sauce}=\frac{\text{recommended intake}}{\text{concentration of Na in 1 L of sauce}}$$
$$=\frac{2.5}{36.6}$$
$$=68.3 \text{ mL}$$

f The light source used in the AAS emits a wavelength that is specific to sodium ions (Na^+) . So the selected wavelength will only be absorbed by the sodium ions (Na^+).

Chapter 6 Metals

6.1 Properties of metals

6.1 Key questions

Retrieval

1 Malleable means the material can be shaped.

2 Transition metals are harder, more dense and have higher melting points. Some transition metals have strong magnetic properties. Transition metal compounds tend to be brightly coloured.

3 Mercury is the only metal that is a liquid at room temperature.

Comprehension

4 Na is in group 1 of the periodic table. It has only one valence electron with a relatively low ionisation energy and is highly reactive. Iron is a transition metal in group 8 of the periodic table. It has delocalised electrons in its valence shell and is positioned towards the right-hand side of the metals.

5 Ductility means the material is able to be drawn into a wire, while malleable means the material can be shaped by beating or rolling.

6 The smaller the first ionisation energy of an element, the greater the metallic character of that element.

Analysis

7 Physical and chemical properties of metals that are important in bridge building would include lack of chemical reaction with water and oxygen, being malleable and ductile when building the bridge, and a high tensile strength to support the traffic on the bridge.

8 a sodium

b silver

c Potassium and sodium are in group 1 as they are alkali metals, while gold and silver are in group 11 as they are transition metals. Alkali metals are very reactive (and as such are kept under oil to stop reaction with moisture in the air) and so are not found as pure metals naturally on Earth. They can only donate one electron and they are soft and shiny. Silver and gold are transition metals. Both of these are very unreactive and can be found as pure metals naturally.

9 a Copper is malleable and ductile, therefore able to form wires, and has high electrical conductivity.

b Metals would have to have low reactivity with air and water and be solid at room temperatures. Cost of the material would also be a factor.

10 Alkali metals are too soft to be used as structural metals, and they also react vigorously with the air and water to produce hydrogen gas.

11 Titanium is a very strong metal, with a high melting point, therefore it is able to withstand the extreme heat of reentry vehicles from space. It is a light, low density metal, therefore the payload of aircraft and space vehicles can be maximised. The low chemical reactivity of the metal means that it is able to withstand environment changes in humidity or oxygen levels in atmosphere, without rusting or decomposing. It also easily forms alloys, which can improve the strength of the material further.

6.2 Metallic bonding

TY 6.2.1 Mg has an atomic number of 12. This means a neutral atom of aluminium has 12 electrons. The electron configuration is $1s^2 2s^2 2p^6 3s^2$.

Mg has two electrons in its outer shell (the $3s^2$ electrons). Mg atoms will tend to lose these two valence electrons to form a cation with a charge of +2. The outer-shell electrons become delocalised and form the sea of delocalised electrons within the metal lattice.

If the Mg is part of an electric circuit, the delocalised electrons are able to move through the lattice towards a positively charged electrode.

6.2 Key questions

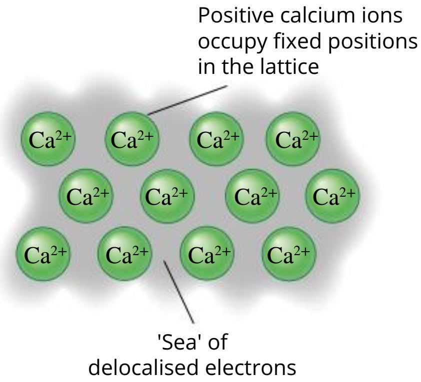
Retrieval

1 Any three of the following: dense, malleable/ductile, good conductors of heat and electricity, lustrous

2 The metallic bonding model states that a solid sample of metal will contain positive cations arranged in a closely packed structure with delocalised electrons moving freely throughout the lattice. The metal appears lustrous as the delocalised electrons reflect light, causing the metal to appear shiny.

Comprehension

3 **a**



b strong electrostatic forces of attraction between Ca^{2+} ions and the delocalised valence electrons

4 Barium has a high melting temperature because there are strong attractive forces between the positive ions and the delocalised electrons. Barium conducts electricity because the delocalised electrons from the outer shell are free to move.

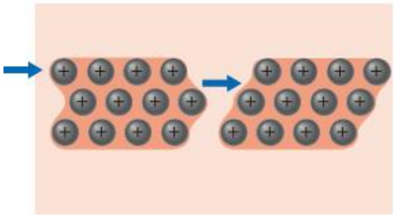
5 **a** lustrous and conduct heat and electricity

b Graphite has two-dimensional lattices separated by delocalised electrons. These delocalised electrons are free to move allowing graphite to conduct electricity. The electrons in graphite collide with each other while moving, allowing the transfer of heat energy from one place to another within the graphite.

Analysis

6 **a** Metals are malleable as the sea of delocalised electrons enables the layers or cations to slide past each other without being repelled.

b When a force is applied, the layers of cations are forced past each other. The layers of ions are still held together as the sea of delocalised electrons stop the repulsion of one cation to another.



7 Copper has an atomic number of 29 with an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. As it only has one electron in its valence shell, it easily loses this electron to the sea of delocalised electrons surrounding the copper cations in the lattice. If the copper is part of an electric circuit, the delocalised electrons are able to move through the lattice towards a positively charged electrode.

8 **a** Aluminium has an atomic number of 13, which means a neutral atom of aluminium has 13 electrons.
The electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^1$. As a result, there will be three electrons in its outer shell that will be donated to the sea of delocalised electrons.

b Aluminium can donate three electrons to the sea of delocalised electrons. This provides many electrons which surround the positive Al^{3+} cations in the lattice. Thus when a force is applied, the layers of the lattice should slide over each other easily as there are three donated electrons per cation to stop repulsion. As sodium is a group 1 metal, it can only donate one valence electron to its sea of delocalised electrons per cation. This will mean there will be fewer electrons to reduce the repulsion effect of the sodium cations if a force is applied. Therefore, you would expect that aluminium can withstand the effect of a force better than sodium.

6.3 Modifying metals

6.3 Key questions

Retrieval

1 In general, the metallic elements added to make substitutional alloys have fairly similar chemical properties and form cations of a similar size to the main metal, while in interstitial alloys, a small proportion of an element with significantly smaller atoms is added to a metal. The added atoms sit in interstices (very small spaces) between metal cations in the metallic lattice.

2 Metals could be given a coating in order to make them more decorative (given a colour coating or image on the metal), better suited to a purpose (e.g. Teflon coating for frying pans makes them non-stick), or resistant to rust or other corrosion ('tin' cans are steel cans with a tin coating to make the can resistant to rusting).

3 **a** A 20-cent coin contains copper and nickel. High-carbon steel contains iron and carbon.

b The 20-cent coin is a substitutional alloy similar to Figure 6.3.3. High-carbon steel is an interstitial alloy, similar to Figure 6.3.2.

4 Powder coating is a method of applying a decorative and protective finish. The powder is a mixture of finely ground particles of pigment and resin, which is sprayed onto a surface electrostatically (Figure 6.3.19). The charged powder particles adhere to the surface. When heated, the particles fuse to form a smooth surface that is resistant to scratching and marking.

5 **a** copper and nickel; harder, more corrosion resistant and a silver colour

b tin and lead; lower melting temperature

c gold, silver and copper; harder

d iron, nickel and chromium; resists rusting, stronger

e mercury and zinc (sometimes a little silver is added); harder, non-toxic

Comprehension

6 Steel is used instead of iron as this alloy is stronger, more flexible and resistant to corrosion. Heating the horseshoe during the process allows the worker to change the crystal structure through heat treatment. The final shaping and hammering is an example of work hardening, which aligns the crystals and increases strength.

Analysis

7 The metal in the hooks becomes work hardened and brittle.

8 **a** Aluminium contains small areas of regular metallic lattices called crystals. When the aluminium is annealed, the crystal structure is changed to contain more large crystals. Larger crystals are more flexible and easier to shape than smaller crystals and the metal is less likely to break along crystal boundaries during shaping.

b Quenching the aluminium by heating to the critical temperature and then rapidly cooling causes the growth of small crystals. These crystals make the metal stiffer and do not allow the metal to deform as easily.

9 Coating of metals can usually be done more quickly, cheaply and uses a lot less energy than heat-treating metals.

10 **a** Chromium would increase hardness and tensile strength and will resist corrosion. These properties could be useful in a bike frame and now many bike frames are made of a mixture of chromium and molybdenum in with the iron.

b Nickel would increase toughness, tensile strength and will resist corrosion, but while it will be light, it might not deal with the stresses a bike frame might undergo.

c Cobalt would improve the magnetic properties of the iron for the frame, but that is where its benefit would end.

d Lithium would be light, but it is highly reactive (as a group 1 metal), so would be of no use outside.

e Aluminium would make the bike lighter, but it does not provide much strength. Bike frames in the past were made of aluminium but this was found to limit the life of the bike if it was ridden a lot or used as a mountain bike.

11 **a** interstitial

b substitutional

c substitutional

12 **a** copper/beryllium: high strength, excellent electrical conductivity

b iron/chromium: increased strength, increased resistance to corrosion

c aluminium/magnesium: high strength, light weight

6.4 Extraction of iron from its ore

6.4 Key questions

Retrieval

1 $Fe_3O_4(l) + 4CO(g) \rightarrow 3Fe(l) + 4CO_2(g)$; magnetite + carbon monoxide → iron + carbon dioxide

2 Iron ore is the source of iron. Limestone is the source of calcium oxide, which reacts with unwanted contaminants to form molten slag. Coke is used to produce heat and carbon monoxide, which reacts with the iron oxides to form iron metal. Air is pumped in as a source of oxygen for the combustion of the coke.

Comprehension

3 Carbon dioxide is the main environmental pollutant from the production of iron in a blast furnace. The carbon dioxide is formed by the reaction of carbon from the coke with oxygen at high temperatures. $C(s) + O_2(g) \rightarrow CO_2(g)$

4 Iron reacts with water and oxygen to produce iron oxides found in ores.

Analysis

5 **a** Haematite (Fe_2O_3) is dark and reflective when polished; magnetite (Fe_3O_4) is a black mineral and magnetic; and pyrite (FeS_2) is a shiny yellow colour, often called ‘fool’s gold’.

b Magnetite ore’s magnetic properties are helpful during the identification and extraction process.

6 The production of iron requires large amounts of coke, which is usually produced on site from coal. Proximity to a coal mine is required. The production of iron may generate local environmental pollution. Iron furnaces should be located away from residential populations, agriculture and conservation areas. The production of iron requires many workers. A site must have suitable accommodation and support for workers. The iron produced must be able to be processed and shipped from the production site. Access is required to transport networks such as ports or railways.

7 **a** False. Slag is less dense and floats on top of the molten iron.

b False. Iron ore contains impurities such as silica, alumina and manganese oxides.

c true

d true

e False. The reaction between carbon in the coke and oxygen to produce carbon dioxide releases a considerable amount of heat energy to maintain the high temperatures required in the furnace.

8 Fuel gas is a by-product from the production of coke and it is used to provide heat to the blast furnace.
Tar, benzene, toluene and ammonium sulfate are by-products of the production of coke, and these can be sold off as raw materials for other chemical industries.
Heat energy from the reaction of coke with oxygen maintains the high temperatures required in the blast furnace.
The carbon monoxide produced when coke reacts with oxygen is then used in a series of reactions to extract iron from iron oxide.
Calcium oxide from the heating of limestone in the furnace reacts with the unwanted materials from the iron ore to become slag, which in molten form covers the newly produced molten iron, saving the iron from reacting with the air and re-forming iron oxide.

6.5 Metallic nanomaterials

6.5 Science as a Human Endeavour

Nanomaterials

1 A substance that kills off the cancer cells, called a tumour necrosis factor (TNF), is hidden inside a gold nanorod. This stops the body’s immune system from identifying the TNF. The TNF accumulates in cancer cells, destroying the tumour, but does not seem to affect the healthy cells.

2 The iron nanoparticles have a high surface area and high reactivity to the CCl_4. They react to produce smaller and nontoxic particles that can be removed from the area or broken down further, as they are no longer harmful.

6.5 Key questions

Retrieval

1 iron + oxygen → iron oxide

Comprehension

2 nanowire

3 8.34×10^{-7} m is 834 nm. This particle would be too big to be classified as a nanoparticle.

Analysis

4 A is a nanoparticle. B is a nanowire, as the length is more than five times the diameter. C is a nanorod, as the length is between three and five times the diameter. D is not a nanoparticle as all of its dimensions are greater than 100 nm.

5 Gold nanoparticles are not large enough to have a ‘sea’ of delocalised electrons or a consistent surface to reflect light.

6 The width of a nanowire is in the range of standard nanomaterials. This changes the properties of the metal atoms as they are not exposed to delocalised electrons in the same way as in bulk metals. The long length of a nanowire is not enough to give it electrons that behave in the same way as electrons in bulk metals.

Chapter 6 Review

Retrieval

1 Mg, Ca, Sr

2 **a** silver

b It is too expensive and tarnishes readily.

c aluminium, copper (combined with stainless steel)

3 **a** rocks and minerals from which iron can be economically extracted
b A mineral is a naturally occurring inorganic substance that is solid and can be represented by a chemical formula, e.g. quartz.
c A mineral containing iron. Haematite is composed of iron oxides with the formula Fe_2O_3 and is the main source of iron ore in Australia.
d a compound containing the elements iron and oxygen, e.g. Fe_2O_3, Fe_3O_4 and FeO

4 **a** Na: period 3, group 1
K: period 4, group 1
Ca: period 4, group 2
b Na: $1s^22s^22p^63s^1$
K: $1s^22s^22p^63s^23p^64s^1$
Ca: $1s^22s^22p^63s^23p^64s^2$
c **i** The atoms of Na are smaller than those of K, so the delocalised valence electrons of Na are closer to the positive nuclear charge than those of K. The electrostatic forces of attraction between delocalised electrons and cations are stronger in Na, so Na requires more energy to overcome the metallic bonding to boil the metal.
ii Valence electrons are in the fourth shell in the atoms of both Ca and K. However, there are twice as many valence electrons in the atoms of Ca. Also, the charge on a calcium cation is +2 as opposed to +1 on the potassium cation. The electrostatic forces of attraction between delocalised electrons and cations are stronger in Ca and so it requires more energy to overcome the metallic bonding to boil the metal.

Comprehension

5 **a** When a current is applied to the copper wire, the free-moving delocalised electrons move from one end to the other and so the copper wire conducts electricity.
b The delocalised electrons in the metal spoon obtain energy from the boiling mixture and move more quickly. These electrons move freely throughout the spoon, colliding with other electrons and metal ions, transferring energy so that the spoon becomes warmer and, eventually, too hot to hold.
c A lot of energy is required to overcome the strong forces of attraction between the iron ions and the delocalised electrons in the metal lattice, so that the iron changes from a solid to a liquid.
d Because of the strong forces of attraction between them, the lead ions and the delocalised electrons form a closely packed three-dimensional structure. Also, the lead atom itself has a higher mass-to-volume ratio than the sulfur atom. This means that the density—the mass per volume—is high.
e As the copper is drawn out, the copper ions are forced apart and the delocalised electrons rearrange themselves around these ions and re-establish strong forces of attraction.

6 **a** **i** valence electrons not restricted to a region between two atoms
ii a regular three-dimensional arrangement of a very large number of positive ions or cations
iii the electrostatic attraction between a lattice of cations and delocalised electrons
b valence electrons

7 A metal wire contains an extended lattice of metal cations surrounded by a sea of delocalised electrons. The electrons are charged and free to move and so can conduct electricity.

8 The metal cations are in a regular three-dimensional arrangement and are surrounded by a mobile sea of delocalised electrons.

9 Iron reacts with water and oxygen to produce iron oxides found in ores.

10 calcium oxide + silica → calcium silicate

11 **a** aluminium Al, copper Cu, gold Au, iron Fe, silver Ag
b copper: period 4, group 11, first transition series, *d*-block
gold: period 6, group 11, third transition series, *d*-block
iron: period 4, group 8, first transition series, *d*-block
silver: period 5, group 11, second transition series, *d*-block
c gold and silver
d copper, gold, iron and silver
e gold

Analysis

12

Atom	Number of valence electrons	Number of electrons in its cation	Number of electrons donated to the sea of delocalised electrons
magnesium	2	10	2
aluminium	3	10	3
sodium	1	10	1

b aluminium > magnesium > sodium. They are ranked like this because of the number of electrons donated to the sea of delocalised electrons. The more electrons in the delocalised sea, the higher the bond strength and so the more energy required to break apart the solid, resulting in a higher melting point.

13 Iron ore is the source of iron. Limestone is the source of calcium oxide. Coke is used to produce heat and carbon monoxide, which reacts with the iron oxides to form iron metal. Air is pumped in as a source of oxygen for the combustion of the coke.

14 needle 2 < needle 3 < needle 1

15 **a** **i** iron (steel) or aluminium
ii Iron and steel are strong. Aluminium has a low density (light) and can be easily coloured.
iii Iron rusts easily. Aluminium is soft and lacks strength.
b **i** copper
ii It is a good conductor of electricity and is ductile.
c **i** gold, silver and platinum
ii They are non-reactive, malleable, ductile and lustrous.

16 **a** Al atom: $1s^22s^22p^63s^23p^1$; stable cation: $1s^22s^22p^6$
b To make a substitutional alloy, the metal that is alloyed with aluminium should have a similar size. Magnesium and sodium would both have similar sizes; however, sodium would be too reactive to be effective in an alloy.

17 Both gold nanoparticles and bulk gold have the same number of protons (79) and the same number of valence electrons will be released (one electron) to become delocalised. However, the valence electrons that are delocalised in gold nanoparticles are too few in number to behave as a ‘sea of delocalised electrons’. This results in the electrical and thermal conductivity and metallic lustre seen in bulk gold but not seen in gold nanoparticles. Gold takes on a ruby colour when the gold particles are reduced to nano-size and they are more sensitive to heat.

18 **a** low density
b high electrical conductivity
c high tensile strength

19 **a** B, F
b coke
c a mixture
d limestone

20 **a** copper and nickel; harder, more corrosion resistant and a silver colour
b tin and lead; lower melting temperature
c gold, silver and copper; harder
d iron, nickel and chromium; resists rusting, stronger
e mercury and zinc (sometimes a little silver is added); harder, non-toxic

21 **a** The positive ions are arranged in a regular, three-dimensional lattice.
b Stress corrosion cracking can occur between the crystal grains.

Knowledge utilisation

22 Appropriate alloys may include iron, manganese, chromium, carbon and nickel.

23 Aluminium is extracted from its ore by electrolysis. There was no source of electricity available for this process until 1886.

24 Galvanising is when the metal is coated with zinc. This is referred to as 'sacrificial coating' because the zinc is used to corrode away, leaving the steel fine underneath. It requires galvanising when this occurs, so that the steel is constantly protected.

Chapter 7 Ionic compounds

7.1 Properties and structures of ionic compounds

7.1 Key questions

Retrieval

1 B

2 *Electro* in electrostatic relates to the presence of charge. Both anions and cations are charged particles resulting from the transfer of electrons and an imbalance between numbers of electrons and protons. Anions have a negative charge and have gained electrons. Cations have a positive charge and have lost electrons. *Static* defines something that is not moving, and this describes the fact that the charge(s) on ions remain with the ion even though it has formed into an ionic compound.

Comprehension

3 Aluminium has an electron configuration of 2,8,3. This means that the most effective way for an aluminium atom to attain a noble gas configuration is to lose three electrons. Therefore, it is likely that aluminium will form an Al^{3+} ion.

4 **a** ionic
b not ionic
c not ionic
d ionic

5 The attractive force of the oppositely charged ions is greater than the repulsive force of two positively charged or two negatively charged ions near each other. The cation to anion ratio is 1 : 1 and the ions are of similar size. This makes the lattice very regular and oppositely charged ions are well shielded from each other.

6 Bonding in ionic substances is strong and has a lattice structure, which will not allow a force to change its shape (malleable) or to stretch it into a wire (ductile). The strength of the lattice means that it is brittle—the crystal will shatter when enough force is applied as the like-charged ions are forced next to each other and will thus repel each other.

7 The ratio of magnesium ions to fluorine ions is 1 : 2 and so the formula unit is MgF_2.

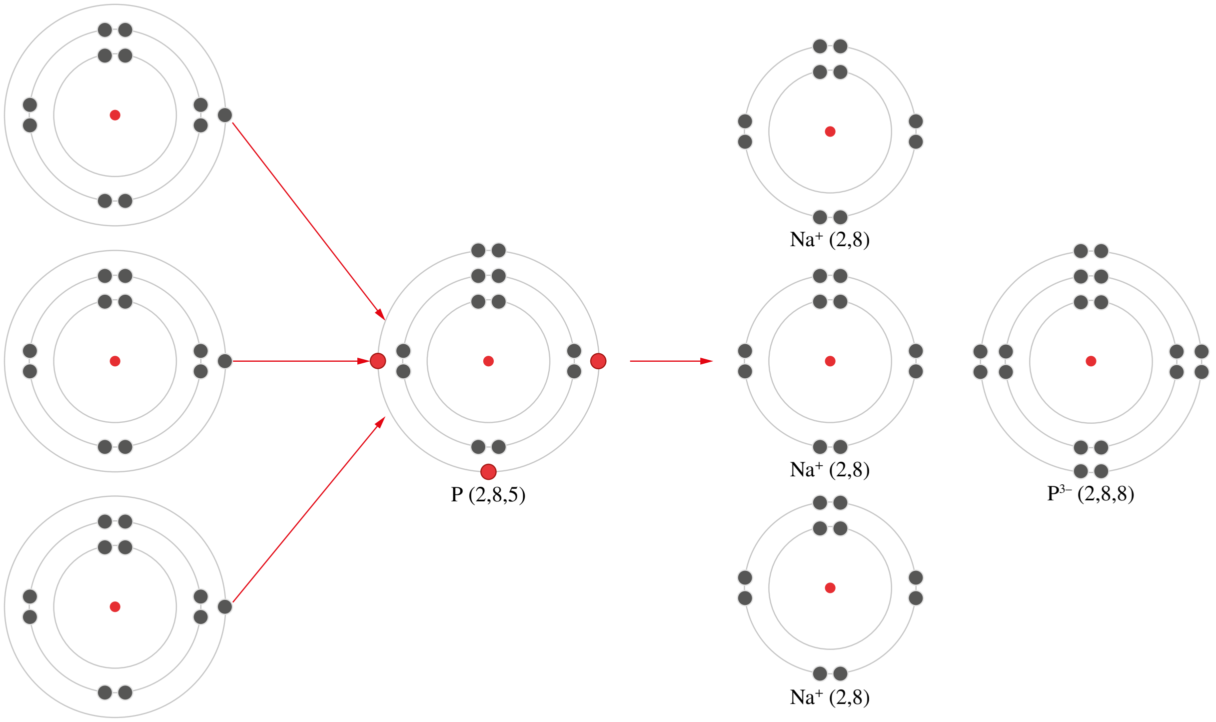
Analysis

8 Tap water has dissolved metal ions and so it is an ionic solution. The dissolved ions are able to move and carry charge and therefore, tap water is able to conduct electricity. This poses a danger when tap water is near a power source. Pure water containing no ions (e.g. deionised or distilled) is a covalent substance, which will not conduct electricity.

9 The lowest melting points are shown when chlorine is bonded to elements from groups 6 and 7, and the highest melting points are seen for compounds where chlorine is bonded to elements in groups 1 or 2. This is because (in period 3) groups 1 to 3 contain metallic elements, Si is a metalloid, and groups 5–7 contain non-metal atoms.
Because Cl itself is a non-metal it will form ionic compounds with elements from the first three groups (metals) and will form covalent compounds with elements from other groups (non-metals). The high melting points that are expected for ionic substances are clearly shown when chlorine bonds with atoms in groups 1–3 (NaCl 801°C, $MgCl_2$ 714°C, $AlCl_3$ 190°C), while lower melting points are shown for covalent substances ($SiCl_4$ −70°C, PCl_5 −162°C, SCl_2 −121°C, Cl_2 −101°C). PCl_5 is an exception to this trend. The electrostatic attraction between anions and cations in ionic substances is strong and a stable lattice structure is formed. So, a significant amount of energy is required to break these bonds, causing a high melting point.

7.2 Formation of ionic compounds

TY 7.2.1



Na (2,8,1) donates one electron each to to form Na^+ (2,8) and P^{3-} (2,8,8).
Both now have filled outer shells.

7.2 Key questions

Retrieval

1 K^+ 18 electrons; Cl^- 18 electrons; Mg^{2+} 10 electrons; H^+ 0 electrons; Al^{3+} 10 electrons; O^{2-} 10 electrons

Comprehension

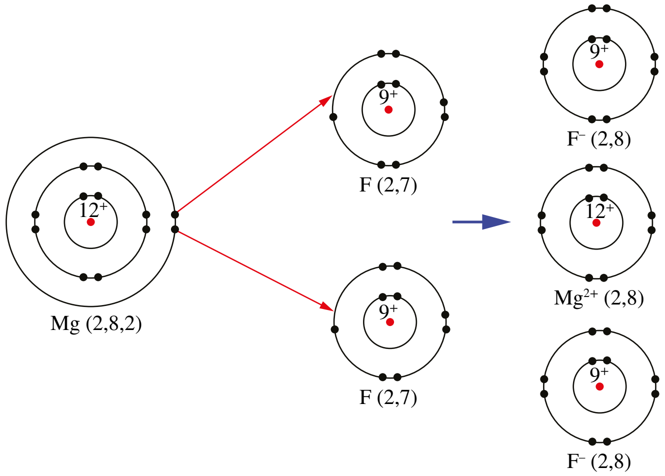
2 **a** Calcium is in group 2 of the periodic table and has an electron configuration of 2,8,8,2. It therefore needs to lose two electrons to have a noble gas configuration (eight electrons in the outer shell). Therefore, it will acquire a noble gas configuration and become a cation. It has a +2 charge because there are now 18 electrons and the nucleus still has 20 protons.

b Fluorine is in group 7 of the periodic table and has an electron configuration of 2,7. It therefore needs to gain one electron to have a noble gas configuration (eight electrons in the outer shell). It will acquire a negative charge and become an anion. It has a –1 charge because there are now 10 electrons and the nucleus still has nine protons.

c Aluminium is in group 3 of the periodic table and has an electron configuration of 2,8,3. It therefore needs to lose three electrons to have a noble gas configuration (eight electrons in the outer shell). Therefore, it will acquire a noble gas configuration and become a cation. It has a +3 charge because there are now 10 electrons and the nucleus still has 13 protons.

d Phosphorus is in group 5 of the periodic table and has an electron configuration of 2,8,5. It therefore needs to gain three electrons to have a noble gas configuration (eight electrons in the outer shell). It will acquire a negative charge and become an anion. It has a –3 charge because there are now 18 electrons and the nucleus still has 15 protons.

3



Formation of magnesium fluoride, MgF_2

4 Group 2 elements have two electrons in their valence shell and so these metals will lose these electrons when they form cations. By losing two electrons, they now have eight electrons in their outer shell, giving them a stable octet of valence electrons. Group 2 atoms will, after losing two electrons, have the electron configuration of the noble gas in the period before.

5 Copper is a transition metal that forms a cation as do other metals by losing electron(s). Cu forms more than one type of ion, namely Cu^+ or Cu^{2+}. This means that Cu will lose either one or two electrons. Copper can make cations of different charges because the *s*, *p* and *d* orbitals have energy levels that are close together.

6 **a** cation Ag^+; anion Br^-
b cation Fe^{2+}; anion I^-
c cation Na^+; anion N^{3-}

7 **a** $Na + Cl \rightarrow Na^+ (2,8) + Cl^- (2,8)$
b $Mg + O \rightarrow Mg^{2+} (2,8) + O^{2-} (2,8)$
c $2Al + 3S \rightarrow 2Al^{3+} (2,8) + 3S^{2-} (2,8,8)$

Analysis

8 **a** $Ca + F_2 \rightarrow Ca^{2+} + 2F^-$
$Ca^{2+} + 2F^- \rightarrow CaF_2$
b $2Na + O \rightarrow 2Na^+ + O^{2-}$
$2Na^+ + O^{2-} \rightarrow Na_2O$
c $Al + P \rightarrow Al^{3+} + P^{3-}$
$Al^{3+} + P^{3-} \rightarrow AlP$

7.3 Chemical formulas of simple ionic compounds

TY 7.3.1 BaF_2

TY 7.3.2 calcium bromide

7.3 Key questions

Retrieval

1 **a** lithium chloride **b** calcium oxide
c magnesium sulfide **d** potassium oxide
e sodium fluoride

Comprehension

2 **a** 2:1 **b** 1:3 **c** 3:2 **d** 1:1 **e** 1:2
3 **a** NaCl **b** KBr **c** $ZnCl_2$ **d** K_2O **e** $BaBr_2$
f AlI_3 **g** AgBr **h** ZnO **i** BaO **j** Al_2S_3
4 **a** potassium chloride **b** calcium oxide
c iron oxide **d** silver(I) iodide
e aluminium sulfide **f** mercury(II) bromide
g calcium nitride **h** copper(I) fluoride
i lead(IV) chloride

Analysis

5 **a** KCl **b** CaI_2 **c** $AlBr_3$ **d** MgO **e** Rb_3P **f** BeSe **g** Li_3N
6 A

7.4 Writing formulas of more complex ionic compounds

TY 7.4.1 $Ba(OH)_2$

TY 7.4.2 calcium carbonate

7.4 Key questions

Retrieval

1 permanganate MnO_4^-; oxalate $C_2O_4^{2-}$; ethanoate CH_3COO^-; hydrogen carbonate HCO_3^-

2 Elements on the left of the periodic table will appear on the left in the chemical formula. Elements on the right of the periodic table will appear on the right in the chemical formula.

3 **a** HSO_3^- **b** PO_4^{3-} **c** CO_3^{2-} **d** CN^-

Comprehension

4 **a** magnesium hydroxide **b** sodium carbonate
c iron(II) phosphate **d** copper(I) sulfate
e barium nitrate **f** copper(I) sulfate
g iron(III) ethanoate **h** ammonium nitrate
i sodium hydrogen phosphate

5 **a** Na_2CO_3 **b** $Ba(NO_3)_2$ **c** $Al(NO_3)_3$ **d** $Ca(OH)_2$ **e** $ZnSO_4$
f KOH **g** KNO_3 **h** $ZnCO_3$ **i** K_2SO_4 **j** $Ba(OH)_2$

6 **a** CuCl **b** Fe_2O_3 **c** $MgCr_2O_7$ **d** $Cr_2(SO_4)_3$
e $FeSO_3$ **f** $Pb(NO_3)_2$ **g** PbO_2 **h** $Sn(OH)_2$
There are the same numbers of cations and anions in **a**, **c** and **e**.

Analysis

7

Compound: ammonium hydrogen phosphate	
Formula: $(NH_4)_2HPO_4$	
Element	Number of atoms in formula unit
N	2
H	9
P	1
O	4

Compound: aluminium carbonate	
Formula: $Al_2(CO_3)_3$	
Element	Number of atoms in formula unit
Al	2
C	3
O	9

7.5 Properties of ionic substances

7.5 Science as a human endeavour

Nanocomputers

1 Examples include:
- in microphones and sensors where the pressure applied to the crystal is a soundwave, which then creates voltage
- in cigarette lighters and barbecue lighters where pressure on the crystal creates a spark.

2 In Figure 7.5.13, the 'ball' on the left of the image can be seen to contain two atoms: carbon and oxygen. The other carbon monoxide molecules are aligned vertically so that they appear as one atom. They are in this position because the scanning metal tip has allowed electrons to partially flow between the metal and non-metal atoms. This association is strong enough to allow lifting and movement of the non-metal atom(s), but not strong enough to create an ionic bond.

3 This temperature is as close as practicable to absolute zero (–273°C). Under these extreme conditions, the kinetic energy of all particles is extremely low (almost zero). At this temperature, the vibrating movement of all atoms in solids, liquids and gases is as close to zero as possible. This means that the electron tunnelling data sent from the scanning tip to the computer is not affected by vibrating and moving atoms, and will provide a clear picture of the structure of the surface.

7.5 Key questions

Retrieval

1 In solid form, the ionic compound forms a crystal lattice. This is a very strong structure as the strong electrostatic forces of attraction between the positively charged cations and negatively charged anions mean that the ions are not free to move. When heated so the ionic compound is now molten, the charged particles are free to move and so are able to conduct electricity.

2 The electrostatic forces of attraction between the positive and negative ions holding the lattice together are very strong and a lot of energy is required to break them apart.

Comprehension

3 **a** When hit with a hammer or hard object, alignment of anions and cations is slightly shifted.

b When the alignment of anions and cations is altered by a heavy blow, like-charged ions line up adjacent to each other. This causes the ions in the lattice to electrostatically repel each other, resulting in the crystal shattering.

Analysis

4

Ore	Cation	Anion	Purified metal
bauxite	Al	O and OH	aluminium
sphalerite	Zn	S	zinc
haematite	Fe	O	iron

5

Property	What this tells us about structure
high melting point	Forces between the particles are strong.
hard, brittle crystals	*Forces between the particles are strong and like charges repel when moved closer to each other.*
does not conduct electricity in the solid state	No free-moving charged particles are present in the solid.
conducts electricity in the molten state or in aqueous solution	*In liquid or dissolved state, the cations and anions are free to move (have been released from the lattice).*
solubility varies between ionic substances	The strength of the interaction of ions with water particles varies as does the strength of bonds within the lattice.

6 **a** Using the trend (size of anion is inversely proportional to melting point), it is proposed that NaF will have a *higher* melting point than NaCl. This is due to F being smaller than Cl (refer to page 81 for trends in atomic size). In fact, the melting point of NaF is 993°C.

b The table of melting points of sodium-based compounds shows that sodium compounds with simple anions that are single atoms have a higher melting point than polyatomic anions where there are multiple atoms. This indicates that the bonding is stronger when there are fewer atoms in the anion. A possible reason for this is the difference in size between the cation and anion when there is a large polyatomic ion. The smaller cation may not be able to 'mask' the negative charge of surrounding larger anions and therefore the anions will repel each other. This is demonstrated in the figure shown where there is a size difference between the two ions in sodium chloride. This effect may be greatly magnified for a polyatomic ion composed of many atoms.

Chapter 7 Review

Retrieval

1 **a** 2,8 **b** 2,8 **c** 2,8 **d** 2,8

2 They represent the number of each atom in the formula unit and indicate the ratio of cations to anions.

3 **a** Dichromate $Cr_2O_7^{2-}$ has nine atoms.

b PO_4^{3-} phosphate has –3 charge.

4 **a** The bonds between the positive and negative ions are strong and will only be overcome when significant energy is applied (only at high temperatures).

b The strong electrostatic forces of attraction between the ions mean that a strong force is needed to break up the lattice, giving the ionic crystals the property of hardness. However, the crystal lattice will shatter when a strong force is applied, suddenly causing ions of like charge to become adjacent to each other and be repelled.

c In the solid state, the ions are not free to move. However, when the solid melts or dissolves in water, the ions are free to move and conduct electricity.

5 A and B

Comprehension

6 D

7 C

8 A

9 A and C

10 Refer to Chapter 6, page 142 for a summary of the properties of metals.

a both **b** both **c** both **d** ionic lattices only **e** both

11 top left: *non-metals*; *gain*
top right: *metals*; *lose*
bottom left: *ionic compound*
bottom right*: ionic bonding* OR *electrostatic forces of attraction*

12 **a** ammonium carbonate **b** copper(II) nitrate
c chromium(III) bromide

13 The strength of the forces between Na^+ and Cl^- ions remains unchanged, but the kinetic energy of the ions increases until the forces can no longer hold the ions in the solid lattice, and the lattice breaks up as the solid melts.

14 **a** Solid ionic compounds do not conduct electricity.
b Ionic compounds are hard.
c Ionic compounds conduct electricity when dissolved and formation of ions involves a gain/loss of electrons.

15 **a** Li (2,1) + Cl (2,8,7) → LI^+(2) + Cl^- (2,8,8)
b Mg (2,8,2) + 2F (2,7) → Mg^{2+} (2,8) + $2F^-$ (2,8)
c 2K (2,8,8,1) + S (2,8,6) → $2K^+$ (2,8,8) + S^{2-} (2,8,8)
d 3Mg (2,8,2) + 2N (2,5) → $3Mg^{2+}$ (2,2) + $2N^{3-}$ (2,8)

16 They have seven electrons in their outer shell so only need to gain one electron to satisfy the octet rule and have noble gas configuration. This means they become more negative by gaining one electron as the number of electrons is one greater than the number of protons.

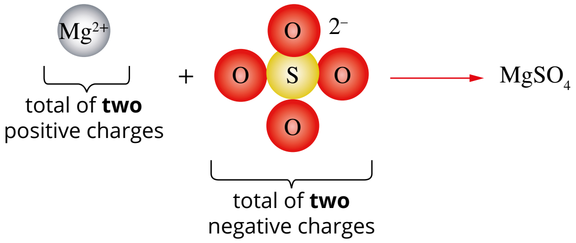
17 **a** KBr **b** MgI_2 **c** CaO **d** AlF_3 **e** Ca_3N_2

18 **a** CuCl **b** Ag_2O **c** Li_3N **d** KI

19

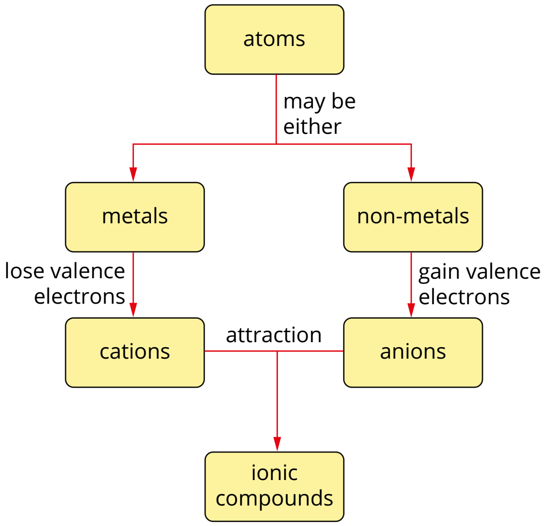
	K^+	Ca^{2+}
F^-	KF	CaF_2
N^{3-}	K_3N	Ca_3N_2

20



Formation of magnesium sulfate

21 (suggested answer, other variations are possible)



Analysis

22 The chemical formula of an ionic compound is known as the formula unit because it shows the simplest whole number ratio of cations to anions. It does not show the actual number of particles in the substance because this may be many thousands or millions of multiples of the formula unit (depending on the size of the crystal).

23 **a** A_2B **b** CD_3 **c** EF **d** H_3G **e** KL

24 **a** $MgCl_2$ or CaF_2 (any group 2 and group 17 using ratio shown)
b NaCl or CaS (any group 1 and group 17 OR any group 2 and group 16 OR any group 3 and group 15 using ratio shown)
c Na_2O or K_2S (any group 1 and group 16 using ratio shown)
d Na_3N or Li_3P (any group 1 and group 15 using ratio shown)
e $AlCl_3$ or BF_3 (any group 3 and group 17 using ratio shown)
f Mg_3N_2 or Ca_3P_2 (any group 2 and group 15 using ratio shown)

25 **a** $Mg + 2Cl_2 \rightarrow Mg^{2+} + 2Cl^-$
$Mg^{2+} + 2Cl^- \rightarrow MgCl_2$
b $2Al + 3O_2 \rightarrow 2Al^{3+} + 6O^{2-}$
$2Al^{3+} + 6O^{2-} \rightarrow Al_2O_3$

26 D

27 **a** Fe_2O_3 **b** CuO **c** $Cr_2(SO_4)_3$ **d** FeO
e $Pb(NO_3)_2$ **f** PbS_2 **g** $Sn(NO_3)_2$

28 **a** **i** –3 **ii** +1 **iii** –2
b **i** Y_2SO_4 **ii** K_2Z **iii** Y_3M **iv** Y_2Z

29 **a** $CuNO_3$ **b** CrF_2 **c** K_2CO_3 **d** $Mg(HCO_3)_2$ **e** $Ni_3(PO_4)_2$

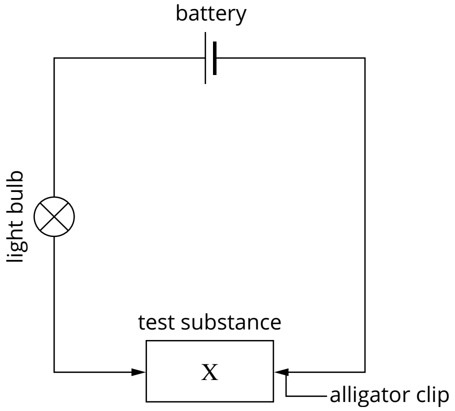
30 **a** Na^+, Cl^-; Mg^{2+}, O^{2-}
b MgO
c The strength of electrostatic attraction between ions will depend on the size of the ions and on their charge. The Mg^{2+} ion is slightly smaller than the Na^+ ion, and the O^{2-} ion is much smaller than the Cl^- ion. More importantly, the Mg^{2+} ion and the O^{2-} ions each have twice the charge of the Na^+ ion and the Cl^- ion. The attraction between the ions in MgO is therefore much stronger than in NaCl. Magnesium oxide therefore has a much higher melting temperature.

Knowledge utilisation

31 **a** $Ca(ClO_4)_2$ **b** $Al_4[Fe(CN)_6]_3$ **c** $Fe(ClO_4)_3$ **d** $(NH_4)_4Fe(CN)_6$

32 In sodium chloride, the lattice structure shows alternating sodium cations (Na^+) and chloride anions (Cl^-). Each ion is surrounded by six ions of the alternate charge. The ions are single and are of a similar atomic number (size) forming a regular array.
In copper sulfate, the lattice structure shows alternating copper cations (Cu^{2+}) and sulfate anions (SO_4^{2-}). Each ion is surrounded by six ions of the alternate charge. The anion is complex and contains five atoms compared to the cation, which is a single atom. This means that the physical arrangement will be different from sodium chloride, and the lattice structure is not as regular. This characteristic can be seen macroscopically (visible with the naked eye) in the diamond-shaped copper sulfate crystal compared to the cuboid sodium chloride crystals.

33 **a** Assemble equipment to test conductivity. Add a globe to a circuit containing a power source such as a battery. When the electrodes are touching the solid magnesium chloride, the globe will not light up.



b Dissolve some solid sodium chloride (about a teaspoon per 200 mL) in deionised water. Using the same equipment, place the two electrodes in the solution but don't allow them to touch; the globe will glow.

c If a crystal of sodium chloride was hit firmly with a hammer, it would shatter. Care is needed—safety glasses must be worn.

34 **a** Agree. In a metallic lattice, the positive ions are surrounded by delocalised electrons; in an ionic lattice, negative ions alternate with the positive ions.

b Agree. In a metallic lattice, each positive ion attracts the delocalised electrons in its region, and each delocalised electron is attracted to all neighbouring positive ions. In an ionic lattice, each positive ion is attracted to the negative ions that surround it, and vice versa.

c Agree. In a metallic solid, there will be repulsion between the positive ions, and between the delocalised electrons. The particles are arranged to minimise these repulsions. In an ionic lattice, the arrangement of alternating positive and negative ions also minimises repulsion between like charges.

d Disagree. In a metal, delocalised electrons are free to move so it conducts electricity; in an ionic solid, the ions are not free to move so it does not conduct electricity.

Chapter 8 Covalent compounds

8.1 Properties of non-metallic substances

8.1 Key questions

Retrieval

1 **a** This is the ionic compound calcium sulfate. The non-metallic elements are sulfur (S) and oxygen (O).

b This is the molecular compound dinitrogen pentoxide. The non-metallic elements are nitrogen (N) and oxygen (O).

c This is the molecular element hydrogen. The only element present is hydrogen and it is a non-metal.

d This is the elemental metal copper. There are no non-metals represented here.

e This is the ionic compound iron(II) chloride. The non-metallic element present is chlorine (Cl).

f This is the non-metallic element iodine. The only element present is iodine (I) and it is a non-metal.

2 A group of two or more atoms covalently bonded together, and representing the smallest fundamental unit of a chemical compound.

Comprehension

3 **a** Non-metallic elements and compounds tend to consist of discrete molecules and do not have mobile ions or delocalised electrons as part of their structure. The absence of mobile charge carriers is the reason that nonmetallic substances tend not to conduct electricity and act as good electrical insulators. Note: one class of notable exceptions is the graphene covalent network structures of carbon (see Module 8.3 for more details).

b Non-metallic substances tend to consist of discrete molecules with relatively weak intermolecular bonds between individual molecules. These weak bonds are easily disrupted and broken with the addition of small amounts of heat and result in low melting points and low boiling points.

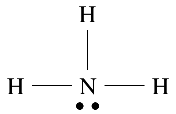
Analysis

4 Melting points and boiling points of molecular compounds (such as methane and water) are determined by the strength of the intermolecular bonds, between individual molecules. The stronger the intermolecular bonds the higher the melting points and boiling points. Given that the boiling point for water is 100°C and the boiling point for methane is –161°C, the intermolecular bonds holding the molecules of water together are clearly much stronger than the bonds holding methane molecules together. In the case of water, the intermolecular bond is called hydrogen bonding and you will explore this idea in more detail in Chapter 12.

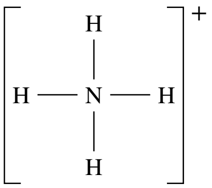
5 When heated gently, the solid is transformed from its solid state to its liquid state. This process is known as melting. During melting, the relatively weak intermolecular bonds are disrupted (or broken to a certain extent), allowing the individual sucrose molecules to roll over each other and allowing the bulk material to flow, i.e. the sucrose is in its liquid state. When the sample is heated even further, a chemical reaction takes place forming carbon dioxide gas and black elemental carbon (or soot). In this chemical reaction, the strong intramolecular bonds of sucrose are broken and the atoms are rearranged to form new substances.
A physical change of state, such as melting, involves disrupting (or breaking) the intermolecular bonds between molecules; a chemical change (i.e. a chemical reaction) involves breaking the intramolecular bonds within molecules.

8.2 Covalent bonding

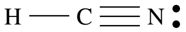
TY 8.2.1



TY 8.2.2



TY 8.2.3



8.2 Key questions

Retrieval

1 Many atoms become more stable if they gain an outer shell of eight electrons when they combine with other atoms. This general observation is summarised by the octet rule: during chemical reactions atoms tend to gain, lose or share electrons so that there are eight electrons in the outer shell.

2 A covalent bond is the force of attraction that exists between two atoms as a result of one or more pairs of electrons being shared between two nuclei.

3 In a single covalent bond, one pair of electrons is shared between two nuclei. In a double covalent bond, two pairs of electrons are shared between two nuclei. In a triple covalent bond, three pairs of electrons are shared between two nuclei.

4 Ethane has a single bond between its two carbon atoms, while ethene and ethyne have a double bond and triple bond respectively between their two carbon atoms. Carbon–carbon bond lengths decrease with increasing number of bonds between the two atoms, while C–C bond strengths increase with increasing numbers of bonds between the two atoms.

a The order of increasing bonding length (shortest to longest) is therefore ethyne–ethene–ethane.

b The order of increasing bond strength (weakest to strongest) is therefore ethane–ethene–ethyne.

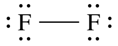
Comprehension

5 **a** one **b** three **c** two **d** one

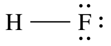
6 **a** one **b** two **c** three **d** four **e** one

f zero (Neon is a noble gas and tends not to undergo chemical reactions or form chemical bonds.)

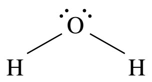
7 **a**



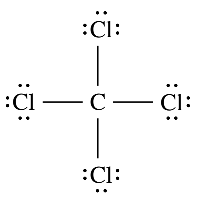
b



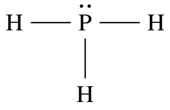
c



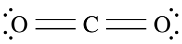
d



e



f



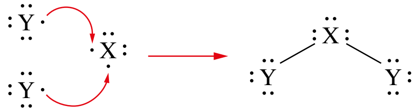
Analysis

8 **a** three bonding pairs and one non-bonding pair surround the central P atom
b four bonding pairs and zero non-bonding pairs surround the central N atom
c three bonding pairs and one non-bonding pair surround the central P atom
d two bonding pairs and two non-bonding pairs surround the central S atom
e four bonding pairs and zero non-bonding pairs surround the central C atom

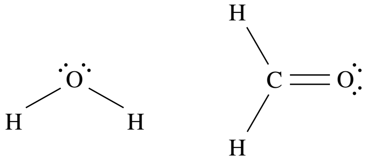
9 **a** ball-and-stick model
b Lewis structure
c space-filling model

10 **a** CCl_4 **b** NBr_3 **c** SiO_2 **d** HF **e** PF_3

11 XY_2. Element X has six valence electrons and therefore requires two more electrons to complete its octet. Element Y has seven valence electrons and therefore requires one more electron to complete its octet. Both atoms can complete their octets by bonding in a 1 : 2 ratio as shown:



12 The Lewis structures for both molecules are shown here.



Both oxygen atoms have two non-bonding pairs and two bonding pairs of electrons, completing the octet. Oxygen is a period 2 non-metal and will always obey the octet rule. The main difference between the structures is that the two bonding pairs of electrons in the H_2O molecule are represented as two single bonds, whereas the two bonding pairs of electrons in the CH_2O molecule are represented as a double bond.

13 The main difference between these elements is that Argon is a monatomic element, while chlorine is a diatomic element. Argon has the electron configuration of $1s^22s^22p^63s^23p^6$ and therefore has eight electrons in its outermost valence shell. The electron dot diagram for argon is shown (on the left below) and you can see that a single argon atom obeys the octet rule and will therefore be stable as a monatomic element. Chlorine on the other hand has the electron configuration of $1s^22s^22p^63s^23p^5$ and therefore has seven electrons in the outermost valence shell. The electron dot diagram for chlorine is shown below (in the middle) and you can see that a single chlorine atom requires another electron to complete its octet. It can do this by bonding to another chlorine atom to form the diatomic Cl_2 molecule where both Cl atoms have a complete octet in their valence shell (see below on the right).

8.3 Carbon lattices and carbon nanomaterials

8.3 Key questions

Retrieval

1 One of the characteristic features of non-metals is a lack of electrical conductivity. In other words, non-metals tend to be very good electrical insulators. This is not the case for some forms of carbon including graphite, graphene and carbon nanotubes. These forms of carbon have delocalised electrons spread across the covalent lattice structure, which can become mobile in the presence of an electric field and therefore conduct electricity.

Comprehension

2 Carbon can readily form single, double and triple covalent bonds with a wide range of non-metals including itself. It can also form long carbon chains as well as large carbon network structures. No other element on the periodic table has this capacity.

3 **a** A substance sublimes when it physically transforms from its solid state to its gaseous state without going through a liquid state.
b The high sublimation temperatures of diamond and graphite are due to the very strong network of carbon atoms covalently bonded together in large lattice structures. Diamond and graphite are not small molecular substances; they therefore do not have any weak intermolecular bonding that lead to low melting points and low boiling points.

4 **a** In diamond, the carbon atoms are bonded to each other in a continuous three-dimensional network with each carbon atom bonded to four others in a tetrahedral arrangement. There are no small, discrete molecules and therefore no weak intermolecular bonds—only strong covalent bonds are present. This is what gives diamond its hardness and strength. In graphite, each carbon atom is covalently bonded to three other carbon atoms in a two-dimensional lattice. While the bonds are very strong between the carbon atoms in each layer, there are weak intermolecular bonds between each layer. This makes graphite hard in one direction but soft and slippery in another direction.
b Electrical conductivity requires the presence of mobile charge carriers—either ions or electrons. In diamond, there are no ions and no delocalised electrons that are freely mobile across the bulk of the substance—the four valence electrons of each carbon atom are locked up in four covalent bonds. Diamond is therefore a poor electrical conductor and an excellent insulator. In graphite, each carbon atom is bonded to three other atoms. The fourth valence electron from each carbon atom is delocalised across the lattice layer and in the presence of an electrical field these valence electrons are freely mobile. Hence graphite has a relatively high electrical conductivity and is a poor electrical insulator.

5 **a** In graphite, each carbon atom is covalently bonded to three other carbon atoms in a two-dimensional lattice. While the bonds are very strong between the carbon atoms within each layer, there are weak intermolecular bonds between each layer. The weak intermolecular bonds allow the different layers to readily slide over one another, imparting a slippery feel to graphite and allowing it to be used as a lubricant.
b In diamond, the carbon atoms are bonded to each other in a continuous three-dimensional network with each carbon atom bonded to four others in a tetrahedral arrangement. There are no small, discrete molecules and therefore no weak intermolecular bonds—only strong covalent bonds. This is what gives diamond its hardness and strength, allowing it to be used in cutting applications such as on the surface of saws and drill bits.

6 The bonding within a C_{60} buckyball network involves each carbon atom being bonded to three other carbon atoms, leading to an overall spherical shape comprising a series of interlocking hexagons and pentagons. The fourth valence electron of each carbon atom is delocalised across the lattice.

7 A graphene sheet can be considered a single layer of graphite. The carbon lattice network consists of each carbon atom being bonded to three other carbon atoms arranged in a series of interlocking hexagonal shapes to form a two-dimensional layer or sheet. The fourth valence electron of each carbon atom is delocalised across the lattice.

Analysis

8 In both graphite and fullerenes, the network of carbon atoms involves each carbon atom being bonded to three other carbon atoms. In the former case, the atoms are arranged in a series of interlocking hexagonal shapes in a two-dimensional layer—with the layers arranged in stacks; in the latter case, the atoms are arranged in a roughly spherical shape in a series of interlocking hexagons and pentagons. Both structures have delocalised electrons spread across the surface of the lattice, allowing for electrical conductivity. Weak intermolecular bonding between both structures allows for the lattices to readily slide over each other, enabling their use as lubricants (although only graphite is used for this application in practice).

9 A graphene sheet can be considered a single layer of graphite. Each graphene structure is a network of carbon atoms formed into interlocking hexagonal shapes. Each carbon atom in the structure is bonded to three other carbon atoms.
The fourth valence electron of each carbon atom is delocalised across the lattice. Graphite is a multilayered stack of graphene sheets.

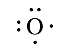
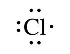
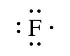
Chapter 8 Review

Retrieval

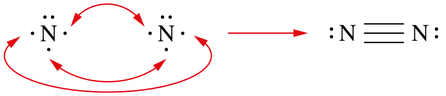
1 D

2 **a** 4 **b** 6 **c** 3 **d** 7 **e** 4

3

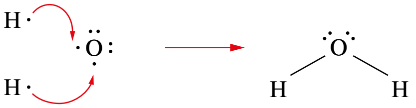
a	
b	
c	
d	
e	

4

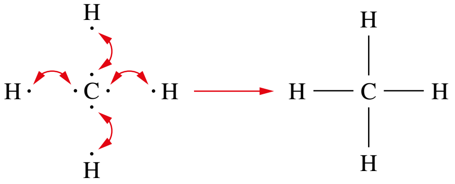
a	
b	
c	

5 **a** the formation of the HCl molecule from a H atom and a Cl atom

b the formation of the H_2O molecule from two H atoms and one O atom



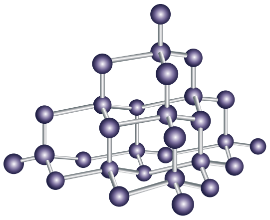
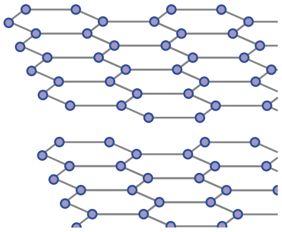
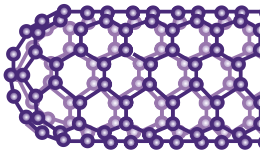
c the formation of the CH_4 molecule from one C atom and four H atoms



6 **a** true b true c false d true e false f true

7 E

8

				
Allotrope name	diamond	graphite	carbon nanotube	graphene
Bonding feature	• strong covalent 3D tetrahedral network lattice throughout	• strong covalent network where each carbon atom is bonded to three others to form a 2D hexagonal lattice sheet • weak intermolecular forces between sheets • delocalised valence electrons	• strong covalent 2D network lattice rolled into a cylindrical shape • delocalised valence electrons	• strong covalent 2D network lattice • delocalised valence electrons
Properties	• strong, hard, heat-resistant (very high sublimation point) • no electrical conductivity	• strong in one direction; weak, soft and slippery in the other direction • heat resistant (high sublimation point) • has appreciable electrical conductivity	• strong, heat resistant • electrically conductive	• strong, heat resistant • electrically conductive

Comprehension

9 Electrical conductivity in materials is the result of the presence of mobile charge carriers such as ions or delocalised electrons. Non-metallic substances, such as molecules, tend not to have either of these charge carriers in their structures. This means that, in general, non-metallic substances do not conduct electricity in their native form. Exceptions include the allotropes of carbon that have delocalised electrons as part of their non-metallic structure, including graphite, graphene, fullerenes and carbon nanotubes.

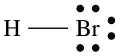
10 The intramolecular bonds within carbon dioxide molecules (i.e. the double covalent bonds between carbon and oxygen) are barely affected when carbon dioxide sublimes. A change of state such as sublimation does not involve the formation of any new substances and does not involve the breaking of covalent bonds. The carbon dioxide molecules remain intact during sublimation.

As carbon dioxide is heated towards its sublimation temperature, the intermolecular bonds between the molecules are disrupted and eventually broken, resulting in the substance transforming from its solid state to its gaseous state.

11 **a** 6 **b** 3 **c** 4

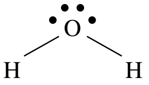
12 Neon is a noble gas with an electron configuration of $1s^2 2s^2 2p^6$. This means there are eight electrons in the valance shell, which is a stable electron arrangement. Since neon already has a stable valence shell arrangement there is no requirement to form covalent bonds (i.e. share electrons) to achieve a favourable electron configuration.

13 **a**



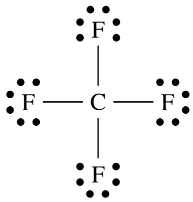
- 2 bonding electrons
- 6 non-bonding electrons

b



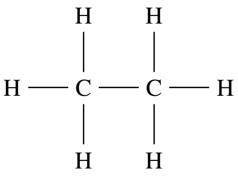
- 4 bonding electrons
- 4 non-bonding electrons

c



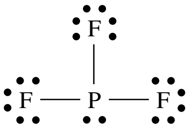
- 8 bonding electrons
- 24 non-bonding electrons

d



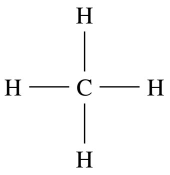
- 14 bonding electrons
- zero non-bonding electrons

e



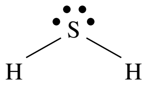
- 6 bonding electrons
- 20 non-bonding electrons

f



- 8 bonding electrons
- zero non-bonding electrons

g



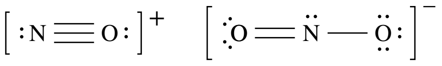
- sulfur: one bonding electron between each hydrogen and four non-bonding electrons
- hydrogen: one bonding electron

14 In diamond, the carbon atoms are bonded to each other in a continuous three-dimensional network with each carbon atom bonded to four others in a tetrahedral arrangement. There are no small, discrete molecules and therefore no weak intermolecular bonds—only strong covalent bonds. This is what gives diamond its hardness and strength.

A graphene sheet can be considered a single layer of graphite. The carbon lattice network consists of each carbon atom being bonded to three other carbon atoms arranged in a series of interlocking hexagonal shapes to form a two-dimensional layer. The fourth valence electron of each carbon atom is delocalised across the lattice.

15 Carbon nanotubes are closely related to graphene and can be described as a graphene sheet rolled into a cylindrical tube with a diameter of around 1 nm and a length many millions of times longer. Like graphene, each carbon atom in a carbon nanotube is covalently bonded to three other carbon atoms. This covalent structure arranged in a cylindrical shape imparts great strength on these nanoparticles. The fourth valence electron of each carbon is delocalised across the nanotube structure, giving these particles relatively high electrical conductivity. Carbon nanotubes can be open ended or capped at either end with a half-fullerene structure.

16 The structure for the NO^+ ion is shown on the left and the structure for the NO_2^- ion is shown on the right.

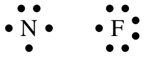


The NO^+ ion has three bonding pairs of electrons and two non-bonding pairs. The NO_2^- ion has three bonding pairs and five non-bonding pairs.

17 **a** X_2Y

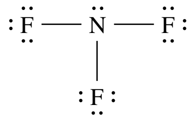
b water (H_2O), hydrogen sulphide (H_2S), hydrogen selenide (H_2Se)

18 The electron dot diagrams for nitrogen and fluorine are shown here.



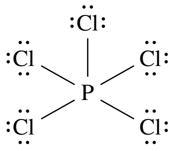
You can see that nitrogen requires three electrons to complete its octet and attain a stable valence shell configuration. You can also see that fluorine needs one electron to complete its octet and gain a stable electron arrangement. When nitrogen and fluorine combine, you can predict that three fluorine atoms can share one electron each with a nitrogen atom to complete the octets of all atoms. The chemical is therefore predicted to be NF_3.

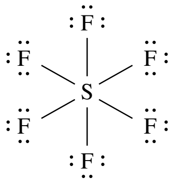
The Lewis structure for the compound that forms between N and F atoms is

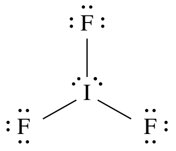


which confirms the prediction

19







All structures break the octet rule for the central atom. All central atoms in these structures are period 3 elements with unfilled *d*-orbitals that can be used in bonding and so can accommodate more than an octet.

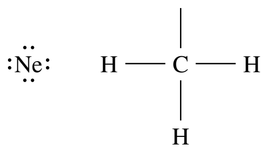
Analysis

20 Given that HCl exists as a gas at room temperature it indicates that the strength of intermolecular bonding is relatively weak. Room temperature provides enough energy to disrupt and ultimately break the attractive forces between individual HCl molecules. When compared to a substance such as water, you can say that, since water is a liquid at room temperature, the intermolecular bonding between water molecules is relatively strong compared with molecules like HCl. It should be said, however, that the forces that keep water molecules together are still very weak compared to the forces that keep ionic compounds together. This is reflected in the very high melting points of ionic compounds compared to molecular compounds.

21 B

22 Valence electrons are the electrons within atoms that are involved in bonding. They are often (but not limited to) the electrons that occupy the highest energy (or outermost) shell.
The electron configuration $1s^2 2s^2 2p^6 3s^2 3p^4$ represents the electron configuration of the element sulfur. Sulfur has six valence electrons, namely those six in the highest energy ($n = 3$) shell. In the electron configuration shown, these six valence electrons are represented by $3s^2 3p^4$.

23 The electron dot diagram for neon (Ne) and the Lewis structure for methane (CH_4) are shown.



Both Ne and the central carbon atom have eight valence electrons—satisfying the octet rule. Ne has four pairs of non-bonding electrons represented by the four pairs of dots; the central carbon atom in methane has four pairs of bonding electrons represented by the four straight lines. The electron configuration for the element Ne is stable and it does not need to gain, lose or share electrons to attain a stable arrangement. On the other hand, the electron configuration for elemental carbon is not stable and carbon needs to share electrons with other atoms to gain a stable arrangement.

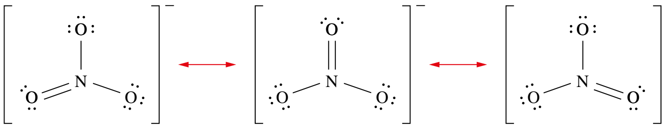
24 The boiling point of –44°C is a relatively low boiling point indicating that weak intermolecular bonding is present.
Ionic compounds tend to have quite high boiling points due mainly to the high strength of the electrostatic attraction between ions of opposite charge. Compound XY is most likely a molecular compound.

25 **a** zero **b** one **c** three **d** four

26 **a** three bonding pairs, one non-bonding pair around the central N atom
b four bonding pairs, zero non-bonding pairs around the central B atom
c four bonding pairs, zero non-bonding pairs around the central C atom.

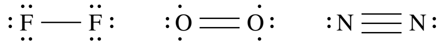
27 Lewis structures require all valence electrons from all atoms to be summed and accounted for. For polyatomic ions, the presence or absence of extra valence electrons needs to be considered. For anions (negatively charged ions), extra valence electrons need to be added; for cations (positively charged ions), valence electrons need to be subtracted.
For polyatomic anions, you need to add the value of the negative charge to the sum of valence electrons from all atoms involved; for polyatomic cations, you need to subtract the value of the positive charge from this total.
The final task required for polyatomic ions is that the entire Lewis structure needs to be placed inside square brackets with the charge of the ion being placed at the top-right of the brackets.

28



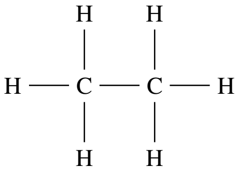
By ensuring that the central nitrogen atom conforms to the octet rule (and has eight valence electrons), the creation of one double bond is required. The question is: which of the three oxygen atoms should give up one of its non-bonding pairs to create the double bond? The answer is that any one of the three is equally likely, so all three structures are possible and equivalent. These structures are known as resonance structures.

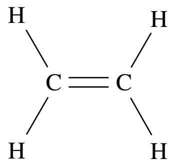
29 **a** The bonds in these molecules are similar in that each involves covalent bonding. Covalent bonding involves sharing of electrons between atoms to gain a stable outer-shell (or valence shell) electron configuration. For each of these molecules, the atoms are period 2 non-metals, which always comply with the octet rule when part of a compound.
Each atom in these molecules therefore has a total of eight electrons in its valence shell.
b The bonds differ in that: F_2 has a single covalent bond (where one pair of electrons are shared across both atoms); O_2 has a double covalent bond (where two pairs of electrons are shared across both atoms); N2 has a triple covalent bond (where three pairs of electrons are shared across both atoms).

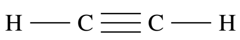


30 **a** ionic **b** metallic **c** molecular **d** molecular
e metallic **f** ionic **g** molecular

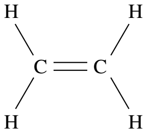
31 Each of these molecules includes two carbon atoms joined by covalent bonding with peripheral hydrogen atoms completing each structure. This means that all atoms in each structure share valence electrons to ensure that all end up with a favourable valence shell electron configuration.
The main difference is the number of covalent bonds holding the carbon atoms together. For ethane, a single covalent bond exists between the two carbon atoms; for ethene, a double covalent bond exists between the two carbon atoms; for ethyne, a triple covalent bond exists between the two carbon atoms.

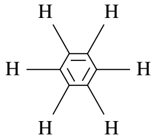

ethane

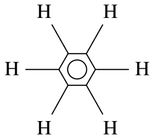

ethene


ethyne

32 The Lewis structure for ethene (C_2H_4) is shown below on the left and is characterised by the presence of a double covalent bond between the two carbon atoms. The C–C bond length for a double bond is ~134 pm, which is shorter than the C–C bond length for a single bond at ~154 pm.
Lewis structures for benzene (C_6H_6) are shown in the middle and right diagrams below. In the middle structure benzene is represented as a cyclic hexagon structure with alternating single and double bonds. This structure suggests that the carbon–carbon bond lengths in benzene are different with the single bonds being longer (i.e. ~154 pm) than the double bonds (~134 pm).
However, spectroscopic analysis confirms that all C–C bond lengths in benzene are equivalent and have a length of ~140 pm, somewhere between that of a double and single bond. A better representation for benzene is shown on the right, where you see that the six electrons associated with the three double bonds are actually spread (or delocalised) across the entire cyclic structure, making each C–C bond equal in both length and strength and a hybrid between a single and double bond.


ethene


benzene: represented with alternating single and double bonds between the carbon atoms


benzene: represented as equivalent bonds between the carbon atoms

33 **a** In the methane molecule, there are strong single covalent bonds between the carbon and hydrogen atoms (i.e. intramolecular bonding). There are also very weak intermolecular bonds between the molecules. In diamond, the carbon atoms are bonded to each other in a continuous three-dimensional network with each carbon atom bonded to four others in a tetrahedral arrangement. There are no small, discrete molecules and therefore no weak intermolecular bonds—only strong covalent bonds throughout the entire structure.

b The bonding in diamond explains why this substance is so hard and strong and has a very high sublimation point.

The absence of small discrete molecules means that there are no weak intermolecular bonds holding molecules together—if there were, then diamond would have low melting and boiling points and likely be soft in the solid state.

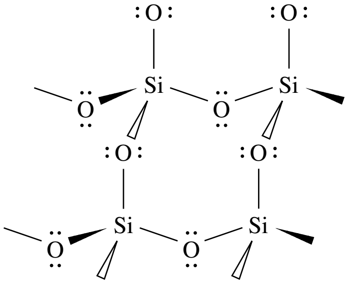
While the intramolecular bonding in methane is strong, the fact that it exists as small discrete molecules with weak intermolecular bonds is the reason it exists as a gas at room temperature and why it has very low melting and boiling points.

Knowledge utilisation

34 Structure A is the most appropriate. The Lewis structure for COF_2 requires the creation of one double bond to ensure that the central carbon atom has eight valence electrons in its valence shell. The placement of the double bond involves a choice between placing it between the C and O atoms or between the C atom and one of the F atoms. Since fluorine is more electronegative than oxygen it is less likely to give up a pair of non-bonding valence electrons to form the double bond (remember, electronegativity is a measure of an atom's ability to draw valence electrons to itself). The double bond will therefore form between the carbon and oxygen atoms.

35 A. The octet rule is not obeyed for the central boron atom in structure A but is obeyed in structure B. Despite this, structure A is still the preferred structure. The reason for this is that structure B requires one of the F atoms to give up a non-bonding pair of electrons to form a double bond—given the large differences in electronegativity between fluorine and boron there is very little chance of this happening. Remember, electronegativity is a measure of an atom's ability to draw valence electrons to itself, and fluorine, highly electronegative, is unlikely to share electrons to form the double bond. In this case the more stable structure is the one where the octet rule is broken for the central boron atom. Spectroscopic analysis reveals that bond lengths are identical for all three B–F bonds in boron trifluoride. This provides further evidence that structure A is the most appropriate since structure B would reveal that one of the B–F bonds would be shorter than the other two, given the presence of a double bond.

36 The high melting and boiling points of SiO_2 indicate that the structure is unlikely to involve the existence of small discrete molecules held together by weak intermolecular bonds. Instead, the structure is more similar to a diamond structure, where each silicon atom is bonded to four oxygen atoms in a 3D tetrahedral network lattice structure (see below). Like diamond, this structure is responsible for physical properties such as high resistance to heat, poor electrical conductivity and high hardness—SiO_2 is rated as 7 on the Mohs hardness scale (not as hard as diamond but still very hard).



Chapter 9 The mole

9.1 The mass of particles

TY 9.1.1 A_r(B) = 10.81 **TY 9.1.2** M_r = 63.02 **TY 9.1.3** 187.57

9.1 Key questions

Retrieval

1 Atomic mass is the mass of an atom in atomic mass units. One atomic mass unit (amu) is defined as a mass exactly equal to $\frac{1}{12}$ the mass of a ^{12}C atom.

2 M_r

3 The atomic mass of an atom is mainly determined by its number of protons and neutrons.

4 In both cases, they are numerically equal to the sum of the relative atomic masses of the elements in the formula.

Comprehension

5 The atomic mass is measured in amu, which is a comparative scale. Mass spectroscopy is used to compare the mass of atoms to a reference atom, which by international agreement is ^{12}C. ^{12}C is given a value of 12 amu, so 1 amu is equal to $\frac{1}{12}$ the mass of a ^{12}C atom. Atoms are then compared in size to the reference atom and given a value accordingly. For example, hydrogen is 8.400% the size of the reference so has an atomic mass of 1.008 amu, which is 8.400% of 12.

6 $CuSO_4$ is an ionic compound and so does not exist as a molecule. Rather than a relative molecular mass, it has a relative formula mass based on the elements in its representative formula.

Analysis

7 Atomic mass is the mass of an atom, in amu, compared to the mass of ^{12}C. Relative atomic mass is the weighted average of the atomic mass of different isotopes of an atom according to their relative abundance.

8 No, because the relative atomic mass is a weighted average of the two isotopes, based on their relative abundance.

9 $A_r = 12.011$

10 32.06

11 **a** 98.08 **b** 17.04 **c** 30.08

12 **a** 74.55 **b** 105.99 **c** 342.14

9.2 Introducing the mole

TY 9.2.1 9.6 × 10^{23} molecules **TY 9.2.2** 8.4 × 10^{23} atoms

TY 9.2.3 0.0013 mol **TY 9.2.4** 1.5 mol

TY 9.2.5 496 g **TY 9.2.6** 7.40 × 10^{21} molecules

9.2 Key questions

Retrieval

1 Avogadro's number

2 The mole is a convenient quantity for counting particles and is given the symbol *n* and the unit *mol*. One mole is defined as the amount of substance that contains the same number of 'specified' particles as there are atoms in 12 g of carbon-12. The number of particles in 1 mole is given the symbol N_A. This is known as Avogadro's number and has the numerical value of 6.02×10^{23}.

3 False. Molar mass is given the symbol *M*, not *m*. The unit g mol^{-1} is correct for molar mass.

4 $n = \frac{m}{M}$

Comprehension

5 It could mean one mole of oxygen atoms (O) or one mole of oxygen molecules (O_2).

6 The molar mass of a compound can be calculated by adding the relative atomic masses for each atom present in the molecular or ionic formula. This value is then expressed in grams.

Analysis

7 **a** 1.2×10^{24} atoms **b** 6.02×10^{22}
c 6.02×10^{21} atoms **d** 2.78×10^{19} molecules
8 **a** 0.5 mol **b** 0.25 mol **c** 70 mol
9 **a** **i** 1.7×10^{-4}
ii 1.7×10^{-4}
iii 1.7×10^{-4}
b They take up different amounts of space and have different masses.
10 **a** 0.8 mol **b** 4.8 mol **c** 0.72 mol **d** 6.0 mol
11 **a** 28.02 **b** 17.04 **c** 98.08 **d** 241.88
e 60.06 **f** 32.06 **g** 176.14 **h** 249.71
12 **a** 23 g **b** 64 g **c** 1.6 g **d** 25 g
13 **a** 5 mol **b** 2.5 mol **c** 0.10 mol **d** 0.025 mol
e 0.0031 mol **f** 0.0063 mol **g** 9.7×10^{-6} mol **h** 3.9×10^{-5} mol
14 **a** 6.02×10^{23} atoms **b** 6.02×10^{22} atoms
c 6.02×10^{21} atoms **d** 3.1×10^{22} atoms
15 **a** **i** 3.01×10^{23} molecules
ii 6.02×10^{22} molecules
b 6.02×10^{22} atoms. **c** 4.08×10^{25} atoms

9.3 Percentage composition and empirical formulas

TY 9.3.1 35% **TY 9.3.2** MgO **TY 9.3.3** N_2O_5 **TY 9.3.4** C_4H_{10}

9.3 Key questions

Retrieval

1 The percentage composition of a given compound tells you the proportion by mass of the different elements in that compound.
2 H_2SO_4, because it is the only formula with the simplest whole number ratio of elements
3 true

Comprehension

4 Formulas tell us the proportions of the atoms, but different atoms have different masses.
5 Ionic compounds do not have molecular formulas because they do not exist as molecules. They do have empirical formulas that describe the fixed ratio of ions that exist in their lattices.

Analysis

6 **a** 69.94% **b** 84.87% **c** 26.18% **d** 51.18%
7 **a** HCl **b** CO
8 **a** MgO **b** CH_4
9 **a** H_2O_2 **b** $C_3H_6O_3$ **c** NO_2 **d** $C_{11}H_{22}$
10 **a** CH_2 **b** C_5H_{10}
11 **a** CH_2O **b** $C_6H_{12}O_6$

9.4 Chemical equations and the mole

TY 9.4.1 $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

TY 9.4.2 0.60 mol

TY 9.4.3 1.484 g

TY 9.4.4 **a** $Pb(NO_3)_2$ is the limiting reactant (it will be completely consumed).
b 0.834 g

TY 9.4.5 79.00%

9.4 Key questions

Retrieval

1 Chemical equations are a method that chemists use to describe chemical reactions. To write a chemical equation you write the formulas of the reactants on the left and the formulas of the products on the right.
2 This equation is not balanced.
3 the study of ratios of moles of substances
4 False. The reactant that is completely consumed is the limiting reagent.
5 percentage yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times \frac{100}{1}$
6 Some reasons why actual yield varies from theoretical yield are:

- some reactions do not go to completion
- the reaction rate is too slow for the reaction to proceed to completion in the time available
- the loss of reactants and products in small quantities occur during transfer between reaction vessels and in separation and purification processes.

Comprehension

7 This is because chemical reactions obey the law of conservation of mass. The law of conservation of mass states that in any physical change, or chemical reaction, matter can neither be created nor destroyed. As such the number of each atom must be equal on the reactant and product sides of the equation.
8 **a** $N_2(g) + O_2(g) \rightarrow 2NO(g)$
b $Sn(OH)_4(s) \xrightarrow{\Delta} SnO_2(s) + 2H_2O(l)$
c $Mg(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Mg(NO_3)_2(aq)$
9 Stoichiometric calculations allow you to use the mole ratio, established in a chemical equation, to predict the amount of a product that will be formed or how much reactant will be used. The moles or mass of one of the reactants or products will be the known chemical and is used to determine the moles or mass of the unknown chemical. You can write the relationship between the known and the unknown chemicals using ratios:
$$\frac{n(\text{unknown chemical})}{n(\text{known chemical})} = \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$$
10 Percentage yield compares the actual yield to the theoretical yield. As such the actual yield is divided by the theoretical yield. This value is then multiplied by 100 to express it as a percentage.

Analysis

11 8.37 g
12

Nitrogen molecules available	Hydrogen molecules available	Ammonia molecules produced	Nitrogen molecules in excess	Hydrogen molecules in excess
2	10	4	0	4
879	477	318	720	0
9 mol	6 mol	4 mol	7 mol	0 mol

13 **a** $2Na(s) + Cl_{2\,(g)} \rightarrow 2NaCl(s)$
b 63.7 g
14 82.4%
15 247 g
16 **a** 0.020 mol **b** $\frac{2}{1}$ **c** 8.52 g
17 **a** $Pb(NO_3)_2$ is in excess by 0.5 mol.
b $Pb(NO_3)_2$ is in excess by 1.75 mol.
c 2 mol of KI reacts with 1 mol of $Pb(NO_3)_2$. In this case, not all of the KI can react and some of it will be in excess. This means that $Pb(NO_3)_2$ is the limiting reactant.
The equation shows that 1 mol of PbI_2 precipitate will form for every 1 mol of $Pb(NO_3)_2$ that reacts.
$n(PbI_2) = n(Pb(NO_3)_2) = 0.00302$ mol
$m(PbI_2) = 0.00302 \times 461.02 = 1.39$ g

Chapter 9 Review

Retrieval

1 D
2 the mass, in grams, of one mole of a particular element or compound
3 The empirical formula of a compound gives the simplest whole number ratio of elements in that compound.
4 The law of conservation of mass states that, in any physical change or chemical reaction, matter can neither be created nor destroyed.

Comprehension

5 The relative atomic mass of carbon is the weighted average of the isotopic masses of all carbon isotopes (i.e. ^{12}C, ^{13}C and ^{14}C). Small amounts of ^{13}C and ^{14}C make this average slightly greater than 12, the isotopic mass of the ^{12}C isotope.

6 **a** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
b $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
c $3Zn(OH)_2(s) + 2H_3PO_4(aq) \rightarrow Zn_3(PO_4)_2(s) + 6H_2O(l)$

7 There are several factors that influence the amount of product that will be produced for a given reaction. Some reactions reach equilibrium rather than going to completion. A reaction may be slow, such that the complete yield cannot be collected in a practical time frame. Finally, reactants and product can be lost when being transferred between containers or filtered.

Analysis

8 $A_r(Pd) = 106.4$
9 **a** $A_r(Pd) = 39.96$ **b** $A_r(K) = 39.11$
10 **a** $n(H_2O) = \frac{4.50 \times 10^{23}}{6.02 \times 10^{23}} = 0.75\,mol$
b $n(CH_4) = 15.0\,mol$
c $n(Cl_2) = 3.82 \times 10^4\,mol$
d $n(C_{12}H_{22}O_{11}) = \frac{1}{6.02 \times 10^{23}} = 1.7 \times 10^{-24}$
11 **a** **i** 8.73×10^{23} molecules
ii 3.49×10^{24} atoms
b **i** 3.47×10^{23} molecules
ii $N(atoms) = N(H_2S) \times 3 = 3.47 \times 10^{23} \times 3 = 1.04 \times 10^{24}$ atoms
c **i** 9.21×10^{21} molecules
ii 4.61×10^{22} atoms
12 **a** 1.8 g **b** 58 g **c** 0.41 g **d** 389 g
13 **a** 0.10 mol **b** 0.0390 mol **c** 1.25 mol **d** 0.00167 mol
14 **a** **i** $n(P_4) = 0.034\,mol$
ii $N(P_4) = 2.04 \times 10^{22}$ molecules
iii total number of atoms = 8.2×10^{22} atoms
b **i** $n(S_8) = 0.292\,mol$
ii $N(S_8) = 1.76 \times 10^{23}$ molecules
iii total number of atoms = 1.41×10^{24} atoms
c **i** $n(HCl) = 0.0088\,mol$
ii $N(HCl) = 5.3 \times 10^{21}$ molecules
iii total number of atoms = 1.1×10^{22} atoms
15 62.0 g
16 $100\,g\,mol^{-1}$
17 **a** $40\,g\,mol^{-1}$ **b** $98\,g\,mol^{-1}$ **c** $44\,g\,mol^{-1}$ **d** $106\,g\,mol^{-1}$
18 **a** $M(antibiotic) = 1.25 \times 10^4\,g\,mol^{-1}$
b $n(antibiotic) = 1.6 \times 10^{-7}\,mol$
c number of molecules = 9.6×10^{16} molecules
19 **a** 93.8% **b** 40.0% **c** 20.0% **d** 59.9%
20 **a** CO_2 **b** C_2H_4O **c** C_6H_5Cl **d** C_7H_{16}
21 **a** C_2H_6O **b** C_2H_6O
22 $58.8\,g\,mol^{-1}$
23 $31.0\,g\,mol^{-1}$
24 **a** $n(KOH) = \frac{2}{1} \times n(Fe(OH)_2)$ **b** $n(FeSO_4) = \frac{1}{2} \times n(KOH)$
c $n(KOH) = \frac{2}{1} \times n(K_2SO_4)$ **d** $n(Fe(OH)2) = \frac{1}{1} \times n(FeSO_4)$
25 373.09 g
26

Carbon atoms available	Oxygen molecules available	Carbon dioxide molecules produced	Carbon atoms in excess	Oxygen molecules in excess
8	20	8	0	12
1000	3000	1000	0	2000
9 mol	6 mol	6 mol	3 mol	0 mol

27 **a** oxygen **b** 1.5 mol
28 **a** P_4O_4 is in excess by 2.78 g **b** 3.37 g **c** 1.26 g
d total mass of products = 3.37 + 1.26 = 4.63 g; 4.63 g of products formed plus 2.78 g unreacted P_4O_6 = 7.40 g, which is consistent with the total mass of reactants used initially
29 **a** 3.26 g **b** 72.09%

Knowledge utilisation

30 **a** D, F, E, A, C, B
b

	Metal	Oxygen
mass (g)	0.542	0.216
relative atomic mass	40.1	16.00
moles	0.0135 mol	0.0135 mol
ratio	1	1

c calcium

Chapter 10 Energy changes in chemical reactions

10.1 Energy change

10.1 Key questions

Retrieval

1 Chemical energy is the energy stored in the bonds between the atoms and molecules of a substance. The SI unit for energy is the joule, J.
2 q
3 Sublimation is a process in which a solid changes state directly to gas without passing through an intermediate liquid state.

Comprehension

4 In chemistry, the system is usually the chemical reaction, whereas the surroundings refer to everything else; for example, the beaker or test-tube in which the reaction takes place.

5 The system would be the chemical reaction taking place in the beaker, so the acid and the base represent the system.

6 The Celsius scale starts at the temperature at which water melts, while the Kelvin scale starts at absolute zero—which is the temperature at which particles have no kinetic energy. The two scales are similar in the amount of each increment. For example, a temperature change of 5°C and 5 K is the same value of the increase. Temperatures on the Celsius scale can be converted to kelvin using the equation T(K) = T(°C) + 273.

7 Heat flow results from the microscopic movement and collisions of particles. When heat moves between two systems, the kinetic energy of the warmer system's particles is transferred to the cooler system's particles. When the temperature of a liquid increases enough for the particles to have sufficient energy to overcome inter-particle forces evaporation will occur.

Analysis

8 Temperature is a measure of the average kinetic energy of the particles of an object, while heat is a measure of the flow of energy. Temperature differences between objects causes heat, while two objects may have the same heat but have different temperatures because of the number and type of particles they contain.

9 This is because the heat energy being absorbed by the substance is being used to break the intermolecular bonds. As a result the average kinetic energy, or temperature, of the substance does not increase.

10 **a** 180 kJ **b** $1.5 \times 10^3\,kJ$
c 0.0100 kJ or $1.00 \times 10^{-2}\,kJ$ **d** $2.0 \times 10^{-6}\,kJ$
11 Use $T(K) = T(°C) + 273$:
a 298 K **b** 89.8 K **c** 1723 K **d** $2.00 \times 10^{-1}\,K$
12 Use $T(K) = T(°C) + 273$:
a −228°C **b** −73.8°C **c** 39°C **d** 182.5°C

10.2 Exothermic and endothermic reactions

TY 10.2.1 $\Delta H = -147$ kJ

TY 10.2.2 $\Delta H = -543$ kJ; The reaction will be exothermic.

10.2 Science as a Human Endeavour

Cellular respiration

1 exothermic
2 Oxidation of the relatively stable C–C bond is difficult, and many intermediate steps are required to release the energy in a controlled way that does not damage the cell.
3 This is an open question, and there are many experiments in which the respiration of glucose can be studied and measured in yeasts. One way that the effect of temperature can be measured is to make yeast cultures using different temperatures of water. The amount of water, yeast powder, and sugar added should be controlled, and if the mixture is placed in a glass measuring cylinder then the height of the foam formed can be used as a measure of respiration. As the yeast respires it releases carbon dioxide, which forms bubbles in the foam. The height of the foam formed over time gives a comparative measure of rate.

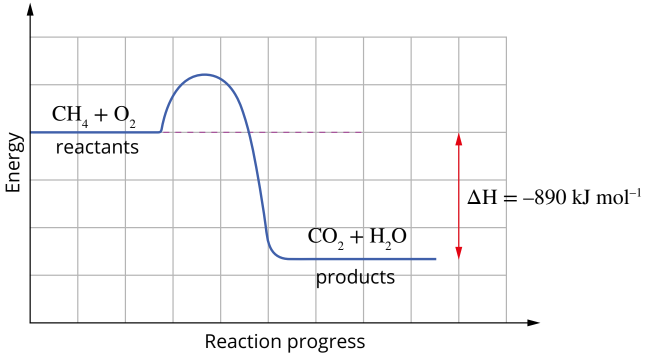
10.2 Key questions

Retrieval

1 The law of conservation of energy states, that during any chemical or physical process, energy is neither created nor destroyed.
2 Enthalpy (*H*) is the chemical energy of a substance.
3 A thermochemical equation shows enthalpy change by writing the ΔH value to the right of the chemical equation.

Comprehension

4 In any reaction, the total amount of chemical energy of the reactants is made up of the bonds between atoms within the reactants. If the total amount of chemical energy within the reactants is less than the total amount of the chemical energy within the products, energy must be supplied to the system, and the reaction is said to be endothermic.
5 Activation energy is the energy required to break the bonds of reactants. As both endothermic and exothermic reactions have reactants, both require activation energy.
6 Average bond enthalpies may differ from the specific molecules involved in your calculation because they assume a molecule is in a gaseous state and are an average of the bonds broken in different molecules.
7 $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = -890$ kJ mol^{-1}
8



Analysis

9 It would be lower because the change of state of the H_2O from liquid to gas will require energy to be absorbed.
10 a endothermic
   b $\Delta H = -1401.25$ kJ
11 a endothermic
   b The total enthalpy of the product (HI) is greater than that of the reactants (hydrogen gas and iodine gas).
   c The activation energy is greater than the ΔH value.

12 **a** B **b** A **c** A
13

Name	Formula	Number and type of bonds
hydrogen	H_2	1 × H–H
water	H_2O	2 × O–H
methane	CH_4	4 × C–H
oxygen	O_2	1 × O=O
ammonia	NH_3	3 × N–H
carbon dioxide	CO_2	2 × C=O

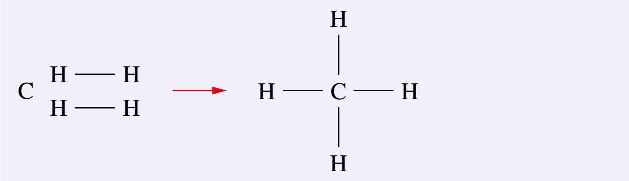
14 **a** $2HF(g) \rightarrow H_2(g) + F_2(g)$
H–F H–F → H–H F–F

		Average bond energy (kJ)
bonds broken	2 × H–F	2 × +567
bonds formed	1 × H–H	1 × –436
	1 × F–F	1 × –155
enthalpy change		+543

b H–H H–H O=O → H–O–H H–O–H

		Average bond energy (kJ)
bonds broken	2 × H–H	2 × +436
	1 × O=O	1 × +495
bonds formed	4 × O–H	4 × –463
enthalpy change		–484

c $C(g) + 2H_2(g) \rightarrow CH_4(g)$ $\Delta H = -780$ kJ mol^{-1}
Note: C(g) is already atomised (it is not bonded).



		Average bond energy (kJ mol^{-1})
bonds broken	2 × H–H	2 × +436
bonds formed	4 × C–H	4 × –413
enthalpy change		–780

10.3 Measuring energy changes

TY 10.3.1 $\Delta H = -111$ kJ mol^{-1}

TY 10.3.2 $\Delta H = -29.3$ kJ mol^{-1}

TY 10.3.3 $q = 12.1$ kJ (3 significant figures)

TY 10.3.4 Energy content = 21.8 kJ g (3 significant figures)

10.3 Key questions

Retrieval

1 False. The enthalpy of individual reactions can be added to give the enthalpy of reaction.
2 A calorimeter is an instrument designed to measure the energy changes in a reaction.
3 A solution calorimeter may be as simple as a polystyrene foam cup with a lid. The insulation provided by the polystyrene foam prevents the transfer of heat to or from the surroundings of the calorimeter. The reaction is carried out in the calorimeter with an accurately known volume of water. The initial and final temperatures are measured and recorded, as are the amounts of reactants used.

4 The heat of combustion is defined as the enthalpy change that occurs when a specified amount of the fuel burns completely in oxygen.

Comprehension

5 Under appropriate conditions chemical reactions are reversible. Δ*H* for an exothermic reaction is negative. Δ*H* for the reverse endothermic reaction is positive and when reactions are reversed, the sign will change. Equations for chemical reactions can be added and subtracted in the manner of ordinary algebraic equations such that complex chemical equations can be reduced to a series of simpler equations of formation with known enthalpies of reaction.

6 Specific heat capacity reflects the types of bonds holding the molecules, ions or atoms together in the substance. In the case of metals, the high conductivity results in lower specific heat capacity.

Analysis

7 $\Delta H = -114.1\text{ kJ}$

8 $\Delta H = +130\text{ kJ mol}^{-1}$

9 **a** $q = 6.35\text{ kJ}$ **b** $q = 25.1\text{ kJ}$ **c** $q = 62.7\text{ kJ}$

10 22.3°C

11 $\Delta H = -12.8\text{ kJ}$

12 energy content = 16.0 kJ g^{-1}

Chapter 10 Review

Retrieval

1 In chemistry, the system is usually the chemical reaction.

2 Temperature is a measure of the average kinetic energy of the particles of an object

3 Heat is the energy that flows from one object to another because of a difference in temperature.

4 **a** true **b** false **c** false **d** true

5 C

6 Specific heat capacity is a measure of the amount of energy (usually in joules) needed to increase the temperature of a specific quantity of that substance (usually 1 gram) by 1°C

7 B

Comprehension

8 Kinetic theory states that all matter consists of particles in motion, and as the temperature increases, the movement of the particles increases. Most substances can exist in all three states, and the state at a given temperature and pressure is determined by the strength of the inter-particle forces.

9 Heat always passes from a warmer object (fingers) to a cooler object (ice).

10 **a** exothermic, because heat and light energy are released to the surrounding environment by the combustion of wood

b endothermic, because thermal energy is absorbed from the surrounding environment to melt the ice

c endothermic, because electrical energy is consumed from a power supply as the battery is recharged

d exothermic, because heat energy is released to the surrounding environment and organisms in the compost heap decompose the plant material. The temperature of the heap rises as a consequence.

Analysis

11 Liquid water molecules are held together by hydrogen bonds, which are relatively strong and require energy to break.

12 **a** 2.21 kJ **b** 152 J **c** 1.89 MJ **d** 12.5 kJ

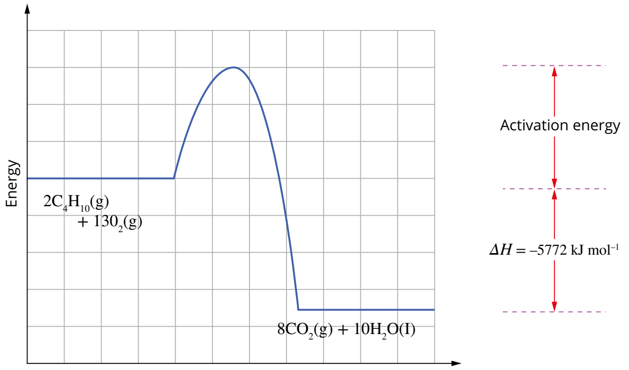
13 Both are types of vaporisation and involve a state change for liquid to gas. The difference is during evaporation, some of the particles have sufficient energy to break away, while during boiling all of the particles have the energy to break away.

14 **a** exothermic because Δ*H* is negative

b The value of Δ*H* would now be doubled as the amount of all the reactants and products is also doubled.
$\Delta H = -5238\text{ kJ}$.

15 **a** higher energy in bonds of reactants

b



16 $\Delta H = -305\text{ kJ mol}^{-1}$; The reaction will be exothermic.

17 If a chemical equation is written for an endothermic reaction, Δ*H* is positive, which tells you that energy is absorbed as the reaction proceeds. The enthalpy of the products must be higher than the enthalpy of the reactants.
If this reaction is reversed, the enthalpy of the reactants is now higher than the enthalpy of the products. The reaction releases energy as the reaction proceeds, so is exothermic. Δ*H* becomes negative.

18 $\Delta H = -853.9\text{ kJ}$

19 **a** $q = 25.1\text{ kJ}$ **b** $q = 78.4\text{ kJ}$ **c** $q = 62.7\text{ kJ}$
d $q = 788\text{ kJ}$ **e** $q = 54.1\text{ kJ}$

20 $q = -32.9\text{ kJ g}^{-1}$

21 energy content = 24.5 kJ g^{-1}

22 $q = 502\text{ J}$

Knowledge utilisation

23 **a** $q = 502\text{ J g}^{-1}$

b The heat of combustion calculated would be lower.
Some of the heat released by the burning wood would be used to heat the beaker holding the water and to heat the air surrounding the beaker. Not all of the energy will be used to heat the water, so the temperature rise will be lower than it should be, making the heat of combustion of the fuel seem lower than the actual value.
Also, the combustion of the wood may have been incomplete, with some carbon and carbon monoxide formed instead of carbon dioxide. As a consequence, less energy would have been released and the temperature rise would be lower.

24 **a** If the temperature of the water increases, it means it has absorbed heat from the combustion of the hydrocarbon.
An exothermic process is one that releases heat to its surroundings, so the combustion of the hydrocarbon mixture must be *exothermic*.

b If the combustion of the hydrocarbon mixture is exothermic, then it means the enthalpy (*H*) of the products of the combustion will be lower than the enthalpy of the reactants. Thus, according to the equation $\Delta H = H_p - H_r$, the Δ*H* will be negative.

c $q = 8.8 \times 10^1\text{ kJ}$

Chapter 11 Fuels

11.1 Fuels in society

11.1 Key questions

Retrieval

1 A fuel and oxygen react to produce carbon dioxide, water and energy. The energy is released in the form of heat, light, and sometimes sound. An example is the combustion of pentane. Energy is produced and so the reaction is exothermic:
$C_5H_{12} + 8O_2 \rightarrow 6H_2O + 5CO_2 + \text{energy}$

2 **a** 180 kJ **b** 1500 kJ **c** 0.01 kJ **d** $2.0 \times 10^{-6}\text{ kJ}$

Comprehension

3 **a** oil

b coal, oil (coal and oil amounts are similar), natural gas, biofuels, hydroelectricity, new technologies, nuclear (hydroelectricity, new technologies, nuclear amounts are similar)

4 $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$

Analysis

5 Wood cannot meet the needs of today's society because, having been used as an energy source for thousands of years, the amount of timber on Earth is dwindling. The energy needs of society are increasing. Global energy use has doubled in the last 50 years. The rate of growth of trees is not fast enough to supply the demand. Massive land area would need to be allocated. Also, three-quarters of our energy needs are provided by coal, oil and gas. This means that the infrastructure currently in use for fuel processing and use does not suit wood.

6 **a** Trends visible are:

1 All sectors of energy consumption have increased between 1976 and 2012.
2 The mining industry had the greatest proportional increase.
3 Between 2008 and 2012 all sectors decreased their energy usage.

b 1 As population and per capita energy use increases, all sectors of the community will increase their usage.
2 The use of fossil fuels has increased significantly (approximately double since 1960, Figure 11.1.2) and the industry has expanded to provide the energy.
3 Australia's commitment to reducing emissions and using clean energy has increased since 2000 and the government ratified the Kyoto Protocol in 2007, committing us to emissions targets. Government bodies have been set up and legislative changes such as a carbon tax and price on carbon have been proposed.

11.2 Fossil fuels

11.2 Key questions

Retrieval

1 The basis for separation of crude oil during fractional distillation is boiling point. The column has high temperatures at its base and lower temperatures at the top, and fractions are collected at different places as the vapours of the fraction condense.

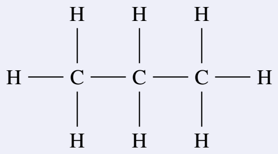
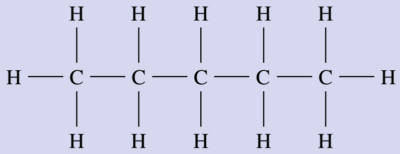
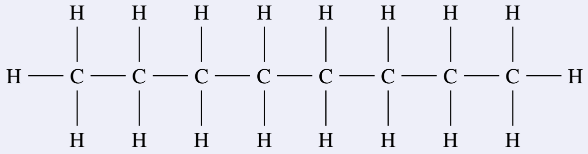
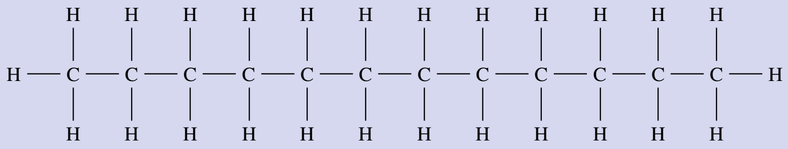
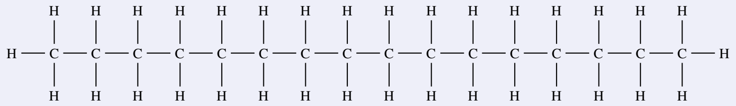
2 Molecules in diesel mostly have between 10 and 14 carbon atoms. Because diesel is a mixture, some components of diesel have 14 to 20 carbon atoms.

3 Natural gas is a fossil fuel because it is the product of degradation of ancient biomass. Large biomolecules have broken down to methane—the key hydrocarbon in natural gas.

4 **a** $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

b $2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(g)$

5

Fuels in fraction	Number of carbons	Example	Structural formula
LPG plastics industry feedstock	1–4	propane	
petroleum ether (petrol)	5–6	pentane	
naphtha petrol	6-10	octane	
kerosene petrodiesel	10–14	dodecane	
gas oil (petrodiesel)	14–20	hexadecane	

Comprehension

6 The crude oil fractional distillation column separates compounds on the basis of their boiling point—high boiling point molecules separate at the base and low boiling point molecules separate at the top. These molecules are hydrocarbons with a similar structure—a carbon chain with bonded hydrogen atoms.
There is little or no polarity in the structures and therefore the intermolecular forces are dispersion forces. Dispersion forces are greater for larger molecules and so these compounds will have a higher boiling point. Similarly, smaller molecules from the top of the column have lower boiling points due to lower intermolecular forces.

7 By using petrol as a fuel, cars are using chemical energy stored in the bonds of the fuel. These bonds are formed during photosynthesis in a plant, and the fuel matured over time underground. This means that the energy in petrol was derived from the sun.

8 **a** Black coal takes the longest time to form.
b Black coal is the better quality fuel because it has low moisture content and high energy density.

9 **a** $C_2H_5OH + 2O_2 \rightarrow 2CO + 3H_2O$ shows incomplete combustion because carbon monoxide is produced rather than carbon dioxide.
b $C_2H_5OH + O_2 \rightarrow 2C + 3H_2O$ shows incomplete combustion because carbon is produced rather than carbon dioxide.

Analysis

10 By considering the relative volumes of the reactants and products in the combustion of octane, explain the origin of the force that moves the pistons in the engine of a car.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g) \quad \Delta H = -10\,900\,\text{kJ}$$

During the combustion of octane the fuel, shown here as a liquid, quickly becomes a gas and reacts with oxygen. This means that the reaction begins with 27 moles of gaseous reactants. As the products are formed, the mole ratio indicates that 34 moles of gas is produced. As these gases will occupy a greater volume, and the cylinder of the engine is a fixed volume the pressure inside the cylinder will quickly increase and push the piston.
Also, the heat produced by combustion will increase the kinetic energy of all gas particles, further increasing the internal pressure of the engine.

11 **a** Energy will be absorbed by the evaporation of water. The energy released per gram will be less for coal with a high water content due to this endothermic change occurring during combustion. Black coal has a lower water content and will release more energy per gram.
b Some water can be removed by drying and/or crushing, which increases surface area for water to evaporate more easily.
c Potential atmospheric pollution from oxides of sulfur require either pre-treatment to remove or reduce sulfur from the brown coal and/or post-treatment to remove or reduce it from exhaust gases.

12 **a** The main component of coal seam gas is methane.
b The production of coal seam gas involves physically breaking apart the coal seam (fracking). An environmental concern associated with fracking is the potential contamination of ground water reserves of various materials used as fracking liquids or gases. Also, as the coal seam gases are released there is considerable risk of amounts of methane escaping into the atmosphere. This is a concern as methane is a powerful greenhouse gas.

13 **a** Calculations of gram and mole amounts of carbon and sulfur in 1 kg coal and the final conversion to a gas volume using the molar gas volume at STP (22.7 L) are summarised in the following table.

	Anthracite coal	Bituminous coal
fixed carbon, average weight %	83.1	61.55
sulfur, average weight %	0.685	0.55
fixed carbon, average mass in 1 kg	831 g	615.5 g
sulfur, average mass in 1 kg	6.85 g	5.5 g
mole carbon in 1 kg	69.19 mol	51.24 mol
mole sulfur in 1 kg	0.21 mol	0.17 mol
volume CO_2 (using molar gas volume, STP, 22.7 L)	1695.15 L	1255.38 L
volume SO_2 (using molar gas volume, STP, 22.7 L)	5.15 L	4.165 L

b When these two gases are released into the atmosphere they can combine with upper level moisture and form acidic compounds (carbonic acid and sulfuric acid). Acid rain is a result of high levels of these compounds in rain, and can be easily measured using pH. Plants and stone structures have shown adverse effects of acid rain.

11.3 Biofuels

11.3 Science as a Human Endeavour

Biogas generation and use at a piggery

1 Anaerobic respiration by bacteria turns the useless pig waste into useful biofuel and fertiliser.

2 Production of methane to be used as a fuel for the farm can reduce the need for using a fossil fuel to warm the farrowing sheds, and would also reduce carbon emissions in the form of methane that would normally escape from waste. The pig waste is not released into the environment as sewage but is converted into useful product.

3 Methane is a much stronger greenhouse gas than carbon dioxide. It is 21 times more effective at absorbing and redirecting infrared radiation. So, by converting methane to carbon dioxide (below), the overall greenhouse effect due to piggery waste gases is reduced.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

4 Hydrogen sulfide gas is a component of biogas and has a rotten egg smell. By collecting the gas in a covered pond, nearby residents would be spared this discomfort.

11.3 Key questions

Retrieval

1 A non-renewable fuel cannot be replenished at the rate at which it is consumed. Renewable fuels are those that can be replenished at a rate at which they are consumed.

2 transesterification

3 **a** true
b true
c false
d true
e False. (The representative structure on page 309 is saturated; however, biodiesel is a complex mixture of compounds.)
f true

4 Methane is the most abundant component of biogas, followed by carbon dioxide and then a mix of other gases in low percentages.

5 soybean, canola and palm oil

Comprehension

6 Fossil fuels are not renewable because they are being used at a rate faster than they can be replaced. This means that they will be exhausted, and the energy needs of our society will not be met. In finding renewable alternative fuels, we will be able to maintain our lifestyle and minimise the negative effects of excess carbon dioxide in the atmosphere.

7 renewable: bioethanol, biogas, biodiesel
non-renewable: coal, oil, LPG, natural gas, coal seam gas

8 renewable source—biogas
non-renewable source—natural gas

9 Biogas is a product of anaerobic bacterial digestion of various wastes. The type of waste, digestion period and growing conditions will alter the composition of gases.

Analysis

10 Coal has the longest expected lifespan based on estimated reserves and current rates of use. Concern for carbon dioxide emission issues is expected to reduce long-term dependence on coal.

11 CO_2 is absorbed from the atmosphere and used in photosynthesis and the production of the triglycerides that biodiesel is made from. These fuels return carbon compounds to the air upon combustion. However, CO_2 produced from the combustion of petrodiesel is derived from fossil fuel deposits underground. This means that when fossil fuels are burnt more carbon compounds are released into the atmosphere than were removed from the atmosphere.

12 Bioethanol and biodiesel will be available for the longest time—they are renewable fuels.

13 Photosynthesis uses six molecules of carbon dioxide to produce one molecule of glucose:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

Fermentation of glucose produces ethanol and two molecules of CO_2:

$$C_6H_{12}O_6(aq) \rightarrow 2CO_2(g) + 2CH_3CH_2OH(aq)$$

Combustion of ethanol releases heat and four molecules of CO_2:

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 3H_2O(g) + 2CO_2(g)$$

Therefore, all of the carbon dioxide released into the atmosphere by the fermentation and combustion of ethanol has been taken in by plants in the process of photosynthesis.
Some energy would be used during the production and transport of ethanol. This expended energy means that the fuel is not entirely carbon neutral.

14 advantages = a, b, c, e, f, g; disadvantages = d, h

15 The rate of production of non-renewable energy (such as fossil fuels) is the rate at which degradation of the biomolecules occurs. This process occurs as a result of heat and pressure underground over millions or hundreds of million years. Progressively, water is lost and coal, oil and gas are formed. This is a very slow process. Conversely, the rate of use of fossil fuels by society is rapid, 4×10^{20} joules per year globally. Non-renewable energy sources provide electricity and are used for transport and many other uses. This use is increasing and so, due to the extremely slow rate of fossil fuel formation, the existing supplies will be exhausted.

11.4 Comparing fuels

TY 11.4.1 energy released = 520 MJ

TY 11.4.2 1.27×10^2 L

11.4 Key questions

Retrieval

1 diesel, petrol, ethanol

2 The term ‘energy efficiency’ is used to describe the percentage of energy from a source that is converted to the form of energy desired.

Comprehension

3 Energy density provides the energy of a fuel per litre. This is a more useful measure when comparing fuels in everyday use because the volume occupied by the fuel is an important practical consideration for portable energy sources.

4 1006.57 MJ

5 4.19 kg

6 2.9 L

7 **a** advantage
b disadvantage
c advantage
d advantage
e disadvantage

8 propane: heat of combustion = -50.3 kJ g^{-1}
octane: heat of combustion = -47.7 kJ g^{-1}
Propane releases more energy per gram; therefore, propane releases more energy per kilogram.

9 **a** From the graph, the energy density of vegetable oil is approximately 34 MJ mol^{-1} so the energy released is 850 MJ.
b From the graph, the energy density of ethanol is approximately 30 MJ kg^{-1} so the energy released is 12 090 MJ.
c From the graph, the energy density of brown coal (lignite) is approximately 15 MJ kg^{-1} so the energy released is 75 000 MJ.

10 161 L

11 44 MJ kg^{-1} of energy is provided by octane but only 26.8 MJ kg^{-1} is provided by ethanol. This means that ethanol will use approximately $\frac{44}{26.8}$ per cent of the fuel that octane uses (164.2%). This is a 64.2% increase in weight needed to be carried.

Analysis

12 **a** For CH_4: 1.80 kg
For C_3H_8: 1.99 kg
b For CH_4: 4.95 kg
For C_3H_8: 5.95 kg
c For CH_4: 2.02×10^4 MJ tonne^{-1}
For C_3H_8: 1.68×10^4 MJ tonne^{-1}
d A lower mass of methane is required to produce 100 MJ of energy, so methane has the higher energy content per kg.
A lower mass of carbon dioxide is released in generating this amount of energy. This means methane has a higher amount of heat energy produced per tonne of carbon dioxide released.

13 **a** In a home, the fuel should have good energy density and produce few particulates (burn cleanly). It should not produce carbon monoxide.
b In a vehicle, the fuel should have high energy density because it needs to be carried, and it should be available from refilling stations (such as petrol stations). It should release low levels of sulfur and nitrogen oxides.

14 40.92 g

11.5 Environmental impact of the use of fuels

11.5 Key questions

Retrieval

1 infrared radiation

2 carbon dioxide, methane and water are greenhouse gases, and oxygen and nitrogen are not greenhouse gases

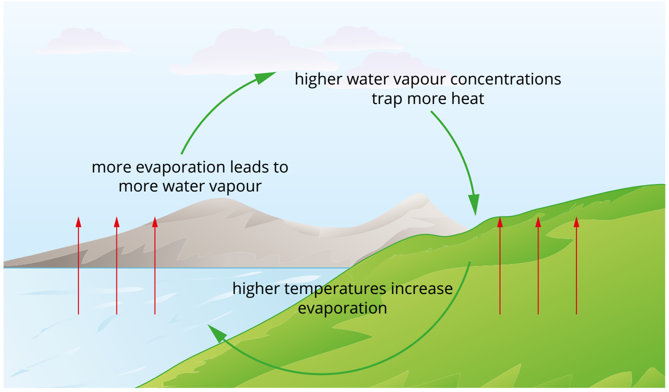
Comprehension

3 The energy content of bioethanol is less than that of octane, and so a greater amount of fuel is required to travel a given distance.

4 Individual student response required

5 **a** Our current method of burning coal to generate electricity produces greenhouse gas emissions, which does not help Australia meet its Paris Agreement obligations. However, the cost of electricity could rise as other energy sources are required to replace coal, which is relatively abundant and cheap. These are still being developed. Also, Australia is a major exporter of coal and so a worldwide decrease in its use would negatively affect our economy.
b Biofuels could potentially be carbon neutral—the carbon dioxide produced from the combustion of the fuels is offset by the carbon dioxide absorbed to grow the crops which are used to make the fuels. This would help Australia meet its Paris Agreement obligations regarding reduction of carbon emissions.

6



Chapter 11 Review

Retrieval

1 D

2 C

3 D

4 $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

5 **a** 2.2 kJ

b 774 J

c 4.42 MJ

d 56 kJ

6 brown coal, ethanol, petrol, diesel

Comprehension

7 41%

8 **a** E10 petrol is a mixture of 10% bioethanol and 90% conventional petrol.

b Use of E10 petrol will extend the availability of petrol as a fuel because bioethanol is not sourced from fossil fuels. Using this amount of E10 petrol will allow more of the C_6–C_{10} fraction from crude oil to be used as a feedstock for other uses rather than being burnt as a fuel.

9 **a** Biofuels (also called biochemical fuels) are derived from renewable resources such as plants. Ethene, the reactant described here as feedstock for the industrial production of ethanol, is derived from the fractional distillation of crude oil, a non-renewable resource. This means that this ethanol is a fossil fuel not, a biofuel.

b Bioethanol is produced by the fermentation of sugar by yeasts as shown here:

$$C_6H_{12}O_6(aq) \rightarrow 2CO_2(g) + 2CH_3CH_2OH(aq)$$

Sugar may be derived from sugar cane or by the hydrolysis of starch from grains such as maize. This means that a wide range of plant-based sugars and starches are suitable for bioethanol production. Newer technologies are developing the use of waste (non-food) cellulosic materials as feedstock. This means that food sources are not compromised by the production of fuels.

10 **a** 496.4 kJ

b 12.27 kJ

c 7.85 mL

11 $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \qquad \Delta H = -3120\,kJ$

$2C_2H_6(g) + 5O_2(g) \rightarrow 4CO(g) + 6H_2O(l) \qquad \Delta H = -1989\,kJ$

Using ethanol as an example, when the ratio of fuel : oxygen is 2 : 7, there is sufficient oxygen for each carbon atom to react with two oxygen atoms to form CO_2. When the ratio of fuel: oxygen is 2 : 5 there is insufficient oxygen present and only one oxygen atom reacts with each carbon atom, forming CO.

There is a concomitant decrease in heat of combustion when there is insufficient oxygen present because there is one fewer high energy C=O bond formed in the products.

Analysis

12 **a** 1976–2008 (approximately)

b The population of Australia in 1976 was 14 million and in 2008 it was 21.3 million. This is an increase of 50%. In comparison, energy consumption doubled (an increase of 100%). This data shows that the rate of increase of energy use is greater than the rate of increase in population. So, people in Australia must have increased their per capita (per person) energy use. One reason would be the vast number and decreasing prices of electronic devices. Also, a rise in living standards means that people, in general, have a greater ability to afford appliances for their homes and businesses.

13 Activities vary but may include:

- food storage and preparation
- plastic items including clothing made from polyester blends
- heating/cooling of home and school/work spaces
- electricity for appliances including hot water
- transport via car or public transport
- and many others.

14 Petrol is a mixture of compounds because fractional distillation of crude oil separates and collects fractions at temperature intervals. Typically, the petrol fraction is collected at temperatures between 110°C and 150°C (Figure 11.2.3). This means that different hydrocarbons that have a boiling point in this range will be collected. These include paraffins, naphthanes, aromatics and asphaltics. Naphthanes are the predominant hydrocarbon type in petrol. The data also shows that the possible range of % composition is high. For example, the amount of paraffins by weight in petrol can be as low as 15% or as high as 60%. The reason for this range is that petrol, like all fossil fuels, is the residue of a biological ecosystem. The type of biomass and the age of the crude oil will determine the hydrocarbons it contains. Crude oil obtained from Western Australia will have formed differently from, for example, that formed in Saudi Arabia.

15 **a** The methane experiment releases the most energy because it produces 890 kJ per mol, and the methanol experiment produces 725 kJ per mole.

b The methanol experiment will produce the most carbon dioxide for 2000 kJ energy because this fuel produces less energy per mole. This means that to produce the same amount of energy more methanol will need to be burnt.

16 **a** 700 g

b 616 g

Knowledge utilisation

17 Answers vary. Impacts of lack of fossil fuels on our lifestyle would be immense. The following would be not viable:

- devices, machines relying on power
- computers, lights, hospitals, work places, train transport
- cars, trucks and industries using petrol and diesel
- agricultural machinery
- materials made of plastics
- heating, cooling of buildings
- food processing and storage
- pharmaceuticals and medicinal organic chemicals
- many more.

18 Summary of ideas for answers

	Uses prime land needed for food	Removal of native ecosystems	Oil yield	Technological understanding
algae	no	no	very high	incomplete
palm oil	no	yes*	high	good
canola	yes	yes*	moderate	good
sunflower	no/yes	yes*	moderate	good
soy	no/yes	yes*	low	good

*depends on agricultural strategy

19 Answers vary.
Electricity can be conserved at home in the following ways.

- Use efficient appliances at home (for example, using new technology).
- Manage temperature of living areas reasonably (for example, setting the thermostat to moderate heating/cooling, and switch off when not in use).
- Use clothing to maintain comfort, not heating/cooling systems.
- Insulate the home (roof, walls).
- Keep doors and/or curtains closed for insulation.
- Wash clothes in cold water.
- Limit hot water use (kitchen, bathroom).

Chapter 12 Intermolecular forces

12.1 Shapes of molecules

TY 12.1.1 bent

12.1 Science as a Human Endeavour

Development of VSEPR theory

1 The earlier models could not accurately predict the shapes of all molecules, specifically those that contain lone pairs of electrons on a central atom (such as ammonia and water). Because VSEPR theory takes into account these lone pairs of electrons as well as the bonding pairs, its predictions about the shapes of molecules are more accurate.

2 Nyholm and Gillespie included rules about the types of electron repulsions and their relative effects. These rules accounted for the observed deviations from expected bond angles in many molecules.

3 The central oxygen atom in each water molecule has two bonding pairs of electrons and two lone pairs. These four pairs are arranged in a tetrahedral shape. However, the lone pairs of electrons occupy more space than the bonding pairs and, as a result, have a greater repulsive effect. The single O–H covalent bonds are therefore pushed closer together and the H–O–H bond angle is reduced from 109° to 104.5°.

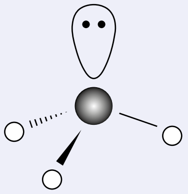
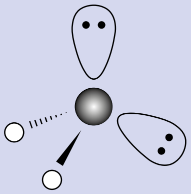
12.1 Key questions

Retrieval

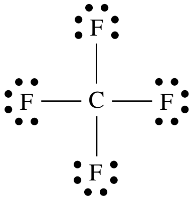
1 **a** VSEPR stands for Valence Shell Electron Pair Repulsion.
b A lone pair is a pair of electrons that is not involved in bonding.

Comprehension

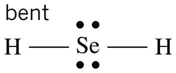
2

Molecule	Bonding groups	Lone pairs	Electron arrangement	Molecular shape
a	3	1		pyramidal
b	2	2		bent
c	2	0		linear

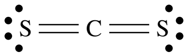
3 **a** tetrahedral



b bent

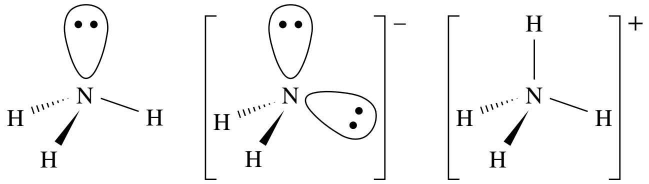


c linear



Analysis

4 NH_3, NH_2^- and NH_4^+ all contain four groups of electrons surrounding the central atom; therefore, each species has a tetrahedral electron arrangement.



NH_3 NH_2^- NH_4^+

However, there are a different number of lone pairs surrounding the central atom in each molecule, which influences the bonding angles. NH_4^+ has four bonding groups of electrons and no lone pairs; the shape of the molecule is tetrahedral with bond angles of 109°. NH_3 has three bonding groups of electrons and one lone pair; the shape of the molecule is pyramidal and the covalent bonds are pushed closer together (compared to those in NH_4^+) due to the increased repulsion from the lone pair of electrons. The bond angles are 107°. NH_2^- has two bonding groups of electrons and two lone pairs; the shape of the molecule is bent and the covalent bonds are pushed even closer together due to the presence of the extra lone pair of electrons. The bond angles are 104.5°.

12.2 Polarity of molecules

TY 12.2.1 HCl is more polar than NO.

12.2 Key questions

Retrieval

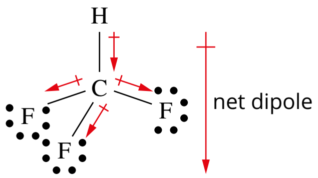
1 **a** Non-polar bond: a bond with an even distribution of valence electrons. Non-polar bonds occur between atoms of the same element.
b Polar bond: a bond with an uneven distribution of valence electrons. Polar bonds occur between atoms of different elements.

2 H = 2.20, Cl = 3.2. Therefore, HCl = 3.2 – 2.20 = 1.0
N = 3.0. Therefore, N2 = 3.0 – 3.0 = 0
H = 2.20, F = 4.0. Therefore, HF = 4.0 – 2.2 = 1.8
C = 2.6, O = 3.4. Therefore, CO = 3.4 – 2.6 = 0.8
In order of increasing polarity: N2 < CO < HCl < HF

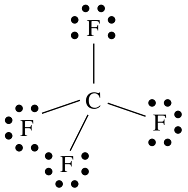
Comprehension

3 **a** non-polar
b polar
c polar
d polar
e non-polar

4 **a**

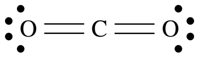


asymmetrical
polar

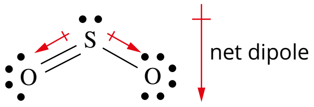


symmetrical
non-polar

b

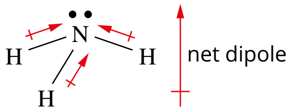


symmetrical
non-polar

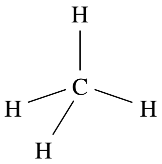


asymmetrical
polar

c



asymmetrical
polar



symmetrical
non-polar

12.3 Intermolecular forces

12.3 Key questions

Retrieval

1 **a** Dipole–dipole forces and dispersion forces exist between polar molecules.

b Only dispersion forces exist between non-polar molecules.

Comprehension

2 hydrogen chloride (HCl) and chloromethane (CH_3Cl)

3 a, g, h: ammonia (NH_3), hydrogen fluoride (HF) and water (H_2O)

Analysis

4 **a** F_2 and I_2 are both non-polar and only form dispersion forces between their molecules. The strength of the dispersion forces increases as the molecular mass increases. Because I_2 has a greater molecular mass than F_2, it would contain the strongest dispersion forces in the liquid state.

b CH_3Cl is polar and forms dipole–dipole forces between its molecules in addition to dispersion forces. CH_3OH is also polar but because it has an electronegative O atom attached to a hydrogen atom, it can form hydrogen bonds between its molecules. Because hydrogen bonding is the strongest type of intermolecular force, CH_3OH contains stronger intermolecular forces than CH_3F in the liquid state.

c F_2 and HCl have similar molecular masses so the strength of their dispersion forces would also be similar. However, F_2 is a non-polar molecule and only forms dispersion forces with neighbouring molecules, whereas HCl is a polar molecule so also forms dipole–dipole forces. These dipole–dipole forces outweigh the dispersion forces and result in HCl having the strongest intermolecular forces in the liquid state.

5 The difference in boiling points is due to the different shapes of the two molecules; pentane is a long molecule, while neopentane is compact.

Being less compact and long means pentane has more contact area to interact with its neighbouring molecules and form stronger dispersion forces, thus its boiling point is higher than neopentane's.

12.4 Properties of covalent molecular substances

TY 12.4.1 CS_2, because it has the weakest dispersion forces

TY 12.4.2 CH_3CHO contains the strongest intermolecular forces and would therefore have the highest melting and boiling points.

TY 12.4.3 The dispersion forces will be stronger in S_8 than O_2 because S_8 has the largest molecular mass. S_8 is more polarisable and will have the highest melting and boiling points.

TY 12.4.4 Iodine will have the highest solubility in hexane.

12.4 Key questions

Retrieval

1 **a** The normal boiling point is the boiling point of a liquid at sea level, where there is an atmospheric pressure of 101.325 kPa (or 1 atm).

b 100°C

2 Identify what attractive forces are broken when the following substances are melted or boiled.

- **a** covalent molecular substances
- **b** covalent network solids
- **c** ionic substances

- **a** intermolecular forces of attraction
- **b** intramolecular covalent bonds
- **c** intramolecular ionic bonds

3 The saying 'like dissolves like' refers to the fact that a solute will dissolve in a solvent if it contains similar intermolecular forces of attraction.

Comprehension

4 **a** Liquids with strong intermolecular forces of attraction have their molecules held relatively tightly, which makes it hard for them to escape from the surface of the liquid and move into the gas phase. Therefore, there will be few gas particles above the surface of the liquid and the vapour pressure will be low.

b As temperature increases, the liquid molecules gain more kinetic energy, which allows them to overcome their intermolecular forces of attraction more easily and move into the gas phase. Therefore, at higher temperatures there will be more gas particles above the surface of the liquid and the vapour pressure will be higher.

5 **a** The molecules in a covalent molecular substance are held together in the liquid state with intermolecular forces. When the liquid boils, these intermolecular forces are broken and the molecules are able to escape the liquid and move into the gas phase.

b Covalent molecular substances are held together in the solid and liquid states by intermolecular forces. Intermolecular forces are much weaker than strong intramolecular bonds such as covalent and ionic bonds and are therefore relatively easily broken. This results in covalent molecular substances having low melting and boiling points. Covalent network substances and ionic compounds are held together in the solid and liquid states with strong intramolecular covalent or ionic bonds; vast amounts of energy are required to break these bonds, which results in the substances having very high melting and boiling points.

6 non-polar solvent: methane and bromine
polar solvent: ethanol and propanone

Analysis

7 Because pentane and octane are both non-polar substances their liquids only contain dispersion forces. Pentane has a smaller molecular mass than octane (and is therefore a smaller molecule with less electrons); this results in pentane having the weaker dispersion forces of the pair and therefore the higher vapour pressure. (Remember the molecules in the liquid with the weaker intermolecular forces are held together less tightly compared to a liquid with stronger intermolecular forces; this makes it easier for the molecules to escape from the surface of the liquid and move into the gas phase. There will be more molecules in the gas phase and therefore a greater vapour pressure.)

8 Bromine and iodine monochloride have a similar molecular mass, which means the strength of their dispersion forces will also be similar. Bromine is non-polar so only contains dispersion forces. However, iodine monochloride is polar so also contains dipole–dipole forces. Consequently, bromine has the weakest intermolecular forces of the pair. Its molecules are less tightly held in the liquid and more able to move into the gaseous phase. Therefore, bromine has the higher vapour pressure.

9 **a** F_2 and HCl have a similar molecular mass so the magnitude of their dispersion forces must also be similar. However, HCl is polar and contains dipole–dipole forces in addition to dispersion forces, whereas F_2 is non-polar and only contains dispersion forces. Dipole–dipole forces are the stronger intermolecular forces and this results in HCl having a higher boiling point than F_2.

b PH_3 has stronger dispersion forces than NH_3 because it is a heavier molecule. However, because NH_3 is polar and contains an electronegative nitrogen atom attached to a hydrogen atom, it can form hydrogen bonds. This hydrogen bonding outweighs the effects of the larger dispersion forces in PH_3 and results in NH_3 having the higher boiling point.

c CH_4 and CCl_4 are both non-polar and contain only dispersion forces. Because CCl_4 is a much heavier molecule than CH_4, it is more polarisable and forms stronger dispersion forces. Consequently, CCl_4 has a higher boiling point than CH_4.

10 **a** propane < methoxymethane < methanoic acid

b Propane, methoxymethane and methanoic acid have a similar molecular mass so the strength of their dispersion forces is also similar. Therefore, the difference in boiling points is due to the different intermolecular forces that are present in each substance. Propane is non-polar so only contains dispersion forces; it has the lowest boiling point. Methoxymethane is polar and its dipole–dipole forces cause it to have a higher boiling point than propane. Methanoic acid is polar and can form hydrogen bonds between its molecules. As a result, methanoic acid has the highest boiling point.

11 Carbon tetrachloride is non-polar and only contains dispersion forces. Therefore, it will dissolve in other non-polar substances that have similar intermolecular forces. Pentane also only contains dispersion forces because it is non-polar so it will readily dissolve in carbon tetrachloride.

However, ethanol is polar and contains hydrogen bonds; because the intermolecular forces are so different, ethanol is insoluble in carbon tetrachloride. (If ethanol was to dissolve in carbon tetrachloride then it could only form dispersion forces with the carbon tetrachloride molecules. The energy that would be released when these ethanol–carbon tetrachloride dispersion forces form would not be enough to compensate for the large amount of energy required to break the hydrogen bonds between the ethanol molecules in the first place. Therefore, the ethanol does not dissolve.)

12 **a** ethanoic acid

b Bromine and carbon tetrachloride are non-polar molecules and only contain dispersion forces, whereas water contains hydrogen bonds. The different types of intermolecular forces prevent both bromine and carbon tetrachloride from dissolving in water. (If they were to dissolve, they could only form dispersion forces with the water molecules. The energy that would be released when these bromine–water and carbon tetrachloride–water dispersion forces form would not be enough to compensate for the large amount of energy that would be required to break the hydrogen bonds between the water molecules in the first place.)

Ethanoic acid contains hydrogen bonds. Therefore, when dissolved in water, it can form similar hydrogen bonds with the water molecules, which makes the dissolving process energetically favourable.

13 **a** substance A

b substance C

c ethanol = substance B, water = substance A and diethyl ether = substance C.

The dominant intermolecular forces in diethyl ether are dipole–dipole forces (diethyl ether molecules are polar). Both ethanol and water contain hydrogen bonds. Dipole–dipole forces are weaker than hydrogen bonds so diethyl ether would have a higher vapour pressure than both ethanol and water.

d The hydrogen bonding in water is more effective as evidenced by the lower vapour pressure for water. This is because water contains two hydrogen atoms attached to the electronegative oxygen atom that can both hydrogen bond with neighbouring molecules. Conversely, ethanol only contains one hydrogen atom that can form hydrogen bonds with neighbouring molecules.

14 **a** butan-1-ol boiling point ~118°C, pentan-1-ol boiling point ~136°C

b Dispersion forces are responsible for the increase in boiling point as the carbon chain length increases.

c As the carbon chain length increases, the strength of the dispersion forces formed between the molecules also increases. This is because the contact area over which the dispersion forces can act with neighbouring molecules is greater in the molecules with the longer carbon chains; the dispersion forces become stronger and the boiling points of the substances increase.

Chapter 12 Review

Retrieval

1 B
2 D
3 B
4 B
5 **a** true
b false
c false
d true
6 dipole–dipole forces, hydrogen bonding and dispersion forces
7 Each molecule must contain a hydrogen atom bonded to an electronegative oxygen, nitrogen or fluorine atom.

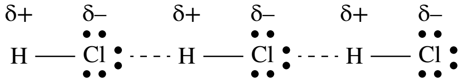
Comprehension

8 C
9 D
10 N–N: The atoms have identical electronegativities so the electronegativity difference is 0.
C–H: C = 2.6, H = 2.2. The electronegativity difference is 0.4.
H–F: H = 2.2, F = 4.0. The electronegativity difference is 1.8.
O–F: O = 3.4, F = 4.0. The electronegativity difference is 0.6.
F–F: The atoms have identical electronegativities so the electronegativity difference is 0.
The HF bond has the biggest difference in electronegativity; therefore, it is the most polar.

11 CO_2 contains C–O bonds. C = 2.6, O = 3.4. The electronegativity difference is 0.8.
H_2O contains O–H bonds. O = 3.4, H = 2.2. The electronegativity difference is 1.2.
H_2 is made up of a H–H bond. The atoms have identical electronegativities so the electronegativity difference is 0.
H_2S contains S–H bonds. S = 2.6, H = 2.2. The electronegativity difference is 0.4.
NH_3 contains N–H bonds. N = 3.0, H = 2.2. The electronegativity difference is 0.8.
The bonds in H_2O are the most polar.

12 **a** tetrahedral
b pyramidal
c tetrahedral
d bent
e pyramidal
13 b, c, d, e

14 Hydrogen chloride is polar and would therefore contain dipole–dipole forces between its molecules.



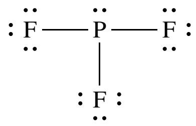
15 **a** hydrogen bonds
b dispersion forces
c Hydrogen bonds are overcome in both the ethanoic acid and the water (new hydrogen bonds form in the solution).

16 The molecules increase in molecular mass going from CH_3F through to CH_3I. As the molecules increase in molecular mass (and hence become larger with more electrons) they become more polarisable and can form stronger dispersion forces with neighbouring molecules. The stronger dispersion forces hold these heavier molecules together more tightly in the liquid state, which results in higher boiling points.

17 Diethyl ether is a polar molecule and can form dipole–dipole forces between neighbouring molecules in the liquid state. Ethanol contains an electronegative oxygen atom attached to a hydrogen atom so it can form hydrogen bonds with neighbouring molecules. The hydrogen bonding in ethanol is stronger than the dipole–dipole forces in diethyl ether; the intermolecular forces in ethanol are therefore harder to overcome when molecules are moving from the liquid to the gaseous phase. Consequently, there are less ethanol molecules in the gaseous phase and this results in a lower vapour pressure.

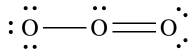
Analysis

18 **a** Lewis structure molecular shape



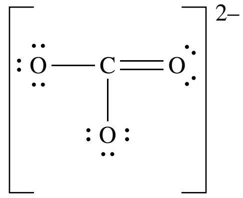
pyramidal

b Lewis diagram (1 resonance form shown) molecular shape



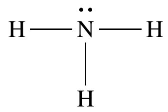
bent

c Lewis diagram (1 resonance form shown) molecular shape



trigonal planar

d Lewis diagram molecular shape



pyramidal

19 **a** 109° **b** 180° **c** 104.5° **d** 107°

20 **a** linear **b** bent **c** tetrahedral **d** pyramidal

21 Ozone only has one lone pair of electrons on the central oxygen atom, whereas water has two. The extra lone pair of electrons on the oxygen atom in the water molecule results in a greater repulsive effect with the bonding electrons so the O–H bonds are pushed closer together.

22 CBr_4 contains polar C–Br bonds. However, as shown below, the molecule has a perfect tetrahedral shape and is symmetrical. Therefore, the individual bond dipoles cancel out resulting in no net dipole. CBr_4 is therefore non-polar. NO_3^- is also a symmetrical molecule and has a trigonal planar shape, as shown below. Even though the N–O bonds are polar, the individual bond dipoles cancel out resulting in no net dipole.

23 Each of the molecules contain polar C–O bonds. However, only CO is a polar molecule. CO is an asymmetrical, linear molecule that contains an overall dipole. CO_2 is also linear but it is a symmetrical molecule. Therefore, its bond dipoles cancel each other out and results in the molecule having no net dipole and being non-polar. CO_3^{2-} is a trigonal planar molecule and it is also symmetrical; therefore, it has no net dipole and is also non-polar.

24 a, b, d, f

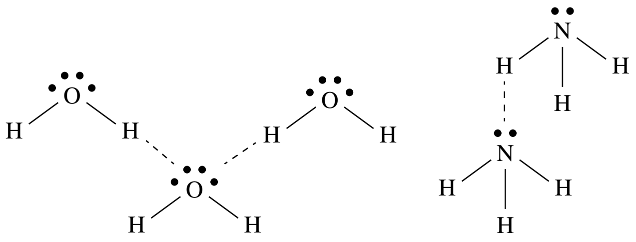
25 NH_3 and PH_3 are both asymmetrical and have a pyramidal shape, as shown below.
Of the two molecules, NH_3 has the most polar bonds and therefore the biggest difference in electron distribution within the molecule. NH_3 will be the most polar.

26 **a** Both compounds are non-polar and only contain dispersion forces. Because it is a liquid at room temperature, CCl_4 must have stronger intermolecular forces of attraction than $CH_{4.}$
b CH_4 and CCl_4 are both non-polar and only contain dispersion forces. The strength of the dispersion forces increases as the molecules increase in molecular mass. Remember that as molecular mass increases, the molecules get larger and become more polarisable. As a result, they can form stronger dispersion forces with neighbouring molecules. Because CCl_4 molecules are much heavier than CH_4 molecules, CCl_4 would contain the strongest dispersion forces and would require more energy to change state from a liquid to a gas.

27 $F_2 < HCl < HF$
F_2 and HCl contain molecules that are roughly of a similar molecular mass. However, F_2 is non-polar so only contains dispersion forces, whereas HCl is polar so also contains dipole–dipole forces. The dipole–dipole forces in HCl are stronger than the dispersion forces in F_2.
HF is also polar but because each HF molecule contains an electronegative fluorine atom attached to a hydrogen atom, HF exhibits hydrogen bonding. In general, hydrogen bonding is the strongest of the intermolecular forces; therefore, HF contains the strongest intermolecular forces of attraction.

28 Higher boiling temperature indicates that forces between molecules are stronger in CF_4. Molecules of CF_4 have a tetrahedral shape and are non-polar. Therefore, the only intermolecular forces that exist between them are dispersion forces. Molecules of OF_2 have a bent shape and are polar. Therefore, dipole-dipole forces as well as dispersion forces exist between them. However, because CF_4 has a larger molecular mass than OF_2 (88.01 amu compared to 54.00 amu) its dispersion forces are stronger and outweigh the dispersion forces and dipole–dipole forces in OF_2. Consequently, CF_4 has the larger boiling point.

29 In general, the hydrogen bonding is more efficient in water due to the availability of two lone pairs of electrons in each water molecule in comparison to one lone pair of electrons in each ammonia molecule. Therefore, in ammonia, some of the partial positive hydrogen atoms will not have access to a lone pair of electrons in a neighbouring molecule so the hydrogen bonding will not be as strong.



30 The melting and boiling points of a compound are closely related to the strength of the intermolecular forces.
Since fluorine and iodine are diatomic molecules of the same element, the electronegativity difference for both is 0. Therefore, the bonds are both non-polar and intermolecular forces must be due to differences in the strengths of the temporary dispersion forces. As you go down the group from fluorine to iodine the atomic

radius of the halogen increases (F: 60 pm; I: 136 pm). This increase in size means that the atom is more polarisable and so the ability to form temporary dipoles increases.

In fluorine, the electrons are tightly held to the nuclei, so the dispersion forces are relatively weak. Since the electrons in iodine are further from the nuclei (and so more polarisable), the dispersion forces are stronger. These stronger dispersion forces increase the melting and boiling points of iodine compared to fluorine.

31 NO_2. Both the CO_2 and NO_2 molecules have a similar molecular mass; this results in the strength of the dispersion forces within each substance also being similar. CO_2 is a non-polar molecule (linear and symmetrical), whereas NO_2 is polar (bent and asymmetrical). CO_2 only contains dispersion forces, whereas NO_2 also contains dipole–dipole forces. The dipole–dipole forces are stronger than the dispersion forces and cause NO_2 to have a higher boiling point than CO_2.

32 Both substances contain hydrogen bonds due to the presence of an electronegative nitrogen atom attached to a hydrogen atom within their molecules. Because hydrazine molecules have two sites where hydrogen bonding can occur, there will be stronger intermolecular forces of attraction in the liquid. The stronger intermolecular forces of attraction result in hydrazine having a higher boiling point than ammonia.

33 The vapour pressure of a substance increases as the strength of the intermolecular forces in the liquid phase decreases. CH_4 is a non-polar molecule and only forms dispersion forces with neighbouring molecules in the liquid. Not only is CH_2F_2 a heavier molecule (which results in stronger dispersion forces with neigbouring molecules) but it is also polar so forms dipole–dipole forces in addition to the dispersion forces. The stronger intermolecular forces in CH_2F_2 holds the molecules more tightly in the liquid so fewer molecules are able to move into the gaseous phase. CH_2F_2 will therefore have a lower vapour pressure than CH_4.

34 Other than NH_3, all of the substances contain non-polar molecules; as a result, they can only form dispersion forces with water molecules when dissolved; the energy released when this occurs is not enough to compensate for the large amount of energy required to break the hydrogen bonds between the water molecules in the first place. However, NH_3 can form hydrogen bonds with water, which makes the dissolving process much more energetically favourable. The energy that is required to break the intermolecular forces of attraction between the water molecules in the first place is compensated for by the release of energy that occurs when new NH_3–H_2O hydrogen bonds form in the solution.

35 **a** As the temperature increases, the water molecules gain more kinetic energy, which allows them to more easily overcome the intermolecular forces (hydrogen bonds) holding them together. More molecules escape from the surface of the liquid and move into the gas phase; this results in an increased vapour pressure.

b 100°C. The normal boiling point is the boiling point at sea level—101.3 kPa.

c about 72–74°C

d Boiling occurs when the vapour pressure of the liquid equals the atmospheric pressure. Because the atmospheric pressure on the summit of Everest is so much lower than at sea level, the vapour pressure of water becomes equal to the atmospheric pressure at a much lower temperature and results in a much lower boiling point.

36 methanol = 64.5°C
ethanol = 78.5°C
propan-1-ol = 97.2°C

As the carbon chain length increases, the dispersion forces also increase and results in progressively higher boiling points. Ethane-1,2-diol has the highest boiling point of 198°C. Molecules of ethane-1,2-diol contain two OH groups in comparison to the one OH group contained in the other three molecules. This extra OH group results in more hydrogen bonds being formed between molecules in the liquid state and results in the higher boiling point. The effect from these extra hydrogen bonds is quite large as evidenced by the boiling point of ethane-1,2-diol being much higher than the boiling points of the other substances.

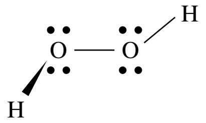
37 The melting points are too close together to use as evidence but the boiling points can be used. Methylamine has a higher boiling point than chloromethane and would therefore have stronger intermolecular forces of attraction in the liquid state.

Methylamine is soluble in water, whereas chloromethane is not soluble to any significant degree. Solutes dissolve in solvents that have similar intermolecular forces of attraction. Water contains strong hydrogen bonds that are difficult to break. Chloromethane is not very soluble in water because it contains weaker intermolecular forces of attraction than those in pure water. If chloromethane was to dissolve, the intermolecular forces of attraction that it would form with the water molecules would not be strong enough to compensate for the energy required to break the hydrogen bonds between the water molecules in the first place. On the other hand, methylamine is very soluble in water, which indicates that its intermolecular forces are of a similar strength and type to those in pure water and can more easily break the relatively strong hydrogen bonds between the water molecules. Therefore, it can be concluded that the intermolecular forces in methylamine must be stronger than those in chloromethane. This last piece of evidence also suggests that methylamine is capable of hydrogen bonding, which is the strongest type of intermolecular force; when methylamine dissolves in water, new hydrogen bonds are formed between the methylamine and water molecules.

Knowledge utilisation

38 **a** The molecule cannot be linear. If it was linear then the molecule would be symmetrical and non-polar (no net dipole).

b A possible structure is



39 Without taking into account the lone pairs, the predicted shape of the water molecule would be linear. If this was the true shape, water would be non-polar and would only form dispersion forces with neighbouring molecules rather than hydrogen bonds. Consequently, the observed properties of water would be drastically different from what they are known to be. For example, the melting and boiling points of water would be so low that it would be a gas at room temperature. Ice would sink rather than float in water. Water would not be able to dissolve any polar solutes such as salts and some biomolecules. There are many other examples that students could consider. Ultimately, if water was a linear molecule then life would not be able to exist on Earth.

40 Ethanol molecules are held together in the liquid state by hydrogen bonds. Water can easily dissolve ethanol because the energy required to break the hydrogen bonds between the water molecules in the pure solvent and the hydrogen bonds between the ethanol molecules in the pure solute is compensated for by the energy released when new ethanol–water hydrogen bonds are formed in the solution.

However, pentan-1-ol molecules are held together in the liquid state predominantly by dispersion forces. There are two main reasons for this.

1. In pentan-1-ol, the polar OH group is much smaller than the longer non-polar carbon chain and it is the longer non-polar carbon chain that has the greatest effect. As a result, the predominant intermolecular forces holding the pentan-1-ol molecules together in the liquid state are dispersion forces and not hydrogen bonds.
2. There is steric hindrance caused by the longer non-polar carbon chain in pentan-1-ol that prevents the OH groups in neighbouring molecules from being able to closely approach each other and form hydrogen bonds. Consequently, the dispersion forces become more important.

Because dispersion forces are the predominant intermolecular forces in pentan-1-ol, if pentan-1-ol was to dissolve in water then the new solute–solvent interactions would also predominantly be dispersion forces. The energy that would be released when the new pentan-1-ol–water dispersion forces form would not compensate for the energy required to break the hydrogen bonds between the water molecules in the first place. As a result, pentan-1-ol is much less soluble in water than ethanol.

41 When most liquids freeze, the intermolecular forces that form bring the molecules closer together and they become more tightly packed; the solid is therefore denser than its liquid and would sink if placed in the liquid. This is the case for benzene. Benzene is a non-polar molecule and only forms dispersion forces. Therefore, when benzene freezes, more dispersion forces form between the benzene molecules, drawing them closer together and packing them more tightly. The solid benzene is therefore denser than the liquid benzene and sinks. This is the case for most substances. However, water is an exception to the rule. When water freezes, the ice that is formed is less dense than the liquid water and the ice floats. This is because the hydrogen bonding holds the water molecules together in ice in a regular crystal lattice. In this lattice, the molecules are held further apart than in liquid water (Figure 12.3.10). As a result, ice is less dense and therefore, floats in liquid water.

42 In their summary, students should consider:

- the polarity of the bonds by determining the difference in the electronegativities of the atoms that make up the bonds; bonds that contain atoms with a large difference in electronegativity will be more polar
- the shape of the molecule and its symmetry
- molecules with non-polar bonds will be non-polar
- molecules with polar bonds that are symmetrical will also be non-polar
- molecules with polar bonds that are asymmetrical will be polar.

Chapter 13 Chromatography

13.1 Principles of chromatography

TY 13.1.1 0.7

13.1 Key questions

Retrieval

1

Term	Description
adsorption	the attraction of one substance to the surface of another
desorption	the breaking of the attraction between a substance and the surface to which the substance is adsorbed
components	the different compounds in the mixture, which can be separated by chromatography
polar molecule	a molecule that acts as a dipole; it has one or more polar covalent bonds, with the charge being distributed asymmetrically
mobile phase	the solvent that moves through the stationary phase in chromatography
stationary phase	the components of a mixture undergo adsorption to this phase

2 In paper chromatography, the paper acts as the *stationary* phase. A small spot of a solution is placed at one end of the paper, called the *origin*. The *sample* solution contains a number of different coloured compounds, the *components*.
The paper is suspended so that the end with the spot is *above* the surface of the solvent. The solvent or *mobile* phase moves up the *stationary* phase. Different coloured spots are observed at various places on the paper, due to the separation of different *components*.

Comprehension

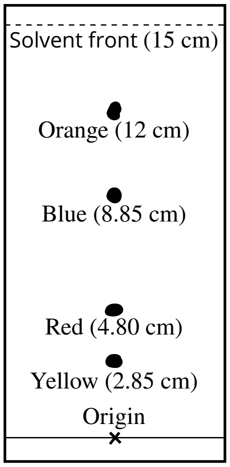
3
1 Dissolve a sample of pure phenacetin in a volume of chloroform. This is the standard solution.
2 Dissolve a tablet of the analgesic in chloroform. This is the sample solution.
3 Place a small spot of the sample solution near the bottom of a thin-layer plate. Place a spot of the standard solution next to it, at the same distance from the bottom of the plate.
4 When the spots are dry, place the plate in a container with a small volume of solvent, such as chloroform. The lower edge of the plate, but not the spots, should be immersed.
5 Allow the solvent to rise until it almost reaches the top of the plate and then remove the plate from the container.
6 Let the plate dry and examine it under ultraviolet light. If a spot from the sample appears at the same distance from the origin as the spot from the standard solution, the tablet is likely to contain phenacetin.

4 The component most strongly adsorbed to the stationary phase is *band A*.
As band B begins to emerge from the column, it appears to separate into two bands. These two components would be more successfully separated if you *increased the column length*.
GC and HPLC could provide clearer results because the particles in the stationary phase are *smaller*.

5 **a** **i** solvent **ii** paper
b **i** solvent
ii aluminium oxide, silica dioxide or other particulate substance packed into a column
c **i** solvent
ii layer of fine powder such as aluminium oxide on plate

6
1 The sample would dissolve in the solvent rather than be carried along by the stationary phase.
2 The mobile phase needs to 'move through' the sample so that components attracted to (soluble in) the mobile phase will be carried to a different place in the paper and R_f can be calculated.

7 **a** 3.2 cm
b 15 cm
c



8 The process **a** to **f** shows separation of the components in purple ink. At **a** the ink is applied to the column and solvent is continually added. **b** to **e** show that, as the sample moves through the column, the red dye in the ink is more soluble in the mobile phase and adsorbs less strongly to the stationary phase than the blue dye. The red dye in the ink moves faster through the column than blue dye, resulting in separate collection of the dyes. First, red is collected, then blue. Polar substances interact and therefore the slower moving substance must be more polar—the blue dye is more polar than the red dye.

Analysis

9

Technique	Problem
paper or thin-layer chromatography	B: difficult or impossible to obtain quantitative data D: samples must be able to be dissolved in solvent
high-performance chromatography (HPLC)	A: requires large amounts of solvents to operate C: expensive equipment needed D: samples must be able to be dissolved in solvent

10 Elution order is ethanol, propanone, ethanoic acid. In terms of polarity, 'like dissolves like'. In a chromatography column containing a non-polar stationary phase, the non-polar substances would be adsorbed more strongly to the stationary phase than non-polar substances. This means that the ethanoic acid would move more slowly than propanone and much more slowly than ethanol.

13.2 Gas chromatography

13.2 Key questions

Retrieval

1 The sample is vaporised by heating before entering the column and the mobile phase is a gas such as nitrogen.
2 a mass spectrometer

Comprehension

3 B, A, C
4 **a** E, C, D, B, A
b An alternative method to identify the herbicide is to obtain R_t values for both the water sample peaks and the herbicide peak, under the sample GC chromatography conditions. An identical R_t will identify the herbicide peak on the chromatogram of creek water.
5 Compared with conventional column chromatography, GC is more sensitive, faster, resolves components better and able to detect colourless components readily because it is usually combined with mass spectrometry.

Analysis

6 **a** Octane is probably peak 4 as it shows the highest peak.
b Using a non-polar stationary phase results in non-polar compounds having a greater elution time. This is because they are adsorbed to the stationary phase and their flow through the column is slowed. So, compounds 5, 6 and 7 must be more non-polar (less polar) than the component shown by peak 4. Conversely, compounds 1, 2 and 3 must be less non-polar (more polar) than the compound shown by peak 4.
7 Hydrocarbons are non-polar substances. As the mixture moves through the stationary phase the more polar molecules (B) have less interaction with the stationary phase than the more non-polar molecules (A). Due to non-polar interactions. A adsorbs to the surface of the stationary phase longer, whereas B is tending to desorb from the stationary phase more quickly (resulting in an earlier / shorter elution time) with the mobile phase. B has a lower retention time than A.

13.3 High-performance liquid chromatography

13.3 Key questions

Retrieval

1 high-performance liquid chromatography
2 The mobile phase and the stationary phase. Temperature will also affect R_t.
3 HPLC can test substances in aqueous solution (without vaporising) and can separate organic molecules with large molecular weights.
4 **a** retention time
b retention time
c peak area

Comprehension

5
1. A solid sample is dissolved in a suitable solvent.
2. The liquid sample is injected into the top of an HPLC column. The stationary and mobile phases are chosen to achieve a good separation of the components in the sample.
3. The sample components alternately adsorb onto the stationary phase and then desorb into the solvent as they are swept forward. The time taken to exit the column increases if the component strongly absorbs onto the stationary phase and has a low solubility in the mobile phase.
4. Detection upon exiting the column could be by absorbance, mass spectrometry or other assay. This provides data about the retention time (R_t) of various components, and an R_t curve.

Analysis

6 B, D, A, C, E
7 Serine, threonine, valine. Molecules of serine and threonine are more polar than valine. They will therefore dissolve more readily in the polar liquid used in this technique. Valine will move most slowly through the column. Because threonine has an extra methyl group in its side group, it will be less polar and move slower than serine.

13.4 Quantitative analysis using chromatography

TY 13.4.1 $0.75\,\text{mg}\,\text{kg}^{-1}$
TY 13.4.2 $3.0\,\mu\text{g}\,\text{mL}^{-1}$

13.4 Key questions

Retrieval

1 **1** Many contaminants of environmental, medical samples and food can have negative effects even at low concentrations. Acceptable levels need to be maintained and the exact level of the component of interest must be monitored.
2 A computer is normally associated with the mass spectrometer—it calculates the area under each peak to provide the concentration of each component. The software program can also link to online databases for compound identification.

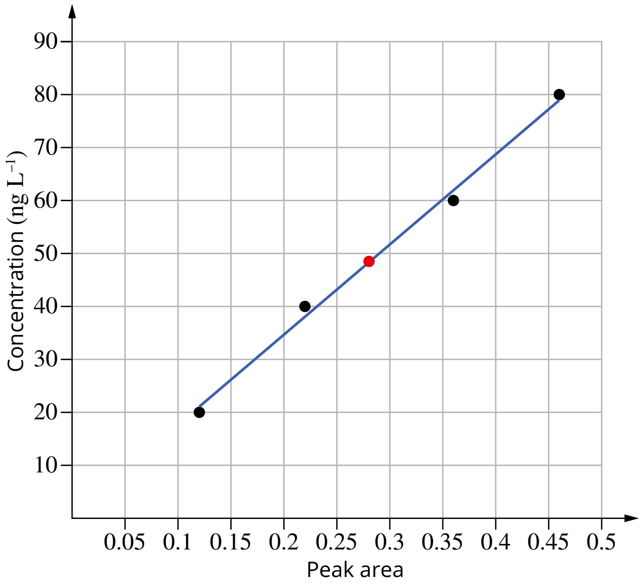
Comprehension

3 **a** 1 mL: mass of dioxin = $3.0 \times 10^{-8} \times 0.001 = 3.0 \times 10^{-11}$ mg
b 600 mL water bottle:
mass of dioxin = $3.0 \times 10^{-8} \times 0.600 = 1.8 \times 10^{-8}$ mg
c 750 ML reservoir:
mass of dioxin = $3.0 \times 10^{-8} \times 750\,000\,000 = 22.5$ mg
4 **a** Chemists can determine the number of components and the area of peaks of components, from which concentration can be determined. Relative polarity of compounds is also evident.
b 4
c The concentration of a component is determined by the area under a peak and is compared to a standard curve (calibration curve). The third peak appears to have the greatest area; however, there are two components eluting between 4 minutes and 5 minutes. So, it is estimated that the second peak (eluting at about 3 minutes) has the greatest area and therefore the highest concentration.

Analysis

5 Paper chromatography or thin-layer chromatography would be used for the initial investigation because it can quickly identify the components of a mixture, and a range of solvents can be used to assess the solubility of components of the mixture. HPLC and GC are more sensitive, capable of greater resolution and are able to provide a concentration for an unknown using the area under its peak on the trace. HPLC and GC, however, require a greater investment in time (to prepare and use the column) and money (cost of equipment).

6 48.3 ng L^{-1} (from graph)



Chapter 13 Review

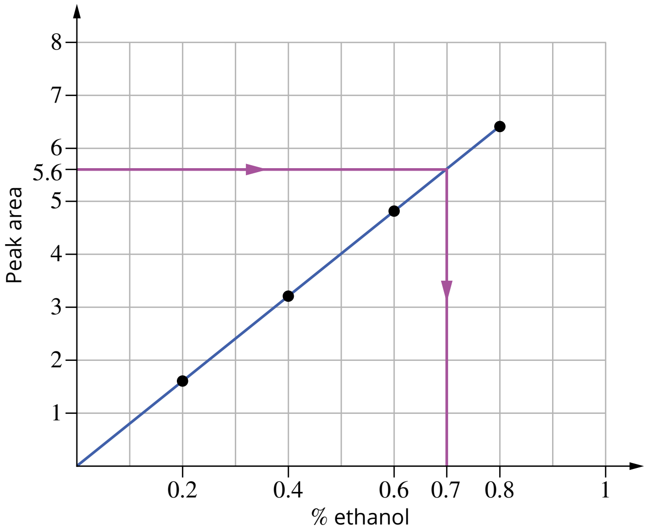
Retrieval

1 A
2 A
3 A
4 **a** Water was absorbed by the towel as the wet swimmer dried himself. A thin layer of grease adsorbed onto the cup when it was washed in the dirty water.
  b Absorb: Atoms or molecules are taken into the material. Adsorb: Atoms or molecules accumulate and bond weakly to the surface of a material.
5 chromatography—a set of techniques used to separate and analyse the components in a mixture
eluent—a liquid or gas used as the mobile phase in chromatography
mobile phase—the phase that moves through the stationary phase in chromatography
stationary phase—a solid, or a solid that is coated in a viscous liquid, used in chromatography
retention time—the time taken for a component to pass through a chromatography column
6 $$R_f = \frac{\text{distance the component travels from the origin}}{\text{distance the solvent front travels from the origin}}$$
7 R_f(blue) = 0.8
R_f(purple) = 0.6
R_f(yellow) = 0.2
8 R_f (blue dye) = 0.83
R_f (red dye) = 0.58
9 The stationary and mobile phases for GC are chosen to achieve a good separation of the components. When the sample is injected into the GC it becomes vapour and is propelled through the column by gas. Components alternately adsorb onto the stationary phase and then desorb into the solvent as they are swept forward and are detected (by mass spectrometry) as they exit. The time taken to exit the column increases if the component strongly absorbs onto the stationary phase and has a low solubility in the mobile phase.

Comprehension

10 B
11 D
12 During column chromatography, the components of the sample *adsorb* onto the stationary phase and *desorb* into the liquid mobile phase. A component that adsorbs most strongly to the *stationary* phase and is least soluble in the *mobile* phase would be expected to take the *longest* time to pass through the column.
13 **a** 2
  b purple
  c A and B
  d R_f(red spot) = 0.28; R_f(green spot) = 0.84
14 **a** If the solvent were above the level of the origin, the compounds under test would dissolve and disperse throughout the solvent.
  b Components in a mixture undergoing chromatography cannot move faster than the solvent that is carrying them over the stationary phase. R_f values must therefore be less than one.
  c 6
15 **a, b** 0.7%

Calibration curve for ethanol samples



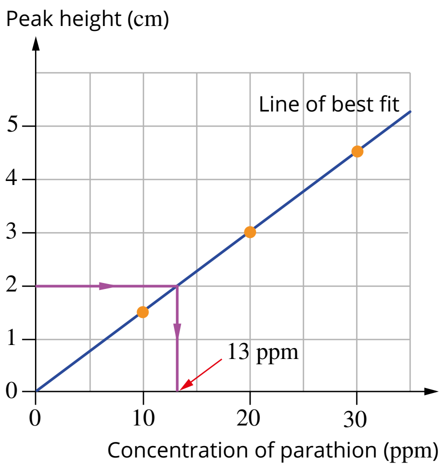
16 Solid samples are dissolved in a suitable solvent. The liquid sample is injected into the top of an HPLC column. The stationary and mobile phases are chosen to achieve a good separation of the components in the sample. The sample components alternately adsorb onto the stationary phase and then desorb into the solvent as they are swept forward. The time taken to exit the column increases if the component strongly absorbs onto the stationary phase and has a low solubility in the mobile phase.

Analysis

17 acetonitrile, methanol, ethanol, 1-propanol, 1-butanol, hexane
18 B
19 C
20 The component at the top of the chromatogram has a greater rate of desorption than adsorption, whereas the components at the bottom of the chromatogram have a greater rate of adsorption compared to desorption.
21 **a**

Standard	Peak height (cm)
Standard 0 ppm parathion	0.2
Standard 10 ppm parathion	1.5
Standard 20 ppm parathion	3.1
standard 30 ppm parathion	4.4
reservoir water	2.0

b



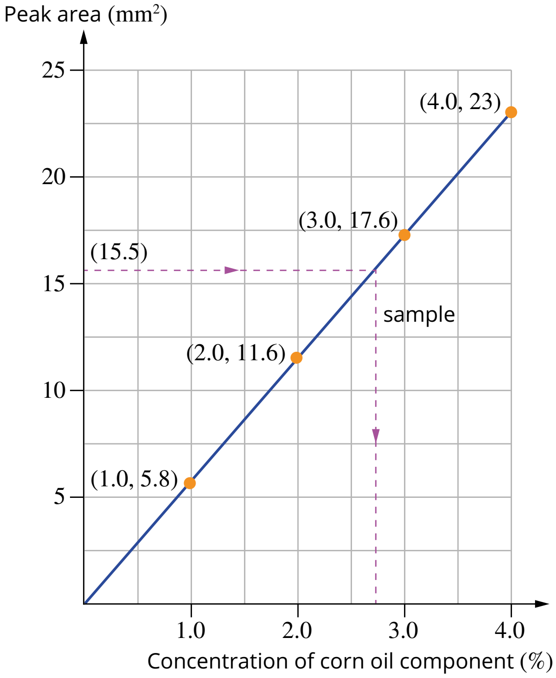
c concentration of parathion in water sample = 13 ppm.

d 13 ppm = 13 μg mL^{-1} = 13 mg L^{-1}. No, this is over 1000 times the allowable concentration.

e 92 mL

22 A, B, C, D

23 A line can be drawn on the calibration curve from the peak area of the sample (15.5 mm^2) to find the corresponding concentration on the graph: 2.7%.



24 For parts **a**, **b**, **c**, it would be useful to set your answers out in a table.

Band	a Distance from origin (mm)	b R_f	c Compound
light green	20	0.33	chlorophyll b
dark green	27	0.45	chlorophyll a
orange	40	0.67	xanthophyll
yellow	50	0.83	β-carotene
solvent front	60	–	–

d Using water as the solvent, the chromatogram would be different because separation of components depends on their solubility in the mobile phase (as well as strength of adsorption to the stationary phase). The polarity of the solvent used in TLC and paper chromatography will affect the R_f of the sample components. A polar solvent will dissolve polar samples readily; a non-polar solvent will dissolve non-polar samples readily.

25 **a** Taurine, glycine and an unknown. Two of the three spots produced by the medicine match those produced by taurine and glycine. The third spot does not match any of the amino acid standards and represents an unknown substance.

b There are several ways of visualising the spots:

- viewing the chromatogram under UV light
- spraying the finished chromatogram with a compound that causes the amino acids to fluoresce
- spraying the finished chromatogram with ninhydrin that causes a brown spot for organic compounds.

c Leucine. The component in sample A is bonded least strongly to the stationary phase as it has moved the greatest distance from the origin.

26 1 An origin (line) is drawn on the thin-chromatography plate and sample positions labelled (pencil).

2 Amino acid standards and the unknown are dotted onto the plate.

3 The plate is placed vertically in solvent so that the liquid is below the origin and solvent rises up the plate.

4 When the solvent front is near the end, the plate is removed from liquid and the front is marked.

5 The plate is sprayed with ninhydrin solution to visualise samples.

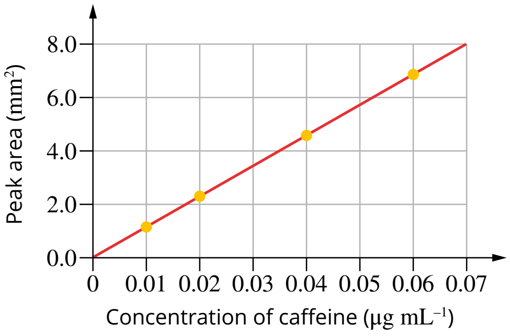
6 R_f for each component is calculated by distance

$$\frac{\text{distance travelled}}{\text{total distance}}.$$

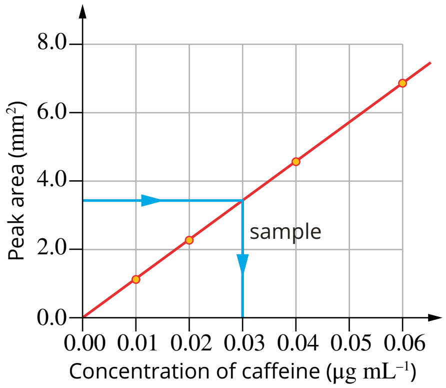
7 Components of the unknown are identified by having the same R_f as a standard.

27 The body fluid is first separated by HPLC to reveal its components (X). The peaks can be tentatively identified by running a sample containing pure insulin under exactly the same conditions as the sample (Y). A component can also be identified by adding a known compound to the sample (spiking). Figure Z shows the sample spiked with insulin. Insulin was added to the sample and the spiked chromatogram shows the second peak to be larger than it was in the unspiked sample. There are no extra peaks, indicating that insulin is very likely to have been present in the sample.

28 Values for standards are plotted.



A line (blue) can be drawn on the calibration curve from the peak area of the sample to find the corresponding concentration on the graph.



The concentration of caffeine in the energy drink is 0.030 μg mL^{-1}.

29 The equipment shows a gas chromatograph because the column is heated in an oven so samples are assayed in a gas state. Also, carrier gas is used compared to HPLC where the mobile phase is a liquid. The features in this diagram that are common to both HPLC and GC are:

- pressure regulator
- flow controller
- injector
- column
- detector
- recorder/data system/electrometer.

Knowledge utilisation

30 **a** A: 0.59; B: 0.49; C: 0.38; D: 0.20; E: 0.13

b A: leucine or isoleucine; B: *β*-phenylamine; C: proline or tyrosine; D: threonine, hydroxyproline or glycine; E: lysine, arginine or taurine

c A: leucine or isoleucine; B: *β*-phenylamine; C: proline; D: glutamic acid; E: arginine

d A two-way chromatogram produces better separation of components of complex mixtures, permitting easier isolation and identification.

31 The stationary phase could be powdered alumina (or powdered alumina on glass sheet). The mobile phase could be water or butanol. These solvents can be used combined in a mixture with ethanoic acid to provide a polar solvent. The mobile phase should be a polar solvent in order to dissolve the polar sample. The amino acids can be visualised after separation by spraying with a ninhydrin stain.

32 A thin-layer chromatography surface with non-polar characteristics could be used. Blood and urine are water-based (polar) and the polar compound of interest will move quickly up the test stick in solution. If a colour-changing substance that reacts with the compound of interest could be placed near the far end of the stick, the fact that the compound of interest is polar means that it will have separated from other materials present when it reaches the far end of the stick. This means that the test at the far end of the stick will be more accurate because the compound has been separated as it moves up the stick (chromatogram).

Chapter 14 Gases

14.1 Introducing gases

TY 14.1.1 7.0 × 10^{-4} m^3 **TY 14.1.2** 0.903 bar

14.1 Key questions

Retrieval

1 Four physical properties of gases:

- low density
- assume the shape and volume of their container
- compress easily
- mix together rapidly.

2 Elastic collisions occur between two gas particles or a gas particle and the side of its container where the kinetic energy is conserved.

3 Pressure is the force exerted by a substance per unit area of the surface to which the force is applied.

4 False. 1 bar is equal to 100 kPa and the SI unit is the Pascal.

Comprehension

5 **a** Molecules of gases are in constant, rapid, random motion and the forces between molecules are negligible. They continue to move outwards until stopped by the walls of the container, filling all the space available.

b Most of the volume occupied by a gas is space, so compression can be achieved by reducing the space between the particles.

c The molecules in a gas are spread much further apart than those of a liquid. A given mass of gas would occupy a much greater volume than the same mass of the liquid phase. Therefore, the density of the gas is less.

6 **a** During a long journey on a hot day, the air in a car tyre warms up. This means the air molecules have increased kinetic energy, and collisions with the walls of the tyres will increase in frequency and exert more force, and so the pressure will increase.

b Particles from the cooking food escape the pot and move randomly through the house. If the food has an odour, and if there are enough particles in the air, you will detect the odour as you enter the house.

c As air is pumped into a balloon, air molecules collide with the rubber of the balloon, forcing it to expand. If too much air is pumped in, the balloon reaches a stage where it cannot stretch any further. If the number of collisions by molecules per given surface area is increased still further, the rubber will break.

Analysis

7 Gases are much lower in density than liquids and are much more easily compressed. Both liquids and gases take the shape of their container, but gases also expand to take the same volume as their container. Both liquids and gases can mix, but liquids mix more slowly.

8 The gas in the inflated air bag can be compressed. As a result, the bag is able to absorb some of the energy of the collision.

9 **a** As temperature increases, the average kinetic energy of gas molecules in the can will increase. This will lead to an increase in the frequency and force of collisions of gas molecules with the inside walls of the aerosol cans. This will cause an increase in pressure.

b As the syringe is compressed, the inside surface area of the syringe will decrease. The number of collisions of molecules per unit area per second with the inside walls of the syringe will increase. This will cause a pressure increase.

10 The intermolecular forces between gas particles are extremely weak, and as gas particles have high kinetic energy their constant movement reduces the effect of the intermolecular forces between the particles, enabling them to move in an independent manner.

11 **a** 2 × 10^3 mL **b** 3.7 × 10^3 L **c** 2.85 × 10^{-4} m^3

12 **a** 1.40 × 10^5 Pa **b** 3.22 × 10^3 mmHg; 4.30 × 10^5 Pa

c 1.84 atm; 1.87 10 Pa; 1.87 bar

14.2 The ideal gas equation

TY 14.2.1 373 K **TY 14.2.2** 80 L **TY 14.2.3** 150 L

14.2 Science as a human endeavour

Decompression chambers

1 HBOT is a medical treatment in which patients breathe pure oxygen while inside a decompression chamber at a pressure higher than 1 atm.

2 If a diver using scuba equipment ascends too quickly, the rapid pressure drop causes the nitrogen to come out of the blood as tiny bubbles. These bubbles cause pain in joints and muscles. If they form in the spinal cord, brain or lungs, they can cause paralysis or death.

3 For gases, solubility decreases as temperature increases. Increased temperature means an increase in kinetic energy of the dissolved gas particles. This higher average kinetic energy means there are more gas particles with sufficient energy to break the intermolecular bonds and escape the solution. The fizz of soft drink comes from dissolved CO_2 leaving the solution as bubbles. Warm soft drink should go flat faster than cold soft drink.

14.2 Key questions

Retrieval

1 The volume of a gas is inversely proportional to its pressure.

2 The mathematical relationship between temperature, *T*, and volume, *V*, can be expressed as $\frac{V}{T} = k$, where *k* is a constant at a given pressure.

3 Molar volume is the amount of space, or volume, occupied by 1 mole of any gas at a particular pressure and temperature.

4 A hypothetical gas whose molecules occupy negligible space and which has no interactions; consequently, it obeys the gas laws exactly.

5 False. *R* is dependent of the units of pressure and volume used. It only has a value of $8.31\ \text{J K}^{-1}\ \text{mol}^{-1}$ if *P* is measured in kPa and volume is measured in L, as *PV* is work and measured in J.

Comprehension

6 Pressure is the force exerted on any surface, compared to the area of surface to which the force is applied. Gases exert pressure on any surface they come into contact with because the constant motion of the particles results in collisions with that surface. If the volume decreases, there will be less internal surface area. The number of gas particles and the force they exert will remain the same, but they will act over a smaller area. Thus, the pressure will increase.

7 Heating the air gives more kinetic energy to the air molecules, causing the balloon to expand, which reduces the density of the air inside, and allows the balloon to rise.

8 Strictly speaking the molar volume does vary slightly. However, except at very low temperatures and high pressures, most gases behave in an ideal way, so their molar volume is relatively similar?

Analysis

9 At higher altitudes, there is also a drop in pressure as well as temperature. In this instance, the pressure change has more of an effect on the volume of the balloon.

10 An ideal gas is one that follows the gas laws at all conditions. Gases that do not obey the gas laws are called real gases. Strictly speaking, ideal gases do not exist. However, most gases, except at low pressures and high temperatures, behave like ideal gases and obey the gas laws. At very low temperatures and very high pressures, gases are more likely to behave like real gases.

11 **a** $T = 100 + 273 = 373\ \text{K}$
b $T = 175 + 273 = 448\ \text{K}$
c $T = -145 + 273 = 128\ \text{K}$

12 **a** 31.4 L **b** 22 mL **c** 1.13 L

13 **a** 2.48 g **b** 70.4 g **c** 0.272 g

14 1.2×10^2 kPa

15 **a** 4.5 L **b** 86.7 L **c** 6.22 L

16 17 g

17 31°C

18 These conversions should be applied as needed:
1.00 atm = 101.325 kPa = 1.01325 bar.
$n(N_2) = 0.155$ mol

$n(\text{He}) = 0.123$ mol

14.3 Gas stoichiometry

TY 14.3.1 468 L

TY 14.3.2 2.24×10^3 L

TY 14.3.3 1 volume of CH_4 produces 1 volume of CO_2 gas, so 50 mL of CH_4 produces 50 mL of CO_2.

TY 14.3.4 **a** As there is 1.36 mol of C_4H_{10}, the limiting reagent is C_4H_{10}
b volume of CO_2 formed is 101.7 L

14.3 Key questions

Retrieval

1 Stoichiometry is the study of ratios of moles of substances.

2 Use $n = \frac{m}{M}$, $n = \frac{V}{V_m}$ or $n = \frac{PV}{RT}$.

Comprehension

3 First, determine the number of moles of the reactant using its mass. Next, use the mole ratio from the balanced chemical equation to determine the number of moles of the unknown. Finally, calculate the volume of the unknown using the ideal gas equation ($PV = nRT$) or the molar volume at STP (22.7 L).

4 It is possible, if the temperature and pressure are constant, because two gases will have the same volume if they are behaving in an ideal way.

Analysis

5 **a** **i** 56.64 L **ii** 33.98 L
b **i** 8.75×10^3 L **ii** 5.25×10^3 L

6 24.3 g **7** 4.07 L **8** 0.5 L

9 $V(\text{CO}) = 150$ mL $\therefore$ $V(O_2) = 75$ mL

Chapter 14 Review

Retrieval

1 D

2 A

3 The volume occupied by the atoms or molecules in a gas is much *smaller* than the total volume occupied by the gas. The particles move in rapid, *straight-line* paths and collide with each other and with the walls of the container. The forces between particles are extremely *weak*. The collisions between particles are *elastic*. The average kinetic energy of the particles is *directly* proportional to the temperature of the gas, in units of *K*.

4 $PV = nRT$

5 $R = 8.31\ \text{J K}^{-1}\ \text{mol}^{-1}$

6 the law of conservation of mass

Comprehension

7 **a** As volume is reduced, there is an increase in the frequency of molecular collisions per unit wall area. This is measured as an increase in pressure.
b When the temperature of a gas is lowered, the average kinetic energy of the particles decreases. The rate of collisions between particles and the walls of the container decreases and particles collide with less force. As pressure is a measure of the force of molecular collisions per unit wall area of the container, pressure is found to decrease.
c In a mixture of gases, the particles of each gas are moving and colliding with the walls of the container, independently of each other. Each gas therefore exerts a pressure. As the gases behave independently of each other, total pressure is simply the sum of the individual gas (or partial) pressures.
d When more gas is added to a container, the total number of particles in the container increases. Provided that the volume of the container and the temperature have not changed, the collisions with the container wall of these additional particles means that the total pressure in the container has increased.

8 The relationship between temperature and volume is directly proportional and linear when graphed. When the graph is extrapolated to a volume of 0 L, it crosses the temperature axis at –273°C.

9 At very low temperatures or very high pressures, the particle volume to container volume ratio becomes measurable. Under these conditions, when gas particles are much closer together than they would normally be, are intermolecular forces are more significant and gases can tend to behave in a non-ideal way.

Analysis

10 For a fixed amount of gas at a constant temperature, as the pressure decreases the volume increases. The relationship between the volume and pressure of a gas is inversely proportional.

11 This is because the gas is assumed to be behaving in as an ideal gas. If a gas is behaving in an ideal way, it is assumed the particles occupy no volume, irrespective of what type of gas. Also, a gas that obeys the gas laws is assumed to have elastic collisions between particles and the sides of the container. This means there are no intermolecular forces.

12 **a** 1.7×10^3 mL **b** 4.41×10^2 kPa **c** 5.02 bar
13 70.4 g **14** **a** 5.68 L **b** 20.29 g **15** 5.55 L
16 13.8 g **17** 497°C **18** **a** 0.22 mol **b** 46 g mol^{-1}
19 container A: 1.66 mol, container B: 0.00197 mol ∴ container A has more oxygen
20 **a** 8.94×10^3 mol **b** 257.7 kg
21 24.45 g **22** 35.0 L
23 **a** $V(CH_4)$ used = 5 L
b $V(O_2)$ used = 10 L
c 4.58 g

Knowledge utilisation

24 **a** There is excess oxygen, which means that all of the propane will react, and that there will be 100 mL of oxygen in excess.
b $V(CO_2)$ = 240 mL; $V(H_2O)$ = 320 mL
c There was an increase in volume of 80 mL.
25 $m(O_2)$ required per minute = 44.8 g
26 1.89×10^{12} L

Chapter 15 Water and aqueous solutions

15.1 Structure and properties of water

15.1 Key questions

Retrieval

1 Physical properties that are unusual to water include:
- relatively high melting and boiling temperatures for its molecular size
- decrease in density on freezing
- high heat capacity
- high latent heat of fusion and evaporation for a substance of its molecular size.

2 When liquid water is cooled, the molecules move more slowly so the molecules form a hexagonal structure. This creates a very open arrangement of molecules, meaning that the water molecules in ice are more widely spaced than in liquid water. Therefore, ice is less dense than liquid water. When ice melts, the water molecules move more freely and move closer together.

3 Molecules of water in the body of the liquid are surrounded in all directions. The sideways forces of attraction to neighbouring molecules are equal in all directions, so for molecules in the body of the liquid, there is no net force.
However, the attraction of surface water molecules to the molecules below them has no opposite force, so the water molecules at the surface are pulled downwards.

4 The difference in electronegativities of hydrogen and oxygen is large, meaning oxygen has a greater attraction for the electrons in the covalent bonds formed with hydrogen. This creates an overall negative dipole around the oxygen atom, producing a positive dipole around the hydrogens. Dipole–dipole interactions can occur between water and other polar molecules. Hydrogen bonds can form between hydrogen atoms on one molecule and atoms of oxygen, nitrogen and fluorine, thus water can form hydrogen bonds with itself.

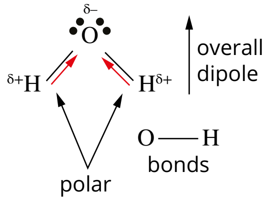
Comprehension

5 Each water molecule has two hydrogen atoms and one oxygen atom. The oxygen atom has two pairs of non-bonding electrons, each of which can form one hydrogen bond.
So the maximum number of water molecules with which one water molecule can form hydrogen bonds is four: up to two hydrogen bonds involving the two hydrogen atoms and up to two hydrogen bonds involving the two pairs of nonbonding electrons on the oxygen atom.

6 With the exception of water, the boiling points of the group 16 hydrides increase going down the group. This is due to the increasing strength of dispersion forces as the molecules increase in mass. Water has a significantly higher boiling point than the other compounds due to the hydrogen bonds between its molecules.

7 The bond between H and O atoms in water is highly polar. As a result, hydrogen bonds exist between water molecules.
Hydrogen bonds are stronger than other intermolecular bonds (although still weaker than the covalent intramolecular bonds) and so require more energy to break. Thus, water has relatively high melting and boiling temperatures.
Hydrogen bonding between water molecules in ice results in a very open arrangement of molecules, so ice is less dense than liquid water.

8



Analysis

9 The high surface tension of water means that some insects are able to move across the water without becoming submerged. If detergents leeched into the natural waterways these organisms would not be able to access food and move through the environment efficiently and safely.

15.2 Water as a solvent

TY 15.2.1 Ammonium phosphate is a water-soluble, phosphate-ion containing compound.

15.2 Science as a human endeavour

Blood chemistry

1 Being proteins, antibodies have a number of polar regions that enable them to interact with water and be soluble in the blood plasma. Due to the nature of different amino acids in the protein chain, some regions of the protein will interact with water by some forming hydrogen bonds and others through dipole–dipole interactions. In contrast, calcium, being an ion, will interact and dissolve in water by forming ion–dipole interactions.

2 Being lipid based, steroid hormones are less soluble in water so these hormones actually attach themselves to protein carriers. This increases their solubility in water, making them able to be transported around the body in the blood plasma. Lipids are the only non-polar macromolecule transported around the body so they need to make use of these protein carriers in the blood.

15.2 Key questions

Retrieval

1 **a** water—solvent **b** wine—solution
c ethanol—solute **d** sugar—solute

2 Hydroxyl functional groups contain a polar region due to the difference in electronegativities of hydrogen and oxygen. Water also contains these same polar regions. Hydroxyl functional groups are capable of forming hydrogen bonds between the electronegative oxygen atom on the water and the hydrogen atom of the hydroxyl group. Likewise, the hydrogen of the water molecule can form hydrogen bonds with the oxygen atom of the hydroxyl group.

3 Vitamin C contains many hydroxyl groups that make it readily dissolve in water in the blood and can be easily excreted in urine, therefore it needs to be consumed regularly to replace what is lost.

Comprehension

4 **a** Ammonia can form hydrogen bonds with water and is likely to be soluble.
  b O_2 is non-polar and is likely to be non-soluble.
  c HCl forms ion–dipole bonds, causing ionisation and hydration of hydronium ion and Cl ion.
  d Methanol forms hydrogen bonds with water, and is likely to dissolve.
  e Methane is, overall, non-polar and is likely to be non-soluble.
  f HF (same as part **c** above)
  g CO_2, although each C=O is polar, the overall molecule is non-polar and is likely to be non-soluble.

5 Dissolving methanol in water: $CH_3OH(l) \xrightarrow{H_2O(l)} CH_3OH(aq)$
Dissolving glucose in water: $C_6H_{12}O_6(s) \xrightarrow{H_2O(l)} C_6H_{12}O_6(aq)$

6
- Hydrogen bonds between water molecules are broken.
- Hydrogen bonds between ethanol molecules are broken.
- Hydrogen bonds between the water molecules and the ethanol molecules will form.

7 Dissolving hydrogen iodide in water:

$$HI(s) + H_2O(l) \rightarrow H_3O^+(aq) + I^-(aq)$$

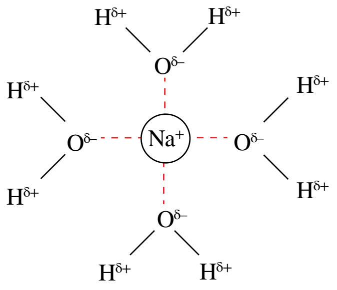
8 Dissolving sodium nitrate in water:

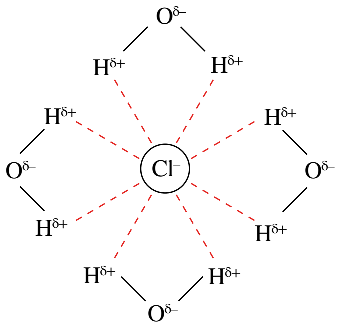
$$NaNO_3(s) \xrightarrow{H_2O(l)} Na^+(aq) + NO_3^-(aq)$$

Dissolving calcium hydroxide in water:

$$Ca(OH)_2(s) \xrightarrow{H_2O(l)} Ca^{2+}(aq) + 2OH^-(aq)$$

9





- - - - represents ion–dipole interaction

Analysis

10 **a** Nitrates are highly soluble in water. If found on Earth, they would dissolve in rainwater and wash into the oceans. Therefore, they are found only in areas of low rainfall.
  b The high solubility of sodium, chloride and sulfate ions results in them dissolving and flowing into the world's oceans.

11 A, B, D, E, H

12 The solute–solvent forces are stronger than the solute–solute and solvent–solvent forces.

13 **a** nitrate (NO_3^-) or ethanoate (CH_3COO^-)
  b lead (Pb^{2+})

14 As chain length increases, the region of the hydrocarbon molecule that is non-polar also increases. The only part of the molecule that is able to form hydrogen bonds and dipole–dipole interactions with water is the carboxylic acid end of the molecule. Therefore, as the length of the chain increases, its solubility in water also decreases.

15 Magnesium ions are cations; they have a positive charge. This means the negatively charged non-bonding electron pairs on the oxygen atoms in the water molecules are attracted to them. The water molecules arrange around the magnesium ion with their oxygen atoms, rather than the hydrogen atoms, closest to the magnesium ion. Chloride ions are anions; they have a negative charge. This means the hydrogen atoms in the water molecules, which have a partial positive charge, are attracted to them. The water molecules arrange around the chloride ion with their hydrogen atoms, rather than the oxygen atoms, closest to the chloride ion.

15.3 Solubility

TY 15.3.1 140 g **TY 15.3.2** 210 g **TY 15.3.3** 10 g

15.3 Key questions

Retrieval

1 **a** saturated **b** unsaturated **c** supersaturated

2 saturated

3 When solutions are cooled slowly the crystals that are formed as the solvent evaporates are larger than those where the solvent evaporates quickly.

4
- Each point on a curve represents the maximum amount of the solute that can be dissolved in 100 g of water at a particular temperature. Therefore, each point on the curve represents a saturated solution.
- Any point below a curve represents an unsaturated solution.
- Any point above a curve represents a supersaturated solution.
- For most solids, as temperature increases, the solubility increases.

Comprehension

5 Generally, as the temperature of a solution increases, the amount of solute that dissolves increases.

6 The solubility curve represents the maximum mass of solute that can be dissolved at a set temperature. The green dot represents a supersaturated solution as it is above the solubility curve. The orange dot represents a saturated solution as it lies on the solubility curve. The blue dot represents an unsaturated solution as it lies below the solubility curve.

7 Different masses of the solute are weighed. Each sample is then placed in a test tube or beaker with a sample of water of a specific mass (for example, 1 g of solute in 20 g of water, 2 g of solute in 20 g of water etc.). The solutions are then heated until the solute dissolves and the heat source is removed. The solutions are watched carefully for the appearance of any crystals. The temperature at which the first crystal appears in each solution is recorded. This is known as the saturation point. The data is then extrapolated to provide an indication of how much solute will dissolve in 100 g of water. A set of axis of temperature vs mass of solute is created and each point plotted, and a smooth curved line of best fit is drawn as the solubility curve.

Analysis

8 40 g

9 20 g

10 **a** 40 g **b** 160 g **c** 20 g

11 **a** 48 g
  b 200 g
  c 70 g
The mass that will dissolve in each case is the mass found from the solubility curve multiplied by 2, as the question asks for the amount in 200 g of solvent, not 100 g.

12 Granite is a rock that is formed from magma cooling beneath Earth's surface. This means the crystals are formed slowly. Crystals formed from slow cooling are larger than those formed from cooling quickly. Magma that cools above the surface forms basalt; the cooling process is faster, and therefore the crystals in basalt are smaller than those in granite.

13 The solution would crystallise rapidly as the crack acts as a seed that disturbs the arrangement of solute in solvent.

14 Gases generally become less soluble with increases in temperature. Therefore, dissolved oxygen, which is vital for respiration, becomes less readily available for respiration of plants and animals. Dissolved carbon dioxide normally produced as a waste product of respiration more readily dissolves out of solution at higher temperatures, so overall the effect of increasing temperature on ocean temperatures would mean a reduction in concentration of both of these gases.

15 nitrogen and methane

16 0.40 mmol L^{-1}

17 A 110 g sample of sodium nitrate forms a saturated solution at 40°C. Therefore, only 10 g is needed to be added to make the solution saturated.

15.4 Precipitation reactions

TY 15.4.1 Compounds containing sodium ions or nitrate ions are usually soluble, so sodium nitrate will not form a precipitate.

Compounds containing sulfide ions are usually insoluble, so copper(II) sulfide will form as a precipitate.

TY 15.4.2 $Na^+(aq)$ and $SO_4^{2-}(aq)$ are spectator ions.

TY 15.4.3 $Mg^{2+}(aq) + S^{2-}(aq) \rightarrow MgS(s)$

15.4 Key questions

Retrieval

1 **a** a precipitate of silver carbonate
  b a precipitate of lead(II) hydroxide
  c a precipitate of magnesium sulfide
  d no precipitate

Silver, lead and magnesium ions are not found in the solubility table. However, the anions that each of these are combined with in these questions all form compounds that are usually insoluble. It is also worth noting that compounds containing sodium and nitrate ions are usually soluble.

The iron(II) ion is also not found in the solubility table. However, iron(II) does not appear in the 'exceptions' column of the table for either nitrates or sulfates, both of which generally form soluble compounds.

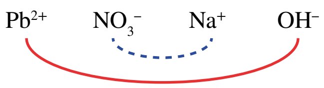
2 Chemicals containing the ions Na^+, K^+, NH_4^+ and NO_3^- almost never form a precipitate.

a **i** magnesium sulfide **ii** silver chloride
  iii aluminium hydroxide **iv** magnesium hydroxide

b **i** $K_2S(aq) + MgCl_2(aq) \rightarrow MgS(s) + 2KCl(aq)$
  ii $CuCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Cu(NO_3)_2(aq)$
  iii $AlCl_3(aq) + 3KOH(aq) \rightarrow Al(OH)_3(s) + 3KCl(aq)$
  iv $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

Comprehension

3



4 $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Analysis

5 A full chemical equation shows all ions present in both the reactant and products formed. In contrast, an ionic equation shows only the reactant ions that change state from aqueous form in the reactants to solid form in the products. For example, the complete chemical equation is $(NH_4)_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NH_4Cl(aq)$, whereas the ionic equation is $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

6 **a–d** Most sulfates and bromides are soluble so they are likely to be aqueous. Phosphates and sulfides are most often insoluble, so when reacted with Fe^{2+} they would form precipitates.

7 Sodium and potassium ions are the most soluble positive ions, so in a precipitation reaction they are more likely to act as spectators.

15.5 Concentration of solutions

TY 15.5.1 20.0 g L^{-1} **TY 15.5.2** 215 ppm
TY 15.5.3 0.48 mol L^{-1} or 0.48 M
TY 15.5.4 0.0666 mol L^{-1} or 0.666 M
TY 15.5.5 0.00250 mol
TY 15.5.6 0.0250 M
TY 15.5.7 c(NaOH) = 400 ppm

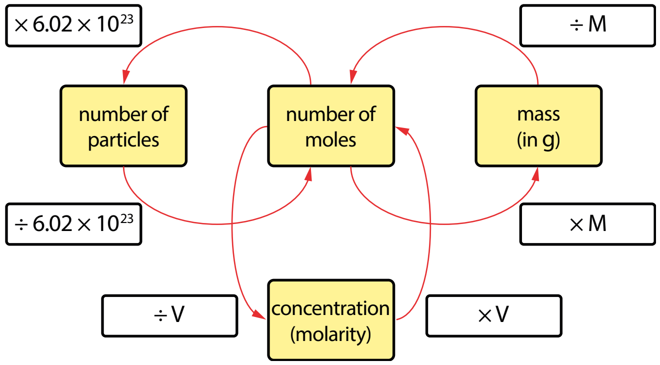
Retrieval

1 sugar = 14.0% (m/v) fat = 3.0% (m/v)

2 Household products most often use %m/m, %m/v or %v/v as these are easily understood measures of concentration using units that people are familiar with.

Comprehension

3



4 The concentration of mercury in fish is critical in ensuring humans are not susceptible to mercury poisoning. The more fish a person eats, the greater the chance of accumulating mercury in the body. The concentration of mercury in ppm is more readily understood by people and provides a more usable value in terms of 1 ppm. %m/m is more difficult and would produce a small number, giving the idea that the value is relatively safe when it in actual fact is not. Using mol L^{-1} is also not workable given the fish is a solid.

Analysis

5 **a** 0.4 M **b** 0.075 M **c** 0.025 M
6 **a** 5.0 ppm **b** 625 ppm **c** 27 ppm
7 0.856 M is less than 2.00 M so the 5.00 g of NaCl is concentrated.
8 0.0975 M **9** 214 ppb **10** 8.0×10^{-2} M
11 4.25 mol L^{-1} **12** 2.398 mol L^{-1}

Chapter 15 Review

Retrieval

1 B **2** C **3** C **4** A

5 **a** oxygen atom **b** hydrogen atom
  c hydrogen bond (and dispersion forces) **d** covalent bond

6 dispersion forces, hydrogen bonds, covalent bonds

7 Potassium ions and bromide ions are held in the ionic lattice by ionic bonds that are based on electrostatic attraction. These ionic bonds in the solute break when it dissolves in water.

Water, the solvent, has hydrogen bonds between water molecules. When potassium bromide is added to water, the hydrogen atoms of the water molecules are attracted to the negative bromide ions, and the oxygen atoms of the water molecules are attracted to the positive potassium ions.

Ion–dipole bonds form between the ions and water molecules and the surface ions are pulled into solution. Gradually, the ionic lattice dissociates and a solution is formed.

8 A solution is most likely to form when the polarity of bonding of the solute is similar to that of the solvent. The bonds formed between solute and solvent are then similar to those that existed between solute particles and between solvent particles. Water, being polar, is therefore a good solvent for ionic and polar substances.

9 7.11×10^{-3} mol

Comprehension

10 According to Tables 15.2.3 and 15.2.4, Fe^{2+} would precipitate in the presence of PO_4^{3-} and S^{2-} to produce $Fe_3(PO_4)_2$ and FeS.

11 **a** $NH_4Cl(aq) + AgNO_3(aq) \rightarrow NH_4NO_3(aq) + AgCl(s)$
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
b $FeCl_2(aq) + Na_2S(aq) \rightarrow FeS(s) + 2NaCl(aq)$
$Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
c $Fe(NO_3)_3(aq) + 3KOH(aq) \rightarrow 3KNO_3(aq) + Fe(OH)_3(s)$
$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
d $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
e $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$
$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

12 Water has a significantly higher melting point than the other hydrides due to the hydrogen bonds between water molecules. The other group 16 elements can only form dipole–dipole and dispersion forces, not hydrogen bonds.

13 $C_6H_{12}O_6$ and C_3H_7OH are polar molecules. They contain the polar –OH group and so are able to form hydrogen bonds with water.
HI and HNO_3 contain polar molecules but they are unable to form hydrogen bonds with water.
They dissolve by ionising. I_2, CH_4 and C_2H_4 are non-polar covalent molecules. They do not dissolve well in polar water.

14 Carbon is slightly more electronegative than hydrogen, so each C–H bond is slightly polar. However, the resulting partial charges are distributed symmetrically across the octane molecule, making the molecule non-polar overall.
The energy released in the formation of solute–solvent bonds is not enough to overcome the intermolecular bonds between solute molecules and the intermolecular bonds between solvent molecules, so octane will not dissolve in water.

15 When sugar dissolves in water, (1) the solute–solute bonds between sugar molecules must be broken, (2) the solvent–solvent bonds between water molecules must also be broken and (3) new solute–solvent bonds form.

16 1.00 M **17** 5.33 ppm

18 The electronegativity value for H is 2.2 and 3.4 for the oxygen atom. This large (1.2) difference in electronegativity between the O and H atoms creates strong intermolecular forces between neighbouring water molecules.
Therefore, it is the high polarity of the water molecule that allows relatively strong hydrogen bonding to occur between molecules. As a consequence, a relatively large quantity of energy is required to break the hydrogen bonds between water molecules when water changes from a liquid to a gas. This gives water a high boiling point.

Analysis

19 D

20 Using the 'like dissolves like' rule, only polar substances will dissolve in water. Nitrogen is a non-polar molecule so will not dissolve well. Ethene is a non-polar hydrocarbon that will not be expected to dissolve in water. Ethanol, however, is a small polar molecule that can be expected to dissolve in water. Full explanation:

- Nitrogen, N_2, is a covalent molecular substance. It is composed of non-polar molecules, so nitrogen would not be expected to dissolve well in water.
- Ethane, C_2H_6, is also a molecular compound. Ethane molecules contain slightly polar covalent bonds that are distributed symmetrically across the molecule. This makes the entire molecule non-polar, so ethane would be expected to not dissolve in water.
- Ethanol, C_2H_5OH, is a molecular compound that dissolves in water by forming hydrogen bonds with water molecules.

21 DDT is most likely non-polar because it is soluble in fats, which are non-polar, and insoluble in polar water.

22 160 g to be dissolved, therefore temperature must be 10°C

23 **a** 180 g **b** 30 g

24 Potassium ions and bromide ions are held in the ionic lattice by ionic bonds that are based on electrostatic attraction. These ionic bonds in the solute break when it dissolves in water. Water, the solvent, has hydrogen bonds between water molecules. When potassium bromide is added to water, the hydrogen atoms of the water molecules are attracted to the negative bromide ions, and the oxygen atoms of the water molecules are attracted to the positive potassium ions. Ion–dipole bonds form between the ions and water molecules and the surface ions are pulled into solution. Gradually the ionic lattice dissociates and a solution is formed.

25 0.80 M

26 Ionic substances form ion–dipole interactions with water. The cation forms these associations with the negative oxygen atom due to its electronegativity and subsequent polarity. The anion associates with the more positive hydrogen atom in the same manner. In some molecular substances, the overall difference in electronegativities of the atoms is so great that the polarity of the molecule means they break apart in water. The electrons that make up the covalent bond remain with the more electronegative atom, giving it a negative charge. The positive and negative ions formed then behave in a similar manner to ionic compounds on dissolving.

27 **a** $Cu^{2+}(aq)$ and $NO_3^{2-}(aq)$ **b** $Zn^{2+}(aq)$ and $SO_4^{2-}(aq)$
c $NH_4^+(aq)$ and $PO_4^{3-}(aq)$

28 **a** Intermolecular forces are those between one molecule and other molecules. For water, these are hydrogen bonds. Intramolecular forces are those holding the atoms together within a molecule. For water, these are covalent bonds.
b Covalent bonds are stronger. Evidence for this is the high temperatures required to break the bonds between the oxygen and hydrogen atoms inside the water molecule and so decompose it into its constituent gases. Changing liquid water into gaseous water involves breaking hydrogen bonds to separate one molecule from another. The lower temperatures needed to do so indicate that hydrogen bonds are weaker.

29 0.300 M

30 **a** Na_2SO_4, K_2SO_4, $(NH_4)_2SO_4$
b $CaSO_4$, $BaSO_4$, $PbSO_4$ or any other suitable answer

31 9 mL

32 a, c, e are all soluble. b and d are both insoluble.

33 concentration = $\frac{2.00}{0.0050}$ = 400 mg L^{-1}, 4.0×10^{-2} %(m/v)

34 A precipitate forms from two aqueous ions that when combined form an insoluble compound whereas crystallised solutes form from soluble solutes when the solution they are found in is cooled or solvent evaporates so the solute is too saturated to dissolve.

35 500 mL **36** 404 ppm

Knowledge utilisation

37 Water has a high latent heat of vaporisation. Water is effective as a cooler because it absorbs a relatively large amount of energy when it evaporates, giving it a high latent heat of vaporization.

38 **a** The crystal lattice of ice is disrupted and molecules have greater freedom of movement. During this time, all solid ice is being converted to liquid water.
b The added energy is taken up in overcoming the hydrogen bonds between molecules, separating the molecules to form a gas.

39 Y, Z, X. As the CH_3OH molecule is polar, it will dissolve readily in polar solvents. Because CH_3OH dissolves in Y, it can be ascertained that Y is a polar solvent. The non-polar nature of a CH_4 molecule means CH_4 will dissolve readily in non-polar solvents. It can be concluded that X is a non-polar solvent because CH_4 completely dissolves in it. CH_3OH molecules and CH_4 molecules partially dissolve in Z, which indicates it is more polar than X. No information is given about methanol dissolving in Z, although it can be implied.

40 Adding aqueous silver nitrate to a sample of each solution would distinguish between the two solutions. The sample that contained NaCl would form a white precipitate of AgCl(s); the glucose solution would not form a precipitate. Note electrical conductivity could also be used as the solution of dissociated ions (NaCl) would conduct an electrical current but glucose is a neutral covalent compound so would not conduct an electrical current.

41 **a** 700 g **b** 45 g **c** 66 g **d** 6.7 M

42 aluminium ions, Al^{3+}

43 Two anions that would not be helpful are NO_3^-, as no precipitate would form, and I^-, as all three precipitates would be a yellowy colour (namely mustard, yellow, light yellow) and it would be difficult to determine which cation is present.

Chapter 16 An introduction to acids and bases

16.1 Introducing acids and bases

16.1 Key questions

Retrieval

1 Acids—ethanoic acid found in vinegar, citric acid found in citrus fruits, carbonic acid found in carbonated soft drinks and beer, ascorbic acid found in citrus fruits. Bases—ammonia found in household cleaning agents, magnesium hydroxide in antacid, sodium carbonate in glass.

2 **a** OH^- **b** H_3O^+ **c** H_3O^+

Comprehension

3 A strong acid is one that readily dissociates in water to form H^+ ions and then reacts with water to form the hydronium ion (H_3O^+). An example of this is hydrochloric acid where almost none of the original HCl molecules remain undissociated in solution. In the case of ethanoic acid, a weak acid, dissociation is only partial and some of the original CH_3COOH molecules remain in solution, as do some of the ethanoate ion (CH_3COO^-) and H_3O^+ ions formed when dissociation does occur.

Analysis

4 The model does not explain why some substances that do not contain hydrogen form acidic solutions when mixed with water. For example, acidic solutions are formed when carbon dioxide, CO_2, or sulfur dioxide, SO_2, are dissolved in water.
The model also does not explain why some substances such as ammonia, NH_3, or sodium hydrogen carbonate, $NaHCO_3$, form basic solutions when mixed with water, even though they do not contain hydroxide ions. The Arrhenius model is also restricted to acids and bases that dissolve in water.

16.2 pH: a convenient way to measure acidity

TY 16.2.1 1.00 M **TY 16.2.2** 60.0 mL

16.2 Key questions

Retrieval

1 **a** pink/magenta **b** wine – solution
c ethanol – solute **d** sugar – solute

2 rose petals, red cabbage and blackberries

3 A solution with a pH of 7.4 is basic.

Comprehension

4 Remember: $pH = -\log_{10}[H_3O^+]$ $pH = 3 \therefore [H_3O^+] = 10^{-3}$ M $pH = 5 \therefore [H_3O^+] = 10^{-5}$ M $\therefore$ the difference is a factor of 100.

Analysis

5 0.075 M

6 30 mL of water needs to be added.

7 100.0 mL **8** 111 mL

16.3 Reactions of acids and bases

TY 16.3.1 $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Note that if hydronium ions are represented as $H_3O^+(aq)$, rather than as $H^+(aq)$, this reaction would be written as:

$H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$

TY 16.3.2 $H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$

Note that if hydronium ions are represented as $H_3O^+(aq)$, rather than as $H^+(aq)$, this reaction would be written as:

$H_3O^+(aq) + HCO_3^-(aq) \rightarrow 2H_2O(l) + CO_2(g)$

TY 16.3.3 $6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

16.3 Key questions

Retrieval

1 **a** $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
b $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
$Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
c $Zn(s) + 2CH_3COOH(aq) \rightarrow Zn(CH_3COO)_2 + H_2(g)$
$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
d $2Al(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3(aq) + 3H_2(g)$
$2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

2 **a** magnesium sulfate **b** calcium chloride
c zinc acetate **d** aluminium nitrate

3 **a** **i** $ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$
ii $ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$
b **i** $Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$
ii $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
c **i** $Cu(OH)_2(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l)$
ii $Cu(OH)_2(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2H_2O(l)$
d **i** $Mg(HCO_3)_2(s) + 2HCl(aq) \rightarrow MgCl_2(s) + 2H_2O(l) + 2CO_2(g)$
ii $Mg(HCO_3)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l) + 2CO_2(g)$
e **i** $SnCO_3(s) + H_2SO_4(aq) \rightarrow SnSO_4(s) + H_2O(l) + CO_2(g)$
ii $SnCO_3(s) + 2H^+(aq) \rightarrow Sn^{2+}(s) + H_2O(l) + CO_2(g)$

Comprehension

4 Limewater is made up of calcium hydroxide, a basic solution. The calcium (Ca^{2+}) ions and hydroxide (OH^-) ions dissociate in solution. When the carbon dioxide comes into contact with the calcium ions, a calcium carbonate salt is produced in water. This calcium carbonate is insoluble and therefore makes the solution appear white or cloudy in appearance.

Analysis

5 **a** $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
b $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
c $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
$Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
d $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
$CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$

6 **a** $KHCO_3(aq) + HF(aq) \rightarrow KF(aq) + H_2O(l) + CO_2(g)$
b $Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
c $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
d $NaHCO_3(s) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l) + CO_2(g)$

7 a metal carbonate

16.4 Human impacts on water quality

16.4 Science as a Human Endeavour

Acid rain

1 $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$

2 Acid rain can cause a reduction in growth of crops for food and affect building materials for shelter, so indirectly humans are affected.

Water quality

1 Students' own answers

2 Water can be used as a coolant in energy production, such as in power plants. It is used to remove wastes or as a cleaning agent in industries ranging from food production to car manufacture. Farming activities use large volumes of water for irrigation.

Chapter 16 Review

Retrieval

1 E **2** A **3** D 4 C

5 **a** $PO_4^{3-}(aq) + H_2O(l) \rightarrow HPO_4^{2-}(aq) + OH^-(aq)$
  b $H_2S(aq) + H_2O(l) \rightarrow HS^-(aq) + H_3O^+(aq)$

6 $H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$

Comprehension

7 **a** $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$
  b $H_2SO_4(aq) + K_2CO_3(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)$
  c $2H_3PO_4(aq) + 3Ca(HCO_3)_2(s) \rightarrow Ca_3(PO_4)_2(s) + 6H_2O(l) + 6CO_2(g)$
  d $2HF(aq) + Zn(OH)_2(s) \rightarrow ZnF_2(aq) + 2H_2O$

8 $HBr(g) \rightarrow H^+(aq) + Br^-(aq)$

9 $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2$

10 Limestone is composed of calcium carbonate. When acid rain reacts with calcium carbonate it forms calcium sulfate, carbon dioxide and water. This calcium sulfate is a soft, slightly water-soluble compound. Limestone statues and buildings therefore appear pitted due to deterioration of the remaining calcium sulfate.

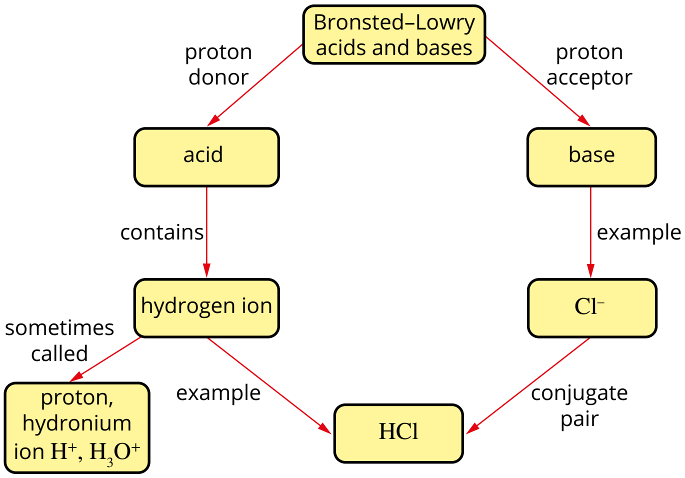
11 **a** $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$
  b $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

Analysis

12 i and ii. The Arrhenius theory gives no indication of why ammonia can form a basic solution when it dissolves in water, rather that a base dissociates in water to produce OH^- ions, which is not provided in this example

13 As the HCl is acidic, its pH would be less than 7. As NaOH is a base and its addition to HCl is a neutralisation reaction (where water and a salt are produced), the pH would increase towards 7. As more and more NaOH is added, eventually the pH would rise above 7 but this would mean the NaOH is in excess.

14



Knowledge utilisation

15 **a** These steps outline a possible experimental design that could qualitatively demonstrate the strengths of the acids.
  1 Ensure strips of zinc are of the same size.
  2 Clean each strip with steel wool and then place one strip into each of four test-tubes.
  3 Pour acid A into the first test-tube and begin timing how long it takes for the zinc strip to disappear.
  4 Write down your observations and time into the results table.
  5 Repeat steps 3 and 4 for each of the subsequent acids.

b The products that you would expect from mixing the zinc with an acid are hydrogen gas and a salt. This salt would be aqueous in solution.

c You could determine the strength of the acids by recording the time it took for the metal to disappear. The strongest acid would be the one with the shortest time and the weakest acid would be the one with the longest time.

d The strongest acid based on the results above is acid B.

e

Acid added to the zinc strip	Observations	Time taken for reaction to go to completion (s)
A	some bubbling and a slow reaction where the metal disappears slowly	342
B	rapid bubbling and a vigorous reaction where the metal disappears almost instantly	22
C	some quite rapid bubbling and a fairly vigorous reaction where the metal disappears quickly	65
D	some bubbling and a slower reaction than with acids B and C; the metal disappears quite slowly	178

f The acids in order from weakest to strongest based on the table are A > S > C > B.

g A list of possible acids that could have been used in this experiment are:
Acid A = ethanoic acid
Acid B = nitric acid
Acid C = hydrochloric acid
Acid D = phosphoric acid

Chapter 17 Rates of chemical reactions

17.1 Investigating the rate of chemical reactions

17.1 Key questions

Retrieval

1 Increasing surface area (for a heterogeneous reaction); increasing the concentration of a reactant in solution (for solution-based reactions in the condensed states); increasing the pressure of a gaseous reactant (for reactions in the gaseous state); increasing temperature (for most reactions); adding a catalyst (for selected reactions).

2 **a** increasing **b** increasing
  c increasing **d** increasing

Comprehension

3 **a** increase **b** increase **c** decrease

4 The spoiling of milk is a biochemical reaction catalysed by bacteria, which occurs faster at room temperature compared to the low temperatures (~4°C) in refrigerators.

Analysis

5 This observation illustrates that increasing the concentration of a reactant in solution increases the rate of reaction.

6 **a** break or cut the wood into small pieces to increase the surface area of the solid reactant
  b increase the temperature of the oven

17.2 Quantifying rates of chemical reactions

TY 17.2.1 1.70×10^{-3} M s^{-1}

TY 17.2.2 Rate $= -\frac{\Delta[F_2]}{\Delta t} = -\frac{1}{2}\frac{\Delta[ClO_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[ClO_2F]}{\Delta t}$

17.2 Key questions

Retrieval

1 M s^{-1} (which can also be expressed as mol L^{-1} s^{-1}).
2 instantaneous rate, initial rate and average rate

Comprehension

3 a

$$CaCO_3(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

b The rate of reaction is decreasing with time. This is evident in the slope of the curve, which is steep in the early part of the reaction and flattens out in the latter part of the reaction. The initial steep part of the curve represents a relatively large change in the amount of CO_2 produced in a relatively short period of time, i.e. a relatively fast rate of reaction. At later times, the curve is flatter and relatively small amounts of CO_2 are produced in the same time period, i.e. a relatively slow rate of reaction.

Analysis

4 The instantaneous rate of reaction is the rate of reaction at any time point during the reaction. The initial rate is the rate of reaction at the beginning of the reaction and can be considered the instantaneous rate at time $t = 0$.

5 To answer this question, it is important to understand the difference between the rate of reaction and the rate of consumption of a reactant. The rate of a chemical reaction at any point in time has only one value and is independent of the reaction stoichiometry. The rate of consumption of a reactant is dependent on stoichiometry and will be different to the rate of reaction if the stoichiometric coefficient is a number other than one—as is the case here.

$$\text{Rate of consumption of X} = -\frac{\Delta[X]}{\Delta t}$$

$$\text{Rate of reaction (with repsect to X)} = -\frac{1}{2}\frac{\Delta[X]}{\Delta t}$$

6 a 5.60×10^{-3} M s^{-1} b 1.5×10^{-3} M s^{-1}
7 a average rate of decomposition of $SO_2 = 1.55 \times 10^{-2}$ M s^{-1}
b average rate of reaction = 7.75×10^{-3} M s^{-1}

17.3 Collision theory

TY 17.3.1 The surface area of the iron anchor is relatively small so the frequency of collisions with reacting particles would be low. The concentration of oxygen at greater depths is also low so the frequency of collisions is further reduced. Therefore, the rate of corrosion is reduced.

17.3 Key questions

Retrieval

1 Reactant particles must collide with sufficient energy and at the correct orientation for products to be formed.
2 The minimum energy required by reactants for a reaction to occur; symbol E_a. This energy is needed to break the bonds between atoms in the reactants to allow products to form.
3 a lower b higher

Comprehension

4 The higher concentration of acid results in an increase in the number of collisions per unit time between the reactant particles. The increase in the rate of collisions results in an increase in the rate of reaction.
5 a very large b low c too low

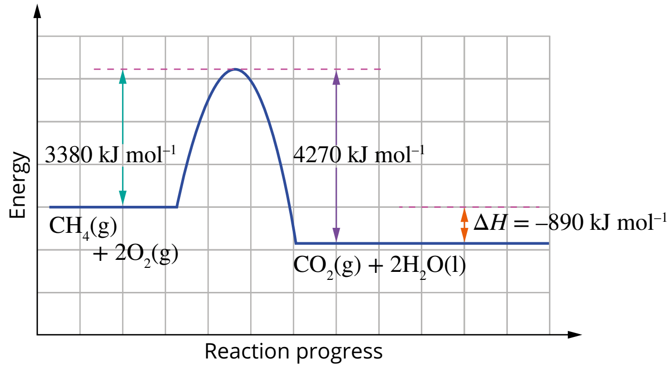
Analysis

6 a At higher temperatures, the molecules that react to form fibreglass plastics have greater energy. They collide more frequently and are more likely to have a total energy exceeding the activation energy of the reaction involved, increasing the rate of reaction. An increased rate of reaction allows for completion of the reaction in a shorter time.

b Fine particles have a large surface area to mass ratio, resulting in a high frequency of collisions of aluminium particles with gas molecules (such as oxygen) in the air and hence rapid reaction rate. The aluminium can burn vigorously and release a large quantity of heat.

c At high altitude, air pressure is considerably lower than at any location in the Australian bush so water boils at a lower temperature (up to 30°C lower). Thus, the average kinetic energy of the molecules in the potato in Nepal is lower, so they are less likely to have a total energy exceeding the activation energy of the reactions involved in cooking a potato, therefore the potato cooks more slowly.

7 a $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ b −890 kJ mol^{-1}
c



8 139 kJ mol^{-1} 9 D, F, B, E, C, A

17.4 Catalysts

17.4 Science as a Human Endeavour

From lock-and-key to induced fit—changing scientific ideas

1 Students' own answers
2 Students' own answers

17.4 Key questions

Retrieval

1 A substance that increases the rate of reaction without being consumed in the reaction. A catalyst provides an alternative reaction pathway with a lower activation energy.
2 a large biological molecule (typically a protein) that acts as a catalyst for a specific biochemical reaction
3 Step 1: E + S → E–S
Initial binding of the substrate with the active site to form the enzyme–substrate complex (E–S).
Step 2: E–S → E–P
Conversion of the enzyme–substrate complex (E–S) to the enzyme–product complex (E–P). This process involves bond stretching and bond breaking and so typically requires the highest energy of the three-step process.
Step 3: E–P → E + P
The breakdown of the enzyme–product complex (E–P) leading to the production of free products and regeneration of the free enzyme, thus completing the catalytic cycle.
4
  1 The activation energy of a chemical reaction can be lowered by distorting the bound substrate molecule(s) to stretch and ultimately break the bonds required for the reaction to proceed.
  2 The activation energy of a chemical reaction can be lowered by better orientating the substrate molecule(s) into an optimal arrangement contributing to a lowering of the energy required to stretch and break bonds.
  3 In some cases, the enzymes form temporary covalent bonds with the substrate molecule(s) to provide a lower energy transition state.
  4 The active site of the enzyme moulds itself to provide an energetically favourable environment to stabilise the transition state at a lower energy.

Comprehension

5 Catalysts lower the activation energy of chemical reactions. This means that at a given temperature, the presence of a catalyst will result in a greater proportion of reactant particles having the required energy to overcome the (lowered) activation energy barrier compared to the uncatalysed reaction. If more reactant particles have the required energy for products to be formed (compared to the uncatalysed reaction) then the rate of reaction will be increased.

6 **a** Reactions involving a heterogeneous catalyst take place at the surface of the catalyst. Reactants form bonds with the catalyst, lowering the activation energy of reactions and allowing them to proceed more rapidly.

b A porous pellet has a much larger surface area than a solid lump. More reactants may be in contact with the surface of a porous pellet at any instant, producing a faster rate of reaction.

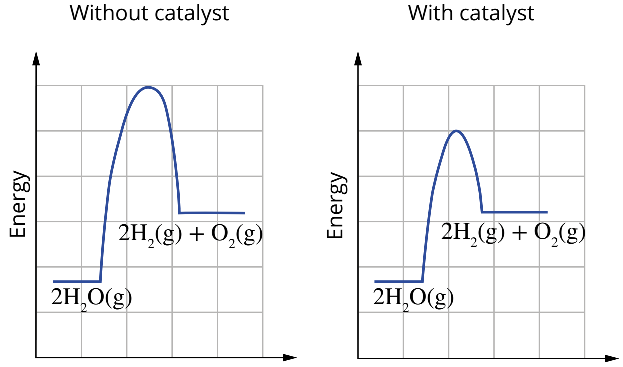
Analysis

7 Increasing the temperature of a reaction increases the kinetic energy of reacting particles and therefore the rate at which reacting particles collide. This in turn increases the frequency of collisions that occur with the correct orientation. Both of these factors increase the rate of reaction. More importantly, however, the increase in temperature increases the energy of collisions meaning a greater proportion of reacting particles will have the energy required to overcome the activation energy barrier. This factor, more than any other, results in an increase in reaction rate.
By contrast, enzymes act to lower the activation energy of chemical reactions by providing an alternative reaction pathway. This means that a greater proportion of reacting particles will have sufficient energy to overcome the (lowered) activation energy barrier, thus allowing the reaction to proceed at a faster rate.
In summary, both increasing temperature and the addition of a catalyst allow a greater proportion of reacting particles to collide with enough energy to allow the reaction to proceed at a faster rate; however, these approaches have different mechanisms for this to occur.

8 When salt is mixed with sugar, the salt acts as a catalyst and lowers the activation energy of the combustion reaction between sugar and oxygen.

9 **a** 1370 kJ mol^{-1} **b** –572 kJ mol^{-1}

c


Without catalyst
With catalyst

d +572 kJ mol^{-1}

Chapter 17 Review

Retrieval

1 D **2** B **3** C, D **4** D

5 Reactant particles must collide with sufficient energy above a minimum threshold and with the correct orientation for products to be produced. The minimum threshold energy is called the activation energy.

Comprehension

6 B **7** B

8 Collision 1 is likely to be more successful since this is the representation in which the orientation of colliding particles is most likely to lead to a successful collision. This is because the orientation in collision 1 shows the parts of the molecules that would react (shown in red and black) are brought into close proximity to each other and are more likely to collide than in the other collisions.

9 The frequency of collisions between the $Cu^{2+}(aq)$ ions and powdered zinc is greater than with the large zinc pieces. This is due to zinc powder having a larger surface area than zinc pieces.

10 Colliding particles must have kinetic energy equal to or greater than the activation energy of the reaction.

Analysis

11 The instantaneous rate of reaction is the rate of reaction at any given time instant throughout the course of a chemical reaction. Instantaneous rates are measured by determining the slope of a concentration–time curve at any time point during a chemical reaction. In practice, measuring the instantaneous rate at any point in time can be difficult for many chemical reactions as it requires the construction of a complete concentration–time curve. The most common rate of reaction that is calculated is the average rate of reaction, which is determined by measuring a change in concentration between two different points in time and dividing that change by the time interval.
The initial rate is the instantaneous rate of reaction measure at time, $t = 0$, and is determined by measuring the initial slope of a concentration–time curve at zero time.

12 **a** The reaction rate will decrease due to a decrease in concentration of HCl, which means that the average rate of collisions between reactant particles decreases.

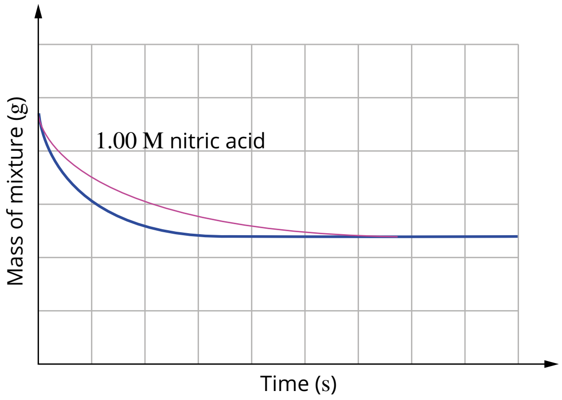
b The reaction rate will increase due to the larger surface area of the solid calcium carbonate powder. This increased surface area means the average rate of collisions between reactant particles will increase.

c There is insufficient information to determine exactly how the reaction rate will change. The decrease in HCl concentration will result in a lower rate of collisions between reactant particles but this will be offset by the increased surface area of the calcium carbonate powder. There is insufficient information to determine which effect will dominate.

13 **a** A gas is produced so mass is lost from the mixture.

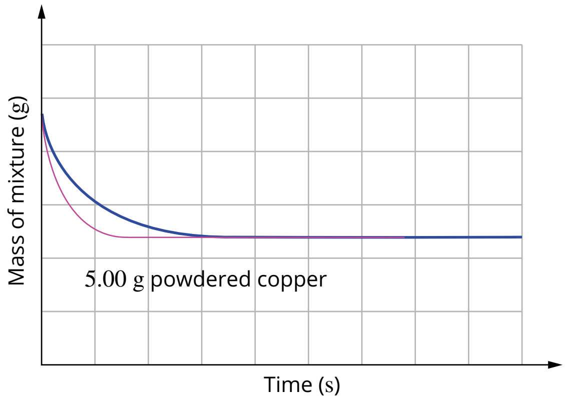
b Nitric acid is in excess, i.e. Cu is limiting.

c



Decreased rate of mass loss due to lower nitric acid concentration.
($n(HNO_3) = cV = 1.00 \times 0.500 = 0.500$ mol)
Copper is still limiting, so final mass remains the same.

d



Increased rate of mass loss due to increased copper surface area. Copper is still limiting, so final mass remains the same.

14 **a** rate of disappearance of A = $-\frac{\Delta[\text{A}]}{\Delta t}$

rate of disappearance of B = $-\frac{\Delta[\text{B}]}{\Delta t}$

rate of appearance of C = $-\frac{\Delta[\text{C}]}{\Delta t}$

b No, the rates would not be equal. This is due to the fact that each of the substances in this reaction has different stoichiometric coefficients. The rate of disappearance of reactant A (with a coefficient of 3) would have the highest rate value; it would disappear three times faster than B and $\frac{3}{2}$ times faster than the appearance of product C. The rate of appearance of product C (with a coefficient of 2) would have the second highest rate value. The rate of disappearance of B (with a coefficient of 1) would have the lowest rate value, i.e.

$-\frac{\Delta[\text{A}]}{\Delta t} > \frac{\Delta[\text{C}]}{\Delta t} > -\frac{\Delta[\text{B}]}{\Delta t}$

c You need to be careful not to confuse expressions for the rate of disappearance (or rate of appearance) of individual substances with the expression for the overall rate of reaction. The rate of reaction at any time during a chemical reaction is independent of the reaction stoichiometry. To account for the different stoichiometric coefficients, you need to divide the rates of appearance/disappearance by the respective stoichiometric coefficient, i.e.

rate of reaction = $-\frac{1}{3}\frac{\Delta[\text{A}]}{\Delta t} = -\frac{\Delta[\text{B}]}{\Delta t} = \frac{1}{2}\frac{\Delta[\text{C}]}{\Delta t}$

All three expressions for rate are now mathematically equivalent.

15 **a** The rate of reaction can be determined from the slope of a concentration–time curve. The region in which the rate of reaction is constant will be the region where the slope of the curve is unchanging (or constant). This is region C. In fact, the slope of the curve in region C is zero indicating that the net overall reaction rate is also zero during this time period. Throughout regions A and B the slope of the curve is changing and therefore the rate of reaction is changing.

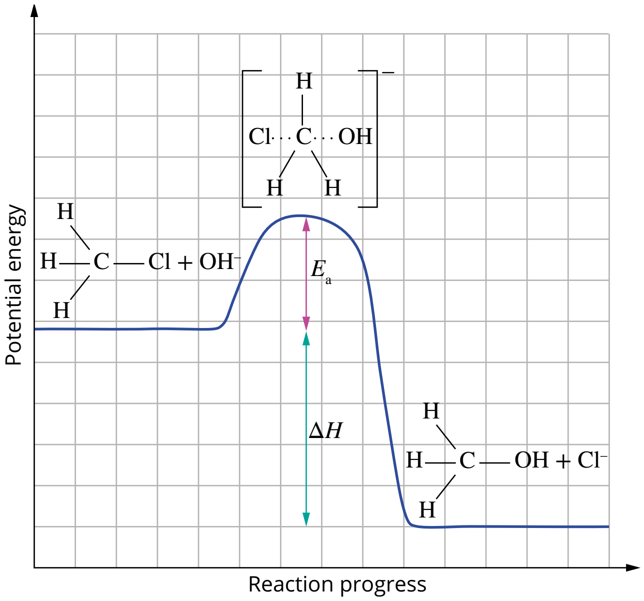
b The region of the curve showing the fastest rate of reaction will be the region where the slope of the concentration–time curve is the greatest. This is region A.

16 **a** average rate of disappearance of $H_2O_2 = 1.50 \times 10^{-2}\,\text{M s}^{-1}$

b average rate of reaction = $7.50 \times 10^{-3}\,\text{M s}^{-1}$

c The average rate of disappearance of H_2O_2 is twice the magnitude of the average rate of reaction. The reason for this is that a measure of the rate of disappearance (or rate of appearance) of a substance does not need to take into account the stoichiometric coefficient of that substance. However, a measure of the average rate of reaction needs to take account of the stoichiometric coefficient—hence the $\frac{1}{2}$ term in the expression for rate of reaction.

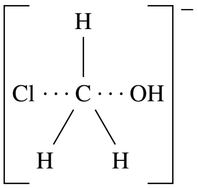
17 **a**



b Activation energy is the minimum amount of energy required by the reactants in order to form products in a reaction. It is related to the energy required to break bonds in the reactants.

c exothermic

d



e C–Cl bond being broken in chloromethane; C–O bond beginning to form between chloromethane and hydroxide ion

18 Heterogeneous catalysts are catalysts that are in a different physical state from the reactants and products of the chemical reaction. Homogeneous catalysts are in the same physical state as the reactants and products of the chemical reaction. Enzyme catalysis can involve both homogeneous catalysis (where the enzyme, substrate and products are freely soluble in solution) or heterogeneous catalysis where the enzyme molecule is embedded into the solid cell wall structure of living tissue (with the substrate and product molecules being soluble in solution).

19 **a** Decrease in reaction rate. Decreasing the concentration of reactants will decrease the frequency of collisions between reactant particles.

b Increase in reaction rate. Increasing temperature will increase the frequency of collisions but more importantly will increase the energy of collisions so that more reactant particles will have the energy required to overcome the activation energy barrier. Both effects (especially the latter) will increase the rate of reaction.

c Increase in reaction rate. Addition of catalyst will lower the activation energy of the chemical reaction, which means that a greater proportion of reactant collisions will have the requisite energy to result in product formation. This will lead to a greater rate of reaction.

20 **a** surface area of a solid reactant, concentration of reactants in a solution, pressure of any gaseous reactants, temperature of the reaction, presence of a catalyst

b **i** surface area of a solid reactant, concentration of reactants in a solution, pressure of any gaseous reactants

ii temperature of the reaction (also increases collision frequency), presence of a catalyst

21 1 Increasing the pressure of the gases increases the number of collisions between the reactant molecules in a given time, so more collisions occur with the correct orientation to react and with energy that is greater than or equal to the activation energy. As a result, the rate of reaction would increase.

2 Adding a catalyst allows the reaction to occur by a different pathway with a lower activation energy. The proportion of collisions with energy greater than the activation energy would thus be increased. As a result, the rate of reaction would increase.

22 **a** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

b $n(CaCO) = \frac{10.0\,g}{100.1\,g\ mol^{-1}} = 0.0999\,mol$

$n(HCl) = 0.1\,M \times 0.1\,L = 0.01\,mol$

$\therefore$ $CaCO_3$ is in excess

c The rate of reaction can be measured by:

- a decrease in mass of reaction mixture as $CO_2(g)$ escapes to the atmosphere
- an increase in pH with a pH probe as acid is consumed.

d The rate of reaction with the smaller lumps will be faster. The smaller lumps have a larger surface area so more collisions can occur per second.

e Increase temperature (which would cause particles to move faster, resulting in more collisions, increasing the likelihood of some of those collisions to be successful collisions, which ultimately results in a faster rate of reaction); increase concentration of hydrochloric acid

Knowledge utilisation

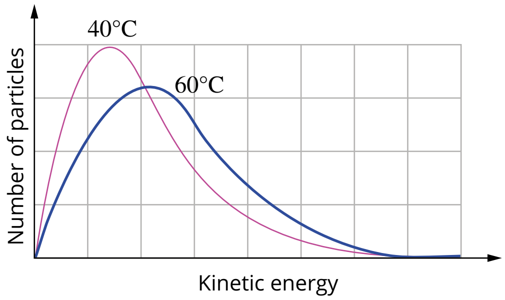
23 **a** The single H–H bond in each hydrogen molecule and the double O=O bond in each oxygen molecule are broken in the reaction.

b Two H–O bonds in each new water molecule are formed.

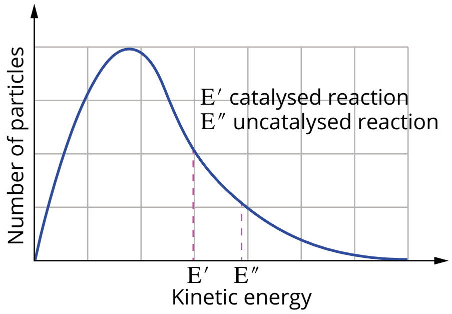
c The energy change for the reaction is the difference between the energy absorbed to break the bonds in the H_2 and O_2 reactants, and the energy released when the bonds in the H_2O product are made.

d No reaction occurs until sufficient energy is supplied to overcome the activation energy.

24 **a**

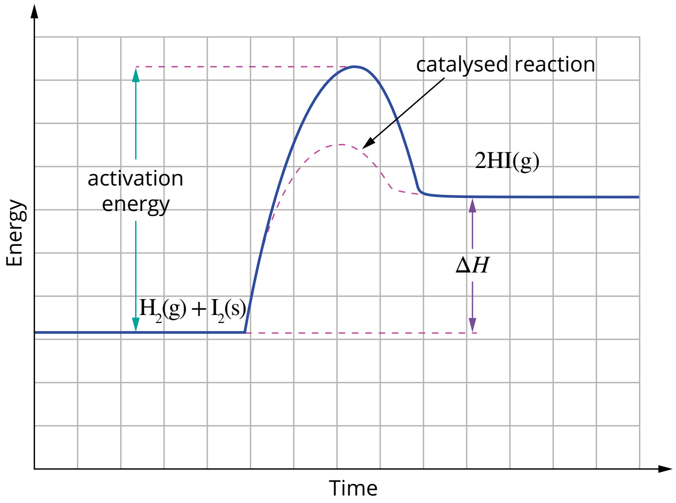


b



c When a catalyst is present, the reaction proceeds by an alternative reaction pathway with a lower activation energy (*E*′) than in the uncatalysed reaction (*E*″). At a given temperature, for the lower activation energy *E*′, a larger proportion of reacting particles will have kinetic energy equal to or greater than the activation energy, than if the activation energy were the higher *E*″. As more reactants have sufficient energy to react, the rate of reaction increases.

25 a, c



b The reaction is endothermic.

26 **a** $S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(g) + H_2O(l)$

b The reaction could be monitored with respect to time by observing the appearance of the solid sulfur precipitate. As the solid forms in the solution, light passing through the solution is scattered. (This is known as the Tyndall effect.) By observing the extent of scattering of light, the rate of reaction can be estimated. One way of observing the scattering of light would be to place the reaction mixture on top of a cross marked on a sheet of paper. The time taken for the cross to disappear when viewed from above the beaker would be a measure of the rate of production of the sulfur precipitate.

c Increasing the temperature should result in an increase in the reaction rate. The major reason for this is that increasing temperature increases the proportion of reactant particles having the necessary energy to overcome the activation energy barrier. This will mean that a higher proportion of collisions will be successful, i.e. resulting in product formation.

d Increasing the HCl concentration should result in an increase in the reaction rate. The reason for this is that increasing reactant concentration increases the rate of collisions between reactant particles. This will increase the rate at which successful collisions occur (i.e. those collisions with the correct orientation and sufficient energy that lead to product formation).

Glossary

RED ENTRIES Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority.

A

absolute uncertainty The uncertainty of a measurement due to the smallest available increment on an instrument.

absorption The process by which one substance permeates another substance.

accelerator mass spectrometry A highly sensitive analytical technique that measures long-lived radionuclides that occur naturally in our environment.

accuracy The condition or quality of being true, correct or exact; freedom from error or defect; precision or exactness; correctness; in science, the extent to which a measurement result represents the quantity it purports to measure; an accurate measurement result includes an estimate of the true value and an estimate of the uncertainty.

acid A substance capable of donating a hydrogen ion (proton).

acid rain Rainwater that has reacted with acidic emissions from industry and vehicles and has a pH less than 5.5.

acid–base reaction A reaction in which one substance, an acid, donates a hydrogen ion (proton) to another substance, a base.

acidic solution An aqueous solution in which the concentration of hydronium ions (H_3O^+) is greater than the concentration of hydroxide ions (OH^-). At 25°C, pH < 7.

acidity The concentration of H_3O^+ ions in an aqueous solution. Acidity is measured using the pH scale.

activation energy The minimum energy required by reactants for a reaction to occur; symbol E_a. This energy is needed to break the bonds between atoms in the reactants to allow products to form.

active site The region of an enzyme where the substrate molecule(s) binds and catalysis takes place.

actual yield The mass of product actually obtained during a chemical reaction. This will be less than or equal to the theoretical yield.

adsorption The attraction and binding of molecules or particles of one substance to the surface of another.

alkali A soluble base or a solution of a soluble base.

alkali metal A group 1 metal—Li, Na, K, Rb, Cs and Fr.

alkaline earth metals A group 2 metal—Be, Mg, Ca, Sr, Ba and Ra.

alkane Saturated hydrocarbons of various lengths with the general formula C_nH_{2n+2}.

allotrope Different forms of the same element in which the atoms combine in different ways.

alloy A substance formed when other materials (e.g. carbon, other metals) are mixed with a metal.

alpha particle A particle composed of two protons and two neutrons, also known as a helium nucleus.

amalgam A homogeneous mixture (or solution) of two or more metals where mercury acts as the solvent and the other metals are solutes dissolved in mercury.

amino acids The monomer molecule used to make proteins.

amphiprotic A substance that has the ability to act as an acid (donate a proton or hydrogen ions) and also to act as a base (accept a proton and hydrogen ions). All amphiprotic substances are amphoteric. Not all amphoteric substances are amphiprotic.

amphoteric The ability to act as an acid and also as a base.

anion A negatively charged ion, e.g. a chloride ion, Cl^-.

annealing Heating a metal to a moderate temperature and then allowing it to cool slowly to make it softer and more ductile.

anomaly Deviation from what is expected or considered to be normal or standard.

aqueous When a chemical species has been dissolved in water, the resulting solution is said to be aqueous. This can be shown by writing '(aq)' after the name or symbol of the chemical.

aqueous solution A solution in which water is the solvent. Typically represented with the (aq) after the chemical formula of the solute, e.g. NaCl(aq) represents the solute sodium chloride dissolved in water.

asymmetrical molecule A molecule in which the polar bonds are unevenly (or asymmetrically) distributed. The bond dipoles do not cancel out and a net dipole is formed.

atom The basic building block of matter. It is made up of subatomic particles—protons, neutrons and electrons.

atomic absorption spectroscopy (AAS) An analytical technique that uses light absorption to measure the concentration of a metal in a sample.

atomic mass The mass of an atom in atomic mass units (amu).

atomic mass unit (amu) One atomic mass unit is defined as a mass exactly equal to $\frac{1}{12}$ the mass of a ^{12}C atom.

atomic number The number of protons in an element's nucleus, Z; in a neutral atom, the number of protons will be equal to the number of electrons orbiting the nucleus.

atomic radius A measurement used for the size of atoms; the distance from the nucleus to the outermost electrons.

atomic theory of matter The proposal made by John Dalton that all matter is composed of atoms.

Aufbau principle The explanation of the most stable electron arrangement in an atom; electrons fill lower energy shells and subshells before residing in higher energy shells and subshells.

average bond enthalpy The average energy, taken across a wide variety of different molecules, required to break a specific type of bond.

average rate of reaction The change in concentration of a reactant or product per unit time measured at two different time points during a chemical reaction.

Avogadro's constant See *Avogadro's number*.

Avogadro's number The number of particles in a mole; symbol $N_A = 6.02 \times 10^{23}$ mol^{-1}.

B

balanced equation The number of each type of atom on the reactant side equals the number of each type of atom on the product side.

ball-and-stick model A model that displays both the three-dimensional position of the atoms and the bonds between them. The atoms are represented by coloured spheres that are connected by rods to represent the bonds.

Balmer series A series of spectral lines in the visible part of the hydrogen spectrum that is produced by transitions of electrons between the n = 2 energy level and higher energy levels of the hydrogen atom

bar A unit of pressure. 1 bar = 100 kPa. 1 bar is approximately equal to atmospheric pressure at sea level.

bar graph A graph that shows the measured value (of value of central tendency) of the dependent variable by the length of the horizontal bar.

base A substance capable of accepting a hydrogen ion (proton).

basic solution An aqueous solution in which $[H_3O^+] < [OH^-]$. For a basic solution at 25°C, pH > 7.

beta radiation Electrons emitted by radioactive decay.

bias An inclination or prejudice towards something.

biodiesel A renewable fuel usually produced from vegetable oil from sources such as soyabean, canola or palm oil.

bioethanol A renewable fuel produced from crops such as sugar cane.

biofuel A fuel derived from biomass.

biogas A renewable fuel that can be used to generate electricity.

blast furnace A type of furnace used for the extraction of metals from their ores, including iron from iron oxides.

block (periodic table) One of four main parts of the periodic table where elements have the same highest energy subshell filled, i.e. *s*, *p*, *d* or *f* subshell.

blueshifted The phenomenon that is observed when light or other electromagnetic radiation from an object is decreased in wavelength, or shifted to the blue end of the spectrum. It occurs when the object emitting the light is moving towards the observer.

Bohr model The description of each electron in an atom as occupying a discrete energy level in the region around the nucleus; the movement to a different energy level involves absorbing or emitting energy.

boiling point The temperature at which the vapour pressure of a liquid equals the external pressure on the surface of liquid. The boiling point is therefore dependent on the external pressure. It is typically characterised by the escape of gaseous bubbles of a substance from the liquid state of the substance.

boiling point elevation The increase in temperature between the boiling point of a solution and the boiling point of the pure solvent.

bond angle The angle formed between adjacent bonds on a central atom.

bond enthalpy The energy needed to break one mole of bonds of gaseous molecules under standard conditions.

bonding electrons The electrons used to form a covalent bond.

bonding pairs A pair of electrons that is shared between two atoms to form a covalent bond.

brittle Shatters when given a sharp tap.

buckyball A ball-like polyhedral molecule consisting of carbon atoms of the type found in fullerenes.

C

calibration Adjusting an instrument with a standard scale of readings.

calibration curve A graph of values for standard solutions of compound of interest where concentrations are known. Used as a reference to interpolate and determine the value for concentration of the compound as a sample of interest.

calorimeter An instrument designed to measure energy changes in a reaction. It is made up of an insulated container of water in which the reaction occurs, with a stirrer and thermometer to measure the temperature change during the reaction.

calorimetry The experimental method by which the heat energy released by a chemical reaction is measured.

carbon nanotube A hollow tube of graphene (carbon) with a diameter of one to tens of nanometres.

carbon neutral A process where there is no net change to the amount of atmospheric carbon. The amount of carbon absorbed by the process is equal to the amount produced by the process.

carrier gas in gas chromatography, the mobile phase or the 'moving phase' is a carrier gas, usually nitrogen.

catalysis The increase in the rate of a chemical reaction due to the presence of a catalyst.

catalyst A substance that increases the rate of a reaction but is not consumed in the reaction. The catalyst provides a new reaction pathway with a lower activation energy.

cathode ray tube a tube containing a vacuum or gas through which electron beams travel and produce visible light on a phosphorescent surface.

cation A positively charged ion, e.g. a sodium ion Na^+.

Celsius scale The temperature scale on which the freezing point of water is given a value of 0 and the boiling point a value of 100.

centrifugation A separation technique that uses the centrifugal force of rotational motion to promote rapid settling of solid particles in a heterogeneous solid–liquid mixture.

ceramic Relating to products made from clay and similar materials, such as pottery, brick, etc.

ceramic insulator Used to keep high-voltage power lines insulated from electricity poles and electric fence wires.

change of state When matter changes from one physical state to another.

charge imbalance When there are more protons than electrons in an atom.

chemical (HAZCHEM) code A system of codes and images that provides warning of hazards about items to users.

chemical change A chemical reaction. A change in the chemical composition of matter where one or more types of matter are transformed into a new type or new types of matter.

chemical energy The sum of the chemical potential energy and kinetic energy in a substance. Chemical energy is stored in bonds between atoms and molecules. The energy results from things such as attractions between electrons and protons in atoms, repulsions between nuclei, repulsions between electrons, movement of electrons, and vibrations and rotations around bonds.

chemical equation A symbolic representation of a chemical reaction using the chemical formulas of the substances involved in the reaction. A balanced chemical equation shows the relative amounts of reactants and products in the reaction. The reactants (written on the left-hand side) are connected by an arrow to the products (written on the right-hand side).

chemical formula A shorthand representation of a substance (element or compound) using chemical symbols. The chemical formula of a compound also uses numerical subscripts to convey the relative proportions of the different types of atoms in the compound.

chemical property A property of a substance observed during a chemical reaction.

chemical reaction A rearrangement of atoms where one or more types of matter are transformed into a new type or new type of matter. Typically involves the breaking and forming of chemical bonds.

chemical symbol A symbolic representation of an element, usually one or two letters, where the first letter is capitalised and the second letter is lower case, e.g. the chemical symbol for carbon is C and the chemical symbol for sodium is Na.

chromatogram The output of a chromatography procedure. In TLC and paper chromatography it is the pattern of bands or spots on the paper/plate. In GC and HPLC it is the graph or trace produced by the detector.

chromatography A technique for separating the components of a mixture, involving the movement of the sample through the adsorbent stationary phase by a mobile phase.

claim An assertion made without any accompanying evidence to support it.

climate change A variation in the climate of the Earth related to changes in the atmosphere.

coefficient (in chemistry) Whole numbers placed in front of formulas to make the equation balance.

coefficient of determination A calculated value that indicates the ability to predict the dependent variable value from the independent value.

coke A solid that contains 80–90% carbon. It is produced by strongly heating coal in the absence of air.

collision theory A theoretical model that accounts for the rates of chemical reactions in terms of collisions between particles occurring during a chemical reaction.

column chromatography Chromatography where the stationary phase is in a column, e.g. GC or HPLC.

column graph A graph that shows the measured value (of value of central tendency) of the dependent variable by the height of the column.

combustion A rapid reaction with oxygen accompanied by the release of large amounts of energy. Also called burning.

competing effects Effects whose impact acts to counteract the effects of one another.

complete combustion A hydrocarbon undergoes complete combustion with oxygen at high temperatures when the only products are carbon dioxide and water.

composite A material resulting from the combination of two or more different substances.

compound A pure substance composed of more than one type of atom chemically combined in fixed proportions.

concentrated solution A solution that has a relatively high ratio of solute to solvent.

concentration A measure of how much solute is dissolved in a specified volume of solution.

concentration–time curve A plot of reactant and/or product concentration versus time. It shows how the concentration of substances in a chemical reaction changes with time and can be used to determine the rate of reaction at any time.

condensation The change of physical state where a gas becomes a liquid.

condensed electron configuration A shortened writing of electron configuration by using the noble gas before the element.

conductivity Capacity to transfer electrical current.

conductor An object or type of material that permits the flow of electric charges, e.g. a wire is an electrical conductor that can carry electricity along its length.

confidence interval A calculated range in values that estimates where the true value is likely to be, according to a pre-determined level of confidence.

continuous quantitative data Data that cannot be counted; for example, the numbers between 1 and 2.

continuous variable Numerical variables that are based on a continuum in which infinite fractional values exist.

continuum (emission spectrum) The spacing between successive spectral lines within the emission spectrum becomes so small such that they are indistinguishable from one another.

control group The selected test group (sample group) in which no change to the independent exists and represents the typical natural conditions.

controlled variable A variable that is controlled during an experiment so that it does not change and does not influence the measured results or dependent variable.

covalent bond The force of attraction formed when one or more pairs of electrons are shared between two nuclei.

covalent layer lattice An arrangement of atoms in a lattice in which there are strong covalent bonds between the atoms that have formed in a layer.

covalent network lattice An arrangement of atoms in a lattice in which there are strong covalent bonds between the atoms in all three dimensions.

covalent network solid A giant covalent substance in which all the atoms are held together by covalent bonds to form a continuous three-dimensional structure.

crude oil Unprocessed mixture of fuels found in porous rock areas of the Earth's surface.

crystal A solid made up of atoms or molecules arranged in a repeating three-dimensional pattern.

crystal lattice The lattice of a crystalline solid.

crystallisation The process in which solid crystals are deposited when the concentration of a solute in a solution increases past the point of saturation.

crystallise Form solid crystals.

current A movement or flow of electric charges.

D

data In science, measurements of an attribute or attributes; data may be quantitative or qualitative and be from primary or secondary sources. (ACARA 2015c)

data loggers Portable tablets that record and log data acquired or measured using probes, sensors and instruments.

decanting A process of separating heterogeneous mixtures with at least one liquid phase. Decanting involves carefully pouring off the top liquid layer lying over a solid phase or another liquid phase.

decomposition A reaction in which a compound is broken down into smaller parts.

delocalise Spread out.

delocalised An electron that is not restricted to the region between two atoms.

delocalised electron An electron that is not restricted to the region between two atoms.

density A measure of the amount of mass per unit volume. It has the SI units of kg m^{-3}, but is commonly quoted in g cm^{-3}.

dependent variable The variable that is measured or observed during an experiment, which is assumed to change due to the independent variable.

deposition The change of physical state where a gas becomes a solid.

desorption The breaking of attraction between a substance and the surface to which the substance is adsorbed.

deviation The amount that a measurement differs from a fixed value.

diamond A form of pure carbon that is the hardest naturally occurring substance.

diatomic molecule A molecule formed from two atoms only, e.g. Cl_2.

dilute solution A solution that has a relatively low ratio of solute to solvent.

dilution The addition of a solvent to a solution to reduce its concentration.

dipole The separation of positive and negative charges in a bond or molecule.

dipole–dipole forces The attractive forces between polar molecules that result from the attraction between the positive and negative ends of the dipoles.

discrete (discontinuous) data Data that can be counted and whose values are separate and distinct.

dispersion forces The attractive forces caused by temporary dipoles that are the result of random movement of the electrons surrounding the molecules.

dissociate Break up.

dissociation A process in which molecules or ionic compounds separate or split into smaller particles such as atoms or ions. Examples of dissociation reactions include the solution of NaCl solid in water, forming Na$^+$(aq) and Cl$^-$(aq) ions, and the reaction of HCl gas with water, forming H$^+$(aq) and Cl$^-$(aq).

dissolution The process of dissolving a solute in a solvent to form a solution.

dissolve To incorporate a solid or gas into a liquid so as to form a solution.

distillation A purification technique used for separating mixtures containing a liquid component by first evaporating the liquid to its gaseous state and then recondensing it back to its liquid state. Distillation uses the differences in boiling points of the components of the mixture.

Doppler effect The change in frequency or wavelength observed in electromagnetic radiation by an observer moving relative to the source of the emitted wave.

double covalent bond A covalent bond in which four electrons (two electron pairs) are shared.

ductile Able to be drawn into a wire.

E

E10 petrol blend A blend of petrol containing 10% ethanol.

effective nuclear charge The charge experienced by the outer-shell electrons in an atom. It indicates the attractive force felt by the valence electrons towards the nucleus.

elastic collision An elastic collision is one in which there is no loss of kinetic energy.

electric dipole The separation of positive and negative charges in a bond or molecule.

electrode A solid conductor in a half-cell at which oxidation or reduction reactions occur.

electrolysis A process that produces a non-spontaneous redox reaction by the passage of electrical energy from a power supply through a conducting liquid.

electrolyte A chemical substance that conducts electric current as a result of dissociation into positively and negatively charged ions, which migrate towards the negative and positive terminals of an electric circuit.

electromagnetic radiation Radiant energy consisting of synchronised oscillations of electric and magnetic fields, or electromagnetic waves, propagated at the speed of light in a vacuum.

electromagnetic spectrum All possible frequencies of electromagnetic radiation shown in order of their wavelengths or frequencies.

electron A negatively charged, very light particle found in energy levels outside the nucleus; one of the three subatomic particles.

electron affinity The change in energy when an electron is added to a neutral atom in the gas phase to form a negative ion.

electron arrangement The distribution of pairs of electrons around an atom in a way that minimises repulsion.

electron configuration In the shell model of an atom, the electronic configuration is a means of representing the number of electrons in each shell.

electron configuration notation A written series of numbers and/or letters stating the number of electrons in each shell or subshell.

electron density The concentration of electrons around an atom or a molecule.

electron distribution The arrangement of electrons around an atom or a molecule.

electron dot diagram A representation of the electron arrangement in a molecule in which outer-shell electrons are represented by dots or crosses.

electron shell A region or energy level of the atom in which electrons can be found.

electron transfer diagram A diagram that shows how electrons move from a metal atom to a non-metal to form ions.

electronegative A measure of the ability of an atom of an element to attract electrons.

electronegativity The ability of an atom to attract electrons in a covalent bond towards itself.

electrostatic attraction The force of attraction between a positively charged particle and a negatively charged particle.

electrostatic force Force of attraction or repulsion between particles that are the result of the charge on each particle.

element (1) A component or constituent part of a complex whole; a fundamental, essential or irreducible part of a composite entity; a pure substance containing only one type of atom (same atomic number). (2) Elements cannot be broken down into simpler substances by physical or chemical means.

eluent The solvent that carries the components and passes through the chromatography column. It may be a liquid or a gas.

emission spectrum A spectrum produced when an element is excited by heat or radiation. It appears as distinct lines characteristic of the element.

empirical formula A formula that shows the simplest whole number ratio of the elements in a compound, e.g. CH_2 is the empirical formula of propene (C_3H_6).

endothermic A reaction that absorbs energy from the surroundings; ΔH is positive.

energy The capacity for doing work or generating heat.

energy density The energy released when 1 litre of fuel undergoes complete combustion.

energy level A region or electron shell of the atom in which electrons can be found.

energy profile diagram A diagram that shows the energy changes during the course of a reaction.

enhanced greenhouse effect The increase in the warming effect of the atmosphere due to changes to the increase in production of greenhouses gases by human activity.

enthalpy Heat content. The sum of the chemical, potential and kinetic energies in a substance. Given the symbol H.

enthalpy change The difference in the total enthalpy of the products and the total enthalpy of the reactants. Given the symbol ΔH.

enthalpy level diagram A diagram that shows the energy changes during the course of a reaction.

enzyme A protein molecule that functions as a catalyst of a specific biochemical reaction. A catalyst increases the rate of reaction.

equilibrium A state of balance in which the rate of a forward reaction or process is equal to the rate of its reverse reaction or process.

error A measure of the estimated difference between the observed or calculated value of a quantity and its true value.

error bar A line drawn through a point on a graph that is parallel to one of the axes and which represents the uncertainty or variability of the data of that point.

ester A molecule of the homologous series containing the (–COO–) functional group.

ethics Moral principles that govern a person’s behaviour or the conducting of an activity.

evaporation The change of physical state where a liquid becomes a gas.

excess reactant A reactant that is not completely consumed in a chemical reaction.

excited electron An electron that has absorbed a particular quantity of energy and moved from its ground state to a higher energy level. An atom is said to be excited when an electron is not in the lowest electron energy level possible (i.e. not in the ground state).

excited state A state where energy has been absorbed by an atom and electron(s) have moved to a higher energy level; energy will be released as it returns to its ground state.

exothermic An exothermic reaction releases energy to the surroundings; ΔH is negative.

experimental group The selected test group (sample group) in which the independent variable has changed to elicit a change in the measured results.

extend In science, to extend an experiment is to modify the methodology to overcome limitations of the scope or applicability of the data.

extrapolate Infer or estimate by extending or projecting known information; conjecture; infer from what is known; extend the application of something (e.g. a method or conclusion) to an unknown situation by assuming that existing trends will continue or similar methods will be applicable

F

fermentation The breakdown of sugar solutions, usually by yeast enzymes, producing ethanol and carbon dioxide.

first ionisation energy The energy required to remove one electron from an atom of an element in the gas phase.

flashpoint The lowest temperature at which the vapour of a liquid will ignite when in air and when near an ignition source.

fluoresce To shine or glow brightly.

formula unit Chemical formula that shows the overall ratio of positive and negative ions in an ionic compound.

fossil fuel Fuel produced over a long period of time by the breakdown of organic material. This includes coal, crude oil and natural gas.

fractional distillation Separation of components of a mixture based on their boiling point.

freezing The change of physical state where a liquid becomes a solid.

freezing point The temperature at which a substance is transformed from its liquid state to its solid state and where both states coexist in equilibrium.

freezing point depression The decrease in temperature between the freezing point of a solution and the freezing point of the pure solvent.

full equation A representation of a reaction that uses formulas.

fullerene A molecule composed entirely of carbon, in the form of a hollow sphere or tube. Other shapes are possible. Each carbon atom is bonded to three other carbon atoms.

G

gamma radiation Short wavelength, high energy radiation.

gas A substance which expands to fill any space available, irrespective of its quantity. A phase of matter.

gas chromatography (GC) A very sensitive form of chromatography in which the mobile phase is a gas that passes through a column containing particles (stationary phase).

gas chromatography–mass spectrometry is an analytical method that combines the features of gas chromatography and mass spectrometry to identify different substances within a test sample.

gas constant The constant, R, in the ideal gas equation. $R = 8.31\ \text{J K}^{-1}\ \text{mol}^{-1}$ when pressure is measured in kPa, volume in L and temperature in K.

gas–liquid chromatography (GLC) A form of GC where the stationary phase is a liquid on the surface of resin beads and the mobile phase is a gas.

gas–solid chromatography (GSC) A form of GC where the stationary phase is a solid and the mobile phase is a gas.

gaseous Having the characteristics of a gas; no fixed shape and takes on the shape of the space available.

gaseous phase A region of space in which a substance exists as a gas.

geosequestration The process by which gases such as carbon dioxide can be concentrated and stored underground.

giant molecule A covalently bonded substance with a large number of atoms that exists as a separate, distinct entity.

graph A diagram showing the relationship between variable quantities, usually two variables.

graphene A form of carbon consisting of planar sheets one atom thick in which each carbon atom is bonded to three neighbouring carbon atoms.

graphite A form of carbon in which the carbon atoms are arranged in layers.

green chemistry principles A set of 12 principles that aim to reduce the impact of chemistry on the environment.

greenhouse effect The production of liveable temperatures on Earth by the process absorption and redirection of infrared radiation by atmospheric gasses.

greenhouse gases Atmospheric gases such as carbon dioxide methane and water which have the ability to absorb and redirect infrared radiation and therefore contribute to the greenhouse effect.

ground state A term used to describe an atom in which the electrons occupy the lowest possible energy levels.

group (periodic table) A vertical column of elements in the periodic table.

H

haematite A mineral containing iron. Composed of iron oxides with the formula Fe_2O_3. Haematite is the main source of iron ore in Australia.

half-life The time taken for number of particles emitted from a radioactive source to be reduced by half.

halogen A group 7 element—F, Cl, Br, I, At, Ts.

heat The energy transfered from one system to another because of a difference in temperature.

heat content The enthalpy of a substance. Given the symbol H.

heat of combustion The energy released when specified quantities (e.g. 1 mole or 1 gram) of a substance burns completely in oxygen.

heat of reaction The exchange of heat between a system and its surroundings during a chemical reaction under constant pressure. Given the symbol ΔH, and also known as the enthalpy change.

heat treatment Heating a metal in different ways to alter its structure and physical properties.

Hess's law The amount of heat energy released or absorbed in a chemical reaction is constant, irrespective of the number of steps or the kind of steps by which the reaction is carried out, provided that the same reactants and products are involved.

heterogeneous Diverse, different. A heterogeneous substance or solution possesses two or more different types of phases in the one sample, e.g. a suspension.

heterogeneous catalyst A catalyst that has a different physical state (phase) from the reactants and products.

heterogeneous mixture A mixture consisting of two or more substances that have visibly distinguishable regions, called phases, which have different physical and chemical properties.

Higgs boson Particle that is the carrier of the Higgs field and endows subatomic particles with mass through its interactions with them.

high-performance, or high-pressure, liquid chromatography (HPLC) A sensitive form of chromatography in which the mobile phase is a liquid which passes through a column containing particles (stationary phase).

homogeneous Uniform. The components of a homogeneous substance are uniformly distributed throughout the substance, e.g. a solution is homogeneous because the solute and the solvent cannot be distinguished from each other.

homogeneous catalyst A catalyst that has the same physical state (phase) as the reactants and products.

homogeneous mixture A mixture consisting of two or more substances but only one visibly distinct phase that has uniform physical and chemical properties throughout. Also known as a *solution*.

Hubble Space Telescope An extremely high-resolution space telescope launched into low Earth orbit in 1990. Observations have led to breakthroughs in astrophysics, including the rate of expansion of the universe.

Hund's rule Describes how to arrange electrons when a *p*- or *d*-subshell is not full. It states that every orbital in a subshell is singly occupied before any orbital is doubly occupied.

hydrated An ion surrounded by water molecules. Hydrated ions can be found in aqueous solutions or crystalline solids.

hydride A compound in which hydrogen is bonded to another element. HF, HCl and HI are hydrides of group 17 elements.

hydrocarbon A compound that contains carbon and hydrogen only.

hydrogen bond A type of intermolecular, dipole–dipole force where a hydrogen atom is covalently bonded to a highly electronegative atom such as oxygen, nitrogen or fluorine. Due to the disparity of electronegativity values between the atoms involved, the hydrogen develops a partial positive charge and bonds to lone pairs of electrons on neighbouring atoms of oxygen, nitrogen or fluorine.

hydrogen bonding A particularly strong form of dipole–dipole force that only occurs between highly polar molecules in which a hydrogen atom is covalently bonded to an oxygen, a nitrogen or a fluorine atom.

hydronium ion The $H_3O^+(aq)$ ion.

hydroxide ion The $OH^-(aq)$ ion.

hyperbaric oxygen therapy (HBOT) A medical treatment in which patients breathe pure oxygen while inside a decompression chamber at a pressure higher than 1 atm.

hypothesise Formulate a supposition to account for known facts or observed occurrences; conjecture, theorise, speculate; especially on uncertain or tentative grounds.

I

ideal gas A gas that obeys the gas equations at all temperatures and pressures.

ideal gas equation The equation that describes the behaviour of an ideal gas: $PV = nRT$

incomplete combustion A combustion reaction occurring in conditions of limited oxygen supply, resulting in production of carbon and carbon monoxide as well as carbon dioxide.

independent variable The variable during an experiment that is deliberately altered, or selected to be tested, that is assumed to cause a change in the dependent variable.

indicator A substance that is different colours in its acid and base forms.

induced-fit model A model of enzymatic action which proposes that the active site of an enzyme changes shape to facilitate binding of the substrate molecule(s) and formation of the enzyme–substrate complex.

inductively coupled plasma (ICP) A highly sensitive ionisation technique that when used in conjunction with mass spectrometry is capable of detecting metals and non-metals concentrations as low as one part in 10^{15}.

industrial era The period after about 1760 during which many industrial and manufacturing processes using fossil fuels began.

infrared radiation Electromagnetic radiation between 7×10^2 and 1×10^6 nm wavelength.

initial rate of reaction The instantaneous rate of reaction at time, $t = 0$.

insoluble Incapable of being dissolved.

instantaneous dipole See *temporary dipole*.

instantaneous rate of reaction The rate of reaction at any time instant during the course of a chemical reaction.

intermolecular bond The net forces of attraction that act between neighbouring molecules.

intermolecular force An electrostatic force of attraction between molecules, including dipole–dipole forces, hydrogen bonds and dispersion forces.

International System (SI) A worldwide system used by scientists to measure physical properties using the metric system.

interstitial alloy An alloy made by adding smaller atoms to a metal.

intramolecular bond A force that holds the atoms within a molecule together.

ion A positively or negatively charged atom or group of atoms.

ion–dipole attraction The attraction that forms between dissociated ions and polar water molecules when an ionic solid dissolves in water.

ion–dipole interaction An electrostatic interaction between a charged ion of a molecule that has a dipole.

ionic bonding A type of chemical bonding that involves the electrostatic attraction between oppositely charged ions.

ionic compound A compound composed of positively charged ions (cations) and negatively charged ions (anions) arranged in a three-dimensional lattice. Ionic compounds are characterised by ionic bonds, i.e. the electrostatic attraction of ions of opposite charge.

ionic equation An equation for a reaction that only includes the ions that are involved in the reaction, e.g. $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)^2(s)$.

ionic radius A measurement used for the size of ions; the distance from the nucleus to the outermost electrons.

ionisation (i) The removal of one or more electrons from an atom or ion; (ii) the reaction of a molecular substance with a solvent to form ions in solution.

ionisation energy The energy required to remove one electron from an atom of an element in the gas phase.

ionise The reaction of a molecular substance with a solvent to form ions in solution. When some polar molecules dissolve in water they ionise to form a hydronium ion, e.g. $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$.

iron oxide A compound containing the elements iron and oxygen, e.g. Fe_2O_3, Fe_3O_4, FeO.

isoelectronic Two elements and/or ions have the same electronic configurations.

isotope Atoms of an element that have the same number of protons but a different number of neutrons.

isotopic abundance The percentage abundance of a particular isotope in a sample of an element.

isotopic mass The mass of an atom of the isotope relative to the mass of an atom of carbon-12 taken as 12 units exactly.

J

joule The SI unit for energy.

K

Kelvin scale The temperature scale in which the freezing point of water is given a value of 273 and the boiling point a value of 373.

kilopascal (kPa) A common unit of pressure.

kinetic energy The energy that a particle or body has due to its motion: $KE = \frac{1}{2}mv^2$

kinetic energy distribution diagram A graph of kinetic energy against number of particles that shows the range of energies in a sample of a gas or a liquid at a given temperature. Also known as a Maxwell–Boltzmann distribution curve.

kinetic theory A theory that explains the states of matter based on the concept all matter is composed of tiny particles that are in constant motion.

L

Large Hadron Collider A massive 27 km underground particle accelerator and proton collider operated by the European Organisation for Nuclear Research (CERN).

lattice A regular arrangement of large numbers of atoms, ions or molecules.

law A statement describing invariable relationships between phenomena in specified conditions, frequently expressed mathematically.

law of conservation of energy The total energy of a system remains constant; energy can neither be created or destroyed, rather, it transforms from one form to another.

law of conservation of mass A scientific law that states that in any physical change, or chemical reaction, matter can neither be created nor destroyed.

Lewis structure A representation of the bonding in a molecule or ion that uses lines to represent two electrons in a covalent bond.

limestone A mineral composed of calcium carbonate ($CaCO_3$).

limewater test A test for carbon dioxide gas. The presence of carbon dioxide is detected by bubbling the gas through a calcium hydroxide solution ($Ca(OH)_2(aq)$). The limewater reacts with the carbon dioxide and turns milky.

limiting reactant A reactant that is completely consumed in a reaction and which determines the amount of products formed.

line graph A graph that displays continuous data using a line from one point of measurement to the next.

line of best fit A straight line drawn through the centre of data points plotted on a scatterplot.

line spectrum See *emission spectrum*.

liquefied petroleum gas (LPG) A fuel comprising a mixture of propane and butane, which are liquefied under pressure.

liquid Matter that does not have a defined shape, and will take the shape of the container it is placed in. The particles in a liquid are in close contact with one another, but the arrangement is not fixed or orderly.

liquid phase A region of space in which a substance exists as a liquid.

literature review An evaluative report that critically analyses the current body of knowledge.

lock-and-key model A model of enzymatic action which proposes that the substrate and the active site of an enzyme have a specific complementary geometric shape that fit exactly into one another.

lone electron pair An outer-shell electron pair that does not form a bond with other atoms.

lone pair See *lone electron pair*.

lustrous Shiny or reflective when freshly cut or polished.

Lyman series A series of spectral lines in the ultraviolet part of the hydrogen spectrum that is produced by transitions of electrons between the ground state energy level and higher energy levels of the hydrogen atom

M

M Mass number. The number of protons and neutrons in the nucleus of an atom.

main group element An element in groups 1, 2 or 13–18 in the periodic table.

malleable Able to be bent or beaten into sheets.

mass A measure of an object's resistance to acceleration. It is related to the amount of matter that an object or particle contains. The SI unit for mass is the kilogram, kg, but the unit of gram, g, is more often used in chemistry studies.

mass number The total number of nucleons in an element's nucleus, A, i.e. the number of protons plus the number of neutrons.

mass spectrometer An instrument that measures the mass-to-charge ratio of particles.

mass spectrometry An analytical technique that uses the mass-to-charge, m/z, ratio of atoms, molecules and fragments of molecules to identify substances.

mass spectrum A plot of the isotopic mass, relative to the mass of carbon-12 taken as 12 units exactly, against the relative abundance of each isotope present in a sample.

matter A physical substance; anything that has mass and occupies space (Macquarie 1981)

Maxwell–Boltzmann distribution curve A graph of kinetic energy against number of particles that shows the range of energies in a sample of a gas or a liquid at a given temperature. Also known as a kinetic energy distribution diagram.

mean A calculated value that represents the centre of a set of numbers that have been averaged.

measured variable A variable that is measured during an experiment because it cannot be controlled, and must be known so that any influence on the measured results or dependent variable can be determined.

median The value of the middle number of a set of numbers arranged from smallest to highest number.

melt The change of physical state where a solid becomes a liquid.

melting point The equilibrium temperature of the solid and liquid phases of a substance.

metal ion Metal atom that loses electrons to non-metallic atoms and so becomes positively charged.

metallic bond The electrostatic attractive forces between delocalised valence electrons and positively charged metal ions.

metallic bonding model A description that explains the properties and behaviour of metals in terms of the particles in metals.

metallic character The ease with which its atoms lose electrons.

metalloid An element that displays both metallic and non-metallic properties, e.g. germanium, silicon, arsenic, tellurium.

methane CH_4; the main component of natural gas.

methodology A systematic, ordered approach to gathering data in a scientific experiment or investigation.

mineral A naturally occurring inorganic substance that is solid and can be represented by a chemical formula, e.g. quartz.

miscible Liquids that can be mixed in any ratio to form a homogeneous solution.

mixture A physical combination of two or more pure substances.

mobile phase The phase that moves through the stationary phase in a chromatographic separation.

mode The value that occurs most often in a list of values.

model In science, a representation that describes, simplifies, clarifies or provides an explanation of the workings, structure or relationships within an object, system or idea. (ACARA 2015c)

modify Change the form or qualities of; make partial or minor changes to something.

mol Unit of moles.

molar mass The mass of one mole of a substance measured in g mol^{-1}; symbol M.

molar volume The volume occupied by one mole of gas at a specified set of conditions. At standard temperature (273 K) and pressure (100 kPa), the molar volume of a gas is 22.4 L mol^{-1}.

molarity The amount of solute, in moles, dissolved in 1 litre of solution (mol L^{-1}).

mole The amount of substance that contains the same number of fundamental particles as there are atoms in 12 g of carbon-12; symbol n; unit mol.

mole ratio The ratio of species involved in a chemical reaction, based on the ratio of their coefficients in the reaction equation.

molecular compound A compound comprised of molecules.

molecular formula A formula of a compound that gives the actual number and type of atoms present in a molecule. It may be the same as or different from the empirical formula.

molecule A group of two or more atoms covalently bonded together, and representing the smallest fundamental unit of a chemical compound.

molten Materials that are normally found as solids but are liquid, melted, due to elevated temperature.

N

N The number of particles in a mole of a substance.

N_A Avogadro's number, $6.02 \times 10^{23} mol^{-1}$. The number of particles in a mole.

nanomaterial A material with nanoscale features.

nanoparticle A particle in the size range 1–100 nm.

nanorod A nanoscale rod in which length and width are in the range 1–100 nm. The length of the rod is 3–5 times its width.

nanoscale The scale used to classify objects 1–100 nm in size.

nanoscience The study of nanoparticles and nanotechnology.

nanotechnology Technologies manipulating structures sized 1–100 nanometres (10^{-9} to 10^{-7} m).

nanotube An allotrope of carbon that consists of layers of carbon atoms formed into a long cylinder.

nanowire A wire that has a diameter measured on the nanoscale. Its length is unrestricted.

natural gas A fossil fuel composed of hydrocarbons under the conditions that they are extracted from Earth. Natural gas consists mainly of methane.

net dipole Uneven distribution of charge over the molecule as a whole.

neutral solution A solution in which the concentrations of H_3O^+ ions equals the concentration of OH^- ions; is neither acidic nor basic. At 25°C, a neutral solution has a pH of 7.

neutralisation reaction An acid reacts with a base in stoichiometric proportions to form a salt plus water.

neutralise To react an acid with a base in stoichiometric proportions to form a solution of a salt and water.

neutron A neutral subatomic particle found in the nucleus of atoms.

noble gas An unreactive gaseous element in group 18 of the periodic table. With the exception of helium, noble gases have eight electrons in their outer shells.

noble gas notation See *condensed electron configuration.*

nominal variable Categorical variables based on a nominated category which is not dependent on order.

non-bonding electron An outer-shell electron that is not shared between atoms.

non-metal ion Non-metal atoms gain electrons from metal atoms and so become negatively charged.

non-polar Bonds or molecules that do not have a permanent dipole. They have an even distribution of charge.

non-polar solvent A liquid or solvent that is a compound of two or more elements whose electronegativities are almost the same, e.g. oil.

nuclear symbol notation The writing of an atom or isotope in the format ^A_ZM; where A = mass number and Z = atomic number.

nucleon A particle in the nucleus; a proton or neutron.

nucleus A dense structure in the centre of the atom comprising almost all the mass of the atom in a very small volume.

O

observation Witnessing an event or phenomenon.

octet arrangement Arrangement of electrons in atoms or ions in which there are eight electrons in the outer shell.

octet rule A rule used as part of the explanation for electron configuration and in bonding. The rule states that during chemical reactions, atoms tend to lose, gain or share electrons so that there are eight electrons in their valence shell.

orbital A three-dimensional region around a nucleus which can contain no more than two electrons.

orbital diagram A depiction of pairs of electrons as arrows with opposite directions in linked boxes; grouped boxes represent subshells.

ordinal variable Categorical variables in which the order of the nominated category is important.

ore A mineral or an aggregate of minerals that contains a valuable constituent, such as a metal, which is mined or extracted.

origin The position that samples are placed to be separated by thin layer or paper chromatography.

outermost shell See *valence shell.*

outlier A value that 'lies outside' (is much smaller or larger than) most of the other values in a set of data.

P

paper chromatography An analytical technique for separating and identifying mixtures which uses paper as the stationary phase.

particulate matter Particles of unburnt solid fuel.

parts per billion (ppb) A unit of concentration which states the number of grams of solute in 1 billion grams of solution. It is equivalent to the number of mg of solute per kg of solution.

parts per million (ppm) A unit of concentration which states the number of grams of solute in 1 million grams of solution. It is equivalent to the number of mg of solute per kg of solution.

Pascal A unit of pressure equal to one Newton per square metre.

Paschen series A series of spectral lines in the infrared part of the hydrogen spectrum that is produced by transitions of electrons between the $n = 3$ energy level and higher energy levels of the hydrogen atom

Pauli exclusion principle No two electrons can occupy the same position or state in an atom; this means that electrons in the same orbital must have opposite spin.

percentage composition The proportion by mass of the different elements in a compound. % by mass of an element in a compound = (mass of the element present/total mass of the compound) × 100.

percentage yield A measure of the quantity of a product obtained from a chemical process compared to the maximum amount possible if the reaction were complete, expressed as a percentage: percentage yield = actual yield / theoretical yield × 100/1.

period (periodic table) A horizontal row of elements in the periodic table. The start of a new period corresponds to the outer electron of that element beginning a new shell.

periodic table (of elements) An arrangement of the elements in order of increasing atomic number in which elements of similar chemical and physical properties are placed in vertical columns known as groups. More specifically, the groups are arranged according to similar outer-shell electron configurations.

periodic trend See *periodicity*.

periodicity The periodic pattern of properties of the elements.

personal protective equipment (PPE) Equipment specifically designed and used during activities to protect the user from hazards.

petrodiesel The most common form of diesel fuel produced from crude oil. It is composed of mostly alkanes ($C_{10}H_{22}$ to $C_{15}H_{32}$) with some aromatic hydrocarbons.

pH scale A measure of acidity and the concentration of hydronium ions, in solution. Acidic solutions have a pH value less than 7 at 25°C and bases have a pH value greater than 7 (at 25°C). Mathematically, pH is defined as $pH = -\log_{10}[H_3O^+]$.

phase A region of matter that is physically and chemically uniform in composition and properties. With respect to heterogeneous mixtures, a phase is physically and mechanically distinct from other phases.

phase of matter Three physical forms that matter can occur in; solid, liquid and gas. Also known as states of matter.

phenomena Events that are not artificial and can be observed through the senses or can be scientifically described or explained.

photochemical smog Atmospheric pollution produced through the action of sunlight on nitrogen oxides and unburned hydrocarbons to form ozone and other pollutants. The nitrogen oxides are formed in high temperature reactions such as those that occur in car engines and lightning strikes.

photosynthesis A reaction that occurs in the leaves of plants between carbon dioxide and water in the presence of sunlight and chlorophyll to form glucose and oxygen.

photovoltaic cell A device constructed from a specialised semiconductor that can produce a flow of electrons from light energy.

physical change A process where the form of matter may be changed without changing its chemical identity or its chemical composition. No new substances are formed during physical changes.

physical property A characteristic of a substance that can be observed or measured.

physical state Refers to the different physical forms that substances can take; most commonly as solids, liquids or gases.

pie chart A spherical display of either qualitative or quantitative data showing the proportion of each data set (measurement, independent variable, observation) as a part of the whole.

plagiarism Taking someone else's ideas or work and passing it off as your own.

plasma An ionised state of matter, similar to a gas.

polar Bonds or molecules with a permanent dipole. They have an uneven distribution of charge.

polar solvent A solvent whose molecules have a permanent dipole (uneven separation of charge).

polarisable Describes an atom or molecule whose electron cloud can be distorted to form a temporary dipole. The more polarisable the atom or molecule is, the more easily its electron cloud can be distorted and the more easily it will form a temporary dipole.

polarised Describes an atom or molecule in which the electron cloud has been distorted to form a temporary dipole.

polarity The measure of how polar a molecule or bond is. It is a measure of the unequal sharing of electrons and is the result of the difference in electronegativities of the bonded atoms.

polyatomic ion An ion that is made up of more than one element, e.g. the carbonate ion (CO_3^{2-}).

polyatomic molecule A molecule that consists of more than two types of atoms, e.g. H_2O.

porous Able to absorb liquids or allow liquids to pass through.

positron emission tomography (PET) A medical imaging technique detecting positrons emitted from short-lived radionucleotides in the body.

potential energy The energy that an object or a piece of matter has due to its position or composition.

precipitate The solid formed during a reaction in which two or more solutions are mixed.

precipitation reaction A reaction between substances in solution in which one of the products is insoluble.

precision Accuracy; exactness; exact observance of forms in conduct or actions.

In science, exactness; how close two or more measurements of the same object or phenomena are to each other

prefix A symbol written at the front of a unit to indicate that the measurement is a power of ten higher or lower than the unit itself.

pressure The force exerted per unit area over a surface.

primary data Data collected directly by a person or group. (ACARA 2015c)

primary source Information created by the person or persons directly involved in a study, investigation or experiment or observing an event. (ACARA 2015c)

principle The known cause of any effect, in a system where the variables and conditions are known.

processed data Manipulated data to produce meaningful information; a set of displayed data in an appropriate form such as tables or graphs; required information that has been extracted from a set.

product A substance produced in a reaction.

property A characteristic feature of a substance or material that can be used to identify or describe matter or to distinguish one type of substance from another. Includes both physical and chemical properties.

proton A positively charged subatomic particle with a mass of one found in the nucleus.

pure substance Matter that has a definite and distinct set of physical and chemical properties that does not change in chemical composition from sample to sample. Substances cannot be broken down into simpler components by physical means and are classified as either elements or compounds.

pyramidal A molecular shape in which a central atom sits at the apex of a pyramid and is bonded to three atoms on the triangular base.

Q

qualitative analysis An analysis to determine the identity of chemical(s) present in a mixture.

qualitative data Information that is not numerical in nature.

quanta Packets of energy released or absorbed by atoms as electrons move between energy levels.

quantitative analysis An analysis to determine the concentration of chemical present in a mixture.

quantitative data Numerical information. (Taylor 1982)

quantum mechanical model A mathematical description of the behaviour of electrons in atoms that links position of an electron to energy level.

quantum mechanics The science of the behaviour and properties of the subatomic particles of atoms.

quenching Heating a metal to a moderate temperature and then cooling it rapidly to make it harder and more brittle.

R

radioactive A nucleus with excess energy that must be relieved by emitting the excess energy.

radioactive dating A technique used to date materials such as rocks or carbon, based on the decay of radioactive isotopes.

radioactivity A spontaneous emission of energy or particles from a nucleus resulting in the degradation of the nucleus.

radioisotope An isotope of a chemical element that emits radioactivity due to its unstable combination of neutrons and protons in the nucleus.

random error Uncontrollable effects of the measurement equipment, procedure and environment on a measurement result; the magnitude of random error for a measurement result can be estimated by finding the spread of values around the average of independent, repeated measurements of the quantity. (ACARA 2015c)

range The variation in values between upper and lower limits on a particular scale or the largest and smallest measurements.

rate of reaction The change in concentration of a reactant or product over a period of time (usually one second): rate of reaction = change in concentration/time.

rationale A set of reasons, or logical basis for a course of action or decision.

raw data Unprocessed and/or unanalysed data; data that has been collected without any additional processing. (Taylor 1982)

reactant A substance present at the start of a reaction.

reaction pathway A series of chemical reactions that converts a starting material into a product in a number of steps.

reactivity The ease with which a chemical can undergo reactions.

real gas Gases that do not obey the gas laws.

redirect In science, to redirect an experiment is to modify the methodology to gain further insight into the phenomena observed in the original experiment.

redshifted The phenomenon that is observed when light or other electromagnetic radiation from an object is increased in wavelength, or shifted to the red end of the spectrum. It occurs when the object emitting the light moves away from the observer.

reference list A list of resources that were cited, or in-text referenced in the document.

refine In science, to refine an experiment is to modify the methodology to obtain more accurate or precise data.

regression The calculated estimate of the strength of a relationship between variables.

relative atomic mass The weighted average of the relative isotopic masses of an element on the scale where ^{12}C is 12 units exactly; symbol Ar.

relative formula mass The mass of a formula unit relative to the mass of an atom of ^{12}C taken as 12 units exactly. It is numerically equal to the sum of the relative atomic masses of the atoms making up the formula. Substances that contain atoms or ions bonded in lattice structures have a relative formula mass. Such compounds include ionic compounds and covalent network substances.

relative isotopic abundance A measure of how much (in percentage) a particular isotope appears in an element.

relative molecular mass The mass of a molecule relative to the mass of an atom of ^{12}C, taken as 12 units exactly; symbol M_r.

relative uncertainty The measured uncertainty displayed as a percentage of the measurement.

reliability In science, the likelihood that another experimenter will obtain the same results (or very similar results) if they perform exactly the same experiment under the same conditions. (ACARA 2015c, Taylor 1982)

renewable energy Energy from a source that can be replenished at the same rate as being used.

repeat trial Repeatedly conducting a process, activity or trial in an attempt to replicate the results.

replication Repeatedly conducting or choosing a process, activity or sample attempting to replicate the results, event or set-up.

research question A question that directs the scientific inquiry activity; it focuses the research investigation or student experiment, informing the direction of the research, and guiding all stages of inquiry, analysis, interpretation and evaluation.

retardation factor (R_f) The ratio of the distance moved by a component to the distance the solvent has moved from the origin.

retention time (R_t) The time taken for a component to pass through a chromatography column.

risk assessment Evaluations performed to identify, assess and control hazards in a systematic way that is consistent, relevant and applicable to all school activities; requirements for risk assessments related to particular activities will be determined by jurisdictions, schools or teachers as appropriate. (ACARA 2015c)

S

safety data sheet (SDS) A document that outlines specific details about items that are required to be considered when handling the item for experimentation.

salt A substance formed from a metal or ammonium cation and an anion. Salts are the products of reactions between acids and bases, metal oxides, carbonates and reactive metals.

sample size The number of individual measurements or observations undertaken for each tested variable.

saturated solution A solution that cannot dissolve any more solute at the given temperature.

scanning tunnelling microscope (STM) A precision instrument able to measure small changes in electron flow from a surface; the shape of an atom is recorded by a scanning tip.

scatterplot A graph which displays the measured values for two variables plotted along two axes.

scientific journal A book containing the entire collection of work related to an experiment or investigation.

scientific method An orderly process of determining the relatedness between phenomena that were originally witnessed causing a query to develop.

scientific model The generation of a physical or conceptual representation of real phenomena or systems.

scientific notation A measured value written as a mathematical expression, between 1 and 10 including decimals displaying all significant figures, multiplied by ten raised to a specific power.

scientific report A document that follows the report genre, systemically organising the main information and findings of scientific research or work.

secondary data Data collected by a person or group other than the person or group using the data. (ACARA 2015c, Macquarie 1981)

secondary source A source of information or data which was not created or measured by its author, rather cited from another.

seed crystal A small crystal from which a large crystal of the same material can typically be grown.

self-ionisation An ionisation reaction of pure water in which water behaves as both an acid and a base.

shielding effect The repulsion between the inner-shell electrons and the valence shell electrons that reduces the effective nuclear charge experienced by the valence electrons.

sieving A physical separation process that sorts matter according to the size of particles.

significant figures The use of place value to represent a measurement result accurately and precisely. (ACARA 2015c)

single covalent bond A covalent bond in which two electrons are shared between two nuclei. It is depicted in a valence structure as a line between the two atoms involved.

slag A mixture of waste materials left over after a desired metal has been removed from its ore.

solid Matter that has a definite shape and volume; the shape does not depend on the shape of the container. The particles in a solid are packed tightly together, often in an orderly arrangement.

solubility A measure of the amount of solute dissolved in a given amount of solvent at a given temperature.

solubility curve A graph of solubility versus temperature for a particular solute dissolved in a particular solvent.

solubility table A reference table that can be used to predict the solubility of ionic compounds.

soluble Capable of being dissolved or liquefied.

solute A substance that dissolves in a solvent, e.g. sugar is the solute when it dissolves in water.

solution A homogeneous mixture of a solute dissolved in a solvent.

solution calorimeter An insulated container that holds a known volume of water and in which a reaction in solution, such as dissolution of a solid or a neutralisation reaction, can be carried out.

solvent A substance, usually a liquid, which is able to dissolve a solute to form a solution. Water is a very good solvent.

solvent front The wet edge or border of the solvent as it moves through the paper or plate in chromatography.

space-filling model A three-dimensional model in which the atoms are represented by spheres with radii proportional to the radii of the atoms involved. The distances between spheres are also proportional to the distances between the atomic nuclei.

specific heat capacity The amount of thermal energy transfer necessary to raise the temperature of one kilogram of a substance by one degree.

spectator ion An ion that remains in solution and is unchanged in the course of a reaction. Spectator ions are not included in ionic equations.

spectroscopic technique A quantitative technique that analyses the way that radiation, such as light and radio waves, interacts with matter.

spectroscopy Study of the absorption and emission of light and other radiation.

spiking A procedure used to tentatively identify a component within a mixture by high-performance liquid chromatography. If a chemical is thought to be one of the components of a mixture, it can be added to the sample. If there are no extra peaks in the chromatogram, just a larger peak for one of the components, it is assumed that the added chemical is one of the components of the mixture.

standard Chemical of known composition used for comparison.

standard atmosphere A unit of pressure equal to the pressure required to support 760 millimetres of mercury (760 mmHg) in a mercury barometer at 25°C. This is the average atmospheric pressure at sea level.

standard deviation A calculated spread of data which estimates the quartile boundaries from the mean, assuming a normalised (bell) curve.

standard solution A solution that has an accurately known concentration.

standard temperature and pressure (STP) Conditions of temperature and pressure relevant to a gas, where the temperature is 0°C (273 K) and pressure is 100 kPa.

state of matter Three physical forms that matter can occur in; solid, liquid, and gas. Also known as phases of matter.

state symbol Symbols used in chemical equations to indicate the state of a reactant or product chemical.

stationary phase A solid that selectively adsorbs components of a mixture as they are carried along by the mobile phase. In GC and HPLC, it is small beads of material that may have a viscous liquid attached.

steel A generally hard, strong, durable and not malleable alloy of iron and carbon, usually containing 0.2–1.5% carbon, often with other constituents such as manganese, chromium, nickel, molybdenum, copper, tungsten, cobalt or silicon, depending on the desired alloy properties.

stoichiometry The calculation of relative amounts of reactants and products in a chemical reaction. Chemical equations give the ratios of the amounts (moles) of the reactants and products.

strong acid An acid that readily donates a hydrogen ion (proton) to a base.

subatomic particle A particle inside the atom; in chemistry, protons, neutrons and electrons.

sublimation The process by which a substance goes directly from the solid phase to the gaseous phase, without passing through a liquid phase.

subshell A part of a shell (energy level) containing orbitals; subshells are named *s*, *p*, *d*, *f* or *g*.

substitutional alloy An alloy made from elements of similar chemical properties and size.

substrate A reactant in a reaction that is catalysed by an enzyme.

supersaturated solution An unstable solution that has more solute dissolved at a given temperature than a saturated solution.

surface area The area of all surfaces of the substance that is exposed to the other reactants. This is proportional to the amount of particles available at the surface to react.

surface tension The resistance of a liquid to increase its surface area.

surroundings The rest of the universe around a chemical reaction.

sustainable Able to support energy and resources into the future without depletion.

symmetrical molecule A molecule in which the polar bonds are evenly (or symmetrically) distributed. The bond dipoles cancel out and no net dipole is formed.

system A group of interacting objects, materials or processes that form an integrated whole; systems can be open or closed. (ACARA 2015c)

systematic error An error that is affected by the accuracy of a measurement process that causes readings to deviate from the accepted value by a consistent amount each time a measurement is made. (Taylor 1982)

T

temperature A measure of the average kinetic energy of the particles of an object.

tempering A process in which a metal that has been quenched is warmed again to a lower temperature to reduce its brittleness but to retain its hardness.

temporary dipole A dipole that forms due to random movement of the electrons surrounding the molecule.

tensile strength The maximum resistance of a material to a force which is pulling it apart before breaking, measured as the maximum stress the material can withstand without tearing.

tetrahedral A molecular shape in which a central atom is bonded to four atoms that sit at the corners of a tetrahedron.

tetrahedral shape The shape of a molecule with a central atom surrounded by four other atoms. The bond angle between two outer atoms and the central atom is 109.5°.

theoretical yield The mass of product that would be formed if the limiting reagent reacted fully. The mass that is expected for a particular product in a chemical reaction, based on the mass of limiting reactant used.

theory In science, a set of concepts, claims and/or laws that can be used to explain and predict a wide range of related observed or observable phenomena; theories are typically founded on clearly identified assumptions, are testable, produce reproducible results and have explanatory power. (ACARA 2015c)

thermochemical equation A chemical equation that includes the enthalpy change of the reaction, ΔH_c.

thermochemistry The study of energy changes that occur during chemical reactions.

thin-layer chromatography An analytical technique that uses a thin layer (such as silica) spread on a glass or plastic plate to separate a mixture.

three-dimensional lattice Cations and anions combine to form a regular arrangement, which has depth, height and breadth, of large numbers of atoms, ions or molecules.

trace contaminants Unwanted components of a mixture present in very small quantities.

transesterification A process in which a triglyceride reacts with an alcohol to form esters.

transition metal An element in groups 3–12 in the periodic table.

transition state An arrangement of atoms in a reaction that occurs when sufficient energy is absorbed for the activation energy to be reached. It represents the stage of maximum potential energy in the reaction. Bond breaking and bond forming are both occurring at this stage, and the arrangement of atoms is unstable.

transportation Using nanoparticles to carry chemicals through air, skin and cells.

trend line A line drawn onto a graph to accurately display a relationship between the variables

trigonal planar A molecular shape in which a central atom is bonded to three atoms that form a triangle in one plane.

triple covalent bond A covalent bond in which six electrons are shared between two nuclei. It is depicted in a valence structure as three lines between the two atoms involved.

U

uncertainty Range of values for a measurement result, taking account of the likely values that could be attributed to the measurement result given the measurement equipment, procedure and environment (ACARA 2015c); indicators of uncertainty may include percentage and/or absolute measurement uncertainty, confidence intervals, inferential statistics, statistical measures of spread, e.g. range, standard deviation.

unsaturated solution A solution that contains less solute dissolved at a given temperature than a saturated solution.

V

valence electron An electron found in the valence shell; an outermost electron in an atom or ion.

valence shell The outermost shell of an atom that contains electrons.

valence shell electron pair repulsion (VSEPR) theory A model used to predict the shapes of molecules by assuming that the valence electron pairs surrounding an atom repel each other, and therefore adopt an arrangement that minimises this repulsion. The shape of the molecule is then determined by considering the positions of the bonded atoms.

validity In science, the extent to which tests measure what was intended; the extent to which data, inferences and actions produced from tests and other processes are accurate. (ACARA 2015c)

vaporisation Conversion of a solid or a liquid into a gas. Evaporation and boiling are both types of vaporisation.

vapour The gaseous state of a substance.

vapour pressure The pressure exerted by a vapour in equilibrium with its liquid in a closed container; vapour pressure is temperature dependent.

variable

adjective

Apt or liable to vary or change; changeable; inconsistent; (readily) susceptible or capable of variation; fluctuating, uncertain;

noun

In science, a factor that can be changed, kept the same or measured in an investigation, e.g. time, distance, light, temperature.

viscosity A measure of a liquid's resistance to flow.

volatile Can change from liquid to a gas (vaporise) easily.

volatility A measure of how readily a substance will vaporise by going from its liquid state to its gaseous state.

volume The physical space occupied by matter. The SI unit for volume is the cubic metre (m^3) but units of litres (L) and millilitres (mL) are commonly used in chemistry.

W

weak acid An acid that is partly ionised in water.

weak base A base that accepts hydrogen ions (protons) from acids to a limited extent.

wedge-dash notation A three-dimensional representation of a molecule in which lines are used to represent bonds in the plane of the page; wedges are used to represent bonds that extend out of the page towards the viewer and dashes are used to represent bonds that extend into the page and away from the viewer.

weighted average A weighted average takes into consideration the proportion of the entire amount that each substance provides, rather than assuming each substance contributes equally to the entire amount.

work hardening The increase in strength that results from processes such as hammering, rolling and drawing a metal.

Index

Attributions

We thank the following for their contributions to our textbook:

Cover: Denys Joannes, Wollemi Gems

Queensland Curriculum & Assessment Authority (QCAA): Chemistry 2019 v1.3 General Senior Syllabus © Queensland Curriculum & Assessment Authority. This syllabus forms part of a new senior assessment and tertiary entrance system in Queensland. Along with other senior syllabuses, it is still being refined in preparation for implementation in schools from 2019. For the most current syllabus versions and curriculum information, please refer to the QCAA website https://www.qcaa.qld.edu.au

123RF: alicenerr, p. 155bl; alila, p. 278; Elnur Amikishiyev, p. 418a; Designua, p. 373; Asaf Eliason, p. 149c, p. 149t; Emirali Koka, p. 152t; Anton Lebedev, p. 458; Cseh Loan, p. 151bl; Gino Santa Maria, p. 147l; Peterhermesfurian, p. 194r; fsstando, p. 152br; Stillfx, p. 151br; Ron Sumners, p. 154tl; Kim Takhyzh-Sviridov, p. 473; Aleksandrs Tihonovs, p. 147, p. 154br; George Tsartsianidis, p. 487; tykhyi, p. 154tr; Tatyana Vyc, p. 537.

Australian Institute of Bioengineering and Nanotechnology, The University of Queensland: pp. 35l, 35r, 36t.

Alamy Stock Photo: age fotostock, p. 69–7; Archive Pics, p. 69–2, p. 69–5; Robert Brook, p. 253; FLPA, p. 525; GL Archive, p. 69–4; Granger Historical Picture Archive, p. 66; Stephanie Jackson, Photographsofaustralia.com, p. 264; Sciencephotos, p. 504; Science History Images, p. 69–6; Sueddeutsche Zeitung Photo, p. 566; World History Archive, p. 266.

Berrybank Farm: Jock Charles/Berrybank Farm, p. 312.

Bridgeman Images: Stained glass of St. George, Southern German, 1400-10/State Hermitage Museum, St. Petersburg, Russia, p. 159t.

Bulimba Creek Catchment Coordinating Committee: Reproduced with permission, p. 528.

California Institute of Technology: Dr. Chi Ma, p. 31tr.

Chemistry Education Association: p. 551.

Chillmaid, Martyn F.: p. 515.

Clifford, Trevor: pp. 266, 271.

Commonwealth Department of Environment and Energy: p. 300.

CSIRO: scienceimage, p. 213.

DK Images: Mike Dunning, p. 455; Howard Shooter, p. 536, Ian O'Leary, p. 265.

Elsevier/Science Direct: p.109, Reprinted from Atmospheric Environment, Volume 41, Issue 10, N.S.Holmes, A review of particle formation events and growth in the atmosphere in the various environments and discussion of mechanistic implications, pp. 2183–2201, 2007, with permission from Elsevier

Fairfax Photo Sales: John Donegan/The Age, p. 312.

Fotolia: Altocumulus, p. 152bl; ASK-Fotografie, p. 475; ASP Inc, p. 490; Barabas Attila, p. 233; Harald Biebel, p. 154bl; BillionPhotos.com, p. 149b; bit24, p. 224a; dabjola, p. 362; Jeffrey Daly, p. 159b; eyeidea, p. 323tl; gmstockstudio, p. 303; Georgios Kollidas, p. 426; lexaarts, p. 162t; Lonely, p. 238; Jan Matoska, p. 516; Mirco251280, p. 140t; osmar01, p. 162b; Pabijan, p. 229; Nicholas Piccillo, p. 229tr; piotr_roae, p. 188t; pixdesign123, p. 464; psdesign1, p. 318r; Rrekemp, p. 420; sehbaer_nrw, p. 362; Tatiana Shepeleva, p. 218c; Angel Simon, p. 404; shaiith, p. 418b; Smileus, p. 115c; Teteline, p. 436; Ina van Hateren, p. 462; vermontalm, p. 428; Monika Wisniewska, p. 161tr.

Getty Images: 145/William Andrew, p. 348, David Butow, p. 524; John Cancalosi, p. 155br; Lowell Georgia/Corbis Documentary, pp. 212, 279; David Munden/Popperfoto, p. 40; John Sanford, p. 127t; Science Photo Library, p. 475; Andrew Unangst, p. 211.

HL Studios: p. 268.

International Business Machines Corporation: © International Business Machines Corporation, p. 190t.

Images of Elements: images-of-elements.com, p. 140c.

John Oxley Library, State Library of Queensland: 'Homeward Laddie', ca. 1895. Anthony Alder, 1838–1915. Oil Painting, 0.01 linear metres; 109 x 135 cm. Ref 28082, Heritage Collections. Licensed under Creative Commons CC-BY http://creativecommons.org/licenses/by/2.5/au, p. 141l.

Dan Lindsay: © 2009 Porsche Ceramic Composite Brakes on Porsche Carrera S (997)., p. 187t.

Lochman Transparencies: p. 475.

McBroom, Alice: p. 149b.

NASA Images: NSF/NRAO/VLA/NASA/courtesy of nasaimages.org, p. 130.

NRM North: p. 527.

Oxford Designers & Illustrators Ltd: pp. 118t, 267, 268.

Powerhouse Museum: Michael Myers, p. 187c.

Purdue Research Foundation: George Bodner, Purdue University/Bodner Group, p. 189.

Sanders, Robert: Courtesy of Robert Sanders, p. 123.

SA Water: Courtesy of the Australian Water Quality Centre, p. 526l.

Science Photo Library SPL: pp. 17b, 19r, 19c, 104–105, 117t, 129br, 244, 389, 564, 567; Dr Juerg Alean, p. 127c; Steve Allen, p. 536; Maurizio De Angelis, pp. 346–347; Andrew Lambert Photography, pp. 505, 509, 510, 517; Animate4.com, p. 39; Samuel Ashfield, p. 403; Stephen Ausmus, p. 282b; Dr Timothy Baker/Visuals Unlimited, Inc., p. 30d; J. Bernholc et al/North Carolina State University, p. 214l; Bill Beatty, Visuals Unlimited, p. 473a; British Antarctic Survey, p. 418c; CC Studio, pp. 54, 549; Centre Jean Perrin, ISM, p. 53; Martyn F. Chillmaid, pp. 19l, 230, 274; Carlos Clarivan, pp. 107, 450; Clive Freeman, Biosym Technologies, p. 371; Custom Medical Stock Photo, p. 138l; Colin Cuthbert, p. 472; Jeff Daly, Visuals Unlimited, pp. 32; Stefan Diller, p. 150b; Dorling Kindersley, p. 11; Peter Eaton-Requimte/University of Porto, Visuals Unlimited, p. 30c; Emilio Segre Visual Archives, p. 567; Eye of Science, p. 34; Clive Freeman/Biosym Technologies, pp. 345, 558a; Nicolle R. Fuller, p. 36b; Pascal Goetgheluck, p. 311; Adam Hart-Davis, p. 537; Gary Hincks, p. 522; Mikkel Jensen, p. 269; Kallista Images/Custom Medical Stock Photo, p. 188c; Russell Kightley, pp. 31b, 218l; Mehau Kulyk, p. 1; Laguna Design, pp. 186c, 558b; Library of Congress/Science Photo Library, p. 426b; Claus Lunau, p. 449; Massimo Breda, The Lighthouse, p. 397; David McCarthy, p. 161tl; Medi-mation, p. 31br; B. Murton/Southampton Oceanography Centre, p. 478; National Optical Astronomy Observatories, p. 127b; Susumu Nishinaga, p. 30a; Northwestern University, p. 30e; Ornl, p. 388; Power and Syred, pp. 30b, 400; Public Health England, pp. 55b, 55c; Ramon Andrade 3Dciencia, p. 214; Paul Rapson, p. 124; J.C. Revy, ISM, pp. 167–168, 477; Alexis Rosenfeld, p. 433b; Science Picture Co, pp. 160c, 218r; TEK Image, pp. 211, 298–299; Sheila Terry, p. 426a; Tina Carvalho, Visuals Unlimited, p. 18tl; Geoff Tompkinson, p. 387; Victor Habbick Visions, p. 213; Peidong Yang/UC Berkeley, p. 160b.

Shutterstock: akiyoko, pp. 534–535; Anneka, pp. 151c, 473b; Egorov Artem, p. 195; Art_of_sun, pp. 26–27; atdigit, p. 453; Marijus Auruskevicius, p. 186b; max blain, p. 323tr; Nadezda Boltaca, p. 155tr; Steve Bower, p. 377; Boykung, p. 141b; Natalia Bratslavsky, p. 160t; Rob Byron, p. 282a; Chromatos, p. 370; Dado Photos, p. 150; Comaniciu Dan, p. 5lc; Antonio Danna, p. 244a; Dainis Derics, p. 98tr; dio5050, p. 366; Designua, p. 563; Everett Historical, p. 69–3; Featureflash, p. 147c; Flegere, p. 5r; Fly_dragonfly, p. 464; Frederic Legrand – COMEO, p. 217; Goodluz, p. 266; Jason and Bonnie Grower, p. 264; Antonio Guillem, p. 14tr; hanabiyori, p. 244b; Brent Hofacker, p. 5l; Ibreakstock, p. 368; indianstockimages, p. 140b; irin-k, p. 214r; Isak55, pp. 74–75; ituchka, p. 5rc; Atiwan Janprom, p. 14br; Szasz-Fabian Jozsef, p. 149c; marcin jucha, pp. 222–223, 224b;

Georgios Kollidas, p. 69–1; Gerard Koudenburg, p. 21l; ktsdesign, p. 18c; Dominique Landau, p. 141r; JP Lavoie, p. 188b; Viacheslav Lopatin, p. 138c; Albert Lozano, p. 55a; Michal Ludwiczak, p. 503; Lumena, p. 1; Magnetix, pp. 67, 353, 371, 508; Alexandr Makarov, p. 156t; Marekusz, p. 14cr; maxstockphoto, p. 417; mdmworks, p. 214; molekuul.be, pp. 208, 368, 561; Vincent Noel, p. 14bl; Nuttapong, p. 229b; Jarun Ontakrai, p. 18r; OPgrapher, p. 523; Alon Othnay, p. 186t; Pakhnyushcha, pp. 31a, 318l; siamionau pavel, p. 229tl; Panom Pensawang, p. 323bl; Ivo Petkov, p. 268; Igor Petrushenko, p. 370; PomInOz, p. 309; Renamarie, p. 17t; Fouad A. Saad, pp. 287, 372, 431; Sciencepics, p. 362; Serger_ Bogomyako, p. 263; zhang sheng p. 137; Sl_photo, pp. 447–448; Yury Smelov, p. 168b; Carolina K. Smith, p. 37r; sspopov, p. 318b; Stable, p. 229; steveball, p. 524; TebNad, p. 323br; Mary Terriberry, p. 523; tonyz20, p. 31tl; Tooykrub, p. 485; Andris Torms, p. 362; JR Trice, p. 98tl; udaix, p. 20; Belovodchenko Vadim, p. 316; Martin Valigursky, p. 42; Joao Virissimo, p. 265; Wallpaper photographer, p. 477; Zelenskaya, p. 14tl; Jiri Vaclavek, p.194l.

TarraWarra Estate: p. 487.

University of Idaho Coeur d' Alene: Marie Pengilly, p. 526r.

University of Munich: p. 119b.

University of NSW: Centre for Quantum Computation and Communication Technology, p. 190.

Volvo Car Corporation: p. 216.

WolframAlpha: p. 129l.

Pearson Australia
(a division of Pearson Australia Group Pty Ltd)
707 Collins Street, Melbourne, Victoria 3008
PO Box 23360, Melbourne, Victoria 8012
www.pearson.com.au

First published 2019 by Pearson Australia

eBook credits
Resource author: Nick Holmes et al
Publisher: Alicia Brown
Project Manager: Michelle Thomas
Production Editor: Virginia O'Brien
Development Editor: Antonietta Anello
Editors: Marcia Bascombe, iEnergizer Aptara Ltd
Copyright & Pictures Editor: Lisa Woodland
Illustrator: Diacritech
eBook Production Coordinator: Louise Grundy
Digital Administrator: Jenny Smirnov

ISBN 978 1 4886 1961 8

Pearson Australia Group Pty Ltd ABN 40 004 245 943

eBook acknowledgements
We would like to thank the following for permission to reproduce copyright material.

The following abbreviations are used in this list: t = top, b = bottom, l = left, r = right, c = *centre*.

Cover: Denys Joannes, Wollemi Gems
Queensland Curriculum & Assessment Authority (QCAA): Chemistry 2019 v1.2 General Senior Syllabus © Queensland Curriculum & Assessment Authority. This syllabus forms part of a new senior assessment and tertiary entrance system in Queensland. Along with other senior syllabuses, it is still being refined in preparation for implementation in schools from 2019. For the most current syllabus versions and curriculum information please refer to the QCAA website https://www.qcaa.qld.edu.au
123RF: alicenerr, p. 155bl; alila, p. 278; Elnur Amikishiyev, p. 418a; Designua, p. 373; Asaf Eliason, p. 149c, p. 149t; Emirali Koka, p. 152t; Anton Lebedev, p. 458; Cseh Loan, p. 151bl; Gino Santa Maria, p. 147l; Ivan Mikhaylov, p. 26; Peterhermesfurian, p. 194r; Thomas Photiou, p. 36t; sstando, p. 152br; Stillfx, p. 151br; Ron Sumners, p. 154tl; Kim Takhyzh-Sviridov, p. 473; Aleksandrs Tihonovs, p. 147, p. 154br; George Tsartsianidis, p. 487; tykhyi, p. 154tr; Tatyana Vyc, p. 537.
Australian Institute of Bioengineering and Nanotechnology, The University of Queensland: pp. 35l, 35r, 36t.
Alamy Stock Photo: age fotostock, p. 69-7; Archive Pics, p. 69-2, p. 69-5; Robert Brook, p. 253; Ray Evans, p. 70; FLPA, p. 525; GL Archive, p. 69-4; Granger Historical Picture Archive, p. 66; ibreakmedia, p. 6; Stephanie Jackson, Photographsofaustralia.com, p. 264; Prisma by Dukas Presseagentur GmbH, p. 2; Sciencephotos, p. 504; Science History Images, p. 69-6; Stephen Barnes/Business, p. 110; Sueddeutsche Zeitung Photo, p. 566; World History Archive, p. 266.
Australian Government Department of Education and Training: p. 9.
Berrybank Farm: Jock Charles/Berrybank Farm, p. 312.
Bridgeman Images: Stained glass of St. George, Southern German, 1400-10/State Hermitage Museum, St. Petersburg, Russia, p. 159t.
Bulimba Creek Catchment Coordinating Committee: Reproduced with permission, p. 528.
California Institute of Technology: Dr. Chi Ma, p. 31tr.
Chemistry Education Association: p. 551.
Chillmaid, Martyn F.: p. 515.
Clifford, Trevor: pp. 266, 271.
Commonwealth Department of Environment and Energy: p. 300.
CSIRO: scienceimage, p. 213.
DK Images: Mike Dunning, p. 455; Howard Shooter, p. 536, Ian O'Leary, p. 265.
University of Munich: p. 119b.
Fairfax Photo Sales: John Donegan/The Age, p. 312.
Fotolia: Altocumulus, p. 152bl; ASK-Fotografie, p. 475; ASP Inc, p. 490; Barabas Attila, p. 233; Harald Biebel, p. 154bl; BillionPhotos.com, p. 149b; bit24, p. 224a; dabjola, p. 362; Jeffrey Daly, p. 159b; eyeidea, p. 323tl; gmstockstudio, p. 303; KAR, p. 67br; Georgios Kollidas, p. 426; lexaarts, p. 162t; Lonely, p. 238; Jan Matoska, p. 516; Mirco251280, p. 140t; osmar01, p. 162b; Pabijan, p. 229; Nicholas Piccillo, p. 229tr; piotr_roae, p. 188t; pixdesign123, p. 464; psdesign1, p. 318r; Rrekemp, p. 420; sehbaer_nrw, p. 362; Tatiana Shepeleva, p. 218c; Angel Simon, p. 404; shaiith, p. 418b; Smileus, p. 115c; Teteline, p. 436; Ina van Hateren, p. 462; vermontalm, p. 428;Sophia Winters, p. 67r; Monika Wisniewska, p. 161tr.
Getty Images: 145/William Andrew, p. 348, David Butow, p. 524; John Cancalosi, p. 155br; Lowell Georgia/Corbis Documentary, pp. 212, 279; David Munden/Popperfoto, p. 40; John Sanford, p. 127t; Science Photo Library, p. 475; Andrew Unangst, p. 211.
HL Studios: p. 268.
International Business Machines Corporation: © International Business Machines Corporation, p. 190t.
Images of Elements: images-of-elements.com, p. 140c.
Ixom: Ixom Operations Pty Ltd, p. 73.
John Oxley Library, State Library of Queensland: "Homeward Laddie", ca. 1895. Anthony Alder, 1838-1915. Oil Painting, 0.01 linear metres; 109 x 135 cm. Ref 28082, Heritage Collections. Licensed under Creative Commons CC-BY http://creativecommons.org/licenses/by/2.5/au, p. 141l.
Dan Lindsay: © 2009 Porsche Ceramic Composite Brakes on Porsche Carrera S (997)., p. 187t.
Lochman Transparencies: p. 475.
McBroom, Alice: p. 149b.
NASA Images: NSF/NRAO/VLA/NASA/courtesy of nasaimages.org, p. 130.
NRM North: p. 527.
Oxford Designers & Illustrators Ltd: pp. 118t, 267, 268.
PASCO Scientific: pp. 65t, 68, 69a, 69b, 69c, 69d, 79tr.

Powerhouse Museum: Michael Myers, p. 187c.
Purdue Research Foundation: George Bodner, Purdue University/Bodner Group, p. 189.
Sanders, Robert: Courtesy of Robert Sanders, p. 123.
SA Water: Courtesy of the Australian Water Quality Centre, p. 526l.
Science Photo Library SPL: pp. 17b, 19r, 19c, 104-105, 117t, 244, 389, 564, 567; Dr Juerg Alean, p. 127c; Steve Allen, p. 536; Amelie-Benoist/BSIP, p. 69t; Maurizio De Angelis, pp. 346-347; Andrew Lambert Photography, pp. 505, 509, 510, 517; Animate4.com, p. 39; Samuel Ashfield, p. 403; Stephen Ausmus, p. 282b; Dr Timothy Baker/Visuals Unlimited, Inc., p. 30d; J. Bernholc et al/North Carolina State University, p. 214l; Bill Beatty, Visuals Unlimited, p. 473a; British Antarctic Survey, p. 418c; CC Studio, pp. 54, 549; Centre Jean Perrin, ISM, p. 53; Martyn F. Chillmaid, pp. 19l, 230, 274; Carlos Clarivan, pp. 107, 450; Clive Freeman, Biosym Technologies, p. 371; Custom Medical Stock Photo, p. 138l; Colin Cuthbert, p. 472; Jeff Daly, Visuals Unlimited, pp. 32, 194l; Stefan Diller, p.150b; Dorling Kindersley, p. 11; Peter Eaton-Requimte/University of Porto, Visuals Unlimited, p. 30c; Emilio Segre Visual Archives, p. 567; Eye of Science, p. 34; Clive Freeman/Biosym Technologies, pp. 345, 558a; Nicolle R. Fuller, p. 36b; Pascal Goetgheluck, p. 311; Adam Hart-Davis, p. 537; Gary Hincks, p. 522; Mikkel Jensen, p. 269; Kallista Images/Custom Medical Stock Photo, p. 188c; Russell Kightley, pp. 31b, 218l; Mehau Kulyk, p. 1; Laguna Design, pp. 186c, 558b; Library of Congress/Science Photo Library, p. 426b; Claus Lunau, p. 449; Massimo Breda, The Lighthouse, p. 397; David McCarthy, p. 161tl; Medi-mation, p. 31br; B. Murton/Southampton Oceanography Centre, p. 478; National Optical Astronomy Observatories, p. 127b; Susumu Nishinaga, p. 30a; Northwestern University, p. 30e; Ornl, p. 388; Power and Syred, pp. 30b, 400; Public Health England, pp. 55b, 55c; Ramon Andrade 3Dciencia, p. 214; Paul Rapson, p. 124; J.C. Revy, ISM, pp. 167-168, 477; Alexis Rosenfeld, p. 433b; Science Picture Co, pp. 160c, 218r; Scientific Picture Research, pp. 72, 78; TEK Image, pp. 211, 298-299; Sheila Terry, p. 426a; Tina Carvalho, Visuals Unlimited, p. 18tl; Geoff Tompkinson, p. 387; Victor Habbick Visions, p. 213; Peidong Yang/UC Berkeley, p. 160b.
Shutterstock: 02creationz, p. 35b; Africa Studio, p. 36b; Stanislav Akimkim, p. 80b; akiyoko, pp. 534-535; AndreyCherkasov, p. 79tl; Anneka, pp. 151c, 473b; Egorov Artem, p. 195; Art_of_sun, pp. 26-27; Mile Atanasov, p. 80a; atdigit, p. 453; Marijus Auruskevicius, p. 186b; max blain, p. 323tr; Nadezda Boltaca, p. 155tr; Steve Bower, p. 377; Boykung, p. 141b; Natalia Bratslavsky, p. 160t; Bro Studio, pp. 64, 75, 124; Charles Brutlag, p. 32b; Rob Byron, p. 282a; Chromatos, p. 370; cunaplus, p. 65b; Dado Photos, p. 150; Comaniciu Dan, p. 5lc; Antonio Danna, p. 244a; Dainis Derics, p. 98tr; dio5050, p. 366; Designua, p. 563; Vereshchagin Dmitry, p. 33r; doomu, p. 80c; Everett Historical, p. 69-3; Featureflash, p. 147c; Flegere, p. 5r; Fly_dragonfly, p. 464; Frederic Legrand - COMEO, p. 217; Gjermund, p. 80d; Goodluz, p. 266; Gjermund, p. 32t; Jason and Bonnie Grower, p. 264; Antonio Guillem, p. 14tr; hanabiyori, p. 244b; Brent Hofacker, p. 5l; Ibreakstock, p. 368; indianstockimages, p. 140b; irin-k, p. 214r; Isak55, pp. 74-75; ituchka, p. 5rc; Atiwan Janprom, p. 14br; Szasz-Fabian Jozsef, p. 149c; marcin jucha, pp. 222-223, 224b; Georgios Kollidas, p. 69-1; Gerard Koudenburg, p. 21l; ktsdesign, p. 18c; Dominique Landau, p. 141r; JP Lavoie, p. 188b; Viacheslav Lopatin, p. 138c; Albert Lozano, p. 55a; Michal Ludwiczak, p. 503; Lumena, p. 1; Magnetix, pp. 67, 353, 371, 508; Alexandr Makarov, p. 156t; Marekusz, p. 14cr; maxstockphoto, p. 417; mdmworks, p. 214; molekuul.be, pp. 208, 368, 561; Vincent Noel, p. 14bl; Sergey Novikov, p. 35t; Nuttapong, p. 229b; Jarun Ontakrai, p. 18r; OPgrapher, p. 523; David Orcea, p. 79br; Alon Othnay, p. 186t; Paket, p. 32t; Pakhnyushcha, pp. 31a, 318l; siamionau pavel, p. 229tl; Panom Pensawang, p. 323bl; Ivo Petkov, p. 268; Igor Petrushenko, p. 370; Leigh Prather, p. 79bl; PomInOz, p. 309; Renamarie, p. 17t; Ronsmith, p. 67l; Fouad A. Saad, pp. 287, 372, 431; Sciencepics, p. 362; Serger_Bogomyako, p. 263; zhang sheng p. 137; Sl_photo, pp. 447-448; Yury Smelov, p. 168b; Carolina K. Smith, p. 37r; sspopov, p. 318b; Stable, p. 229; steveball, p. 524; TebNad, p. 323br; Mary Terriberry, p. 523; tonyz20, p. 31tl; Tooykrub, p. 485; Andris Torms, p. 362; JR Trice, p. 98tl; udaix, p. 20; Belovodchenko Vadim, p. 316; Martin Valigursky, p. 42; Joao Virissimo, p. 265; Wallpaper photographer, p. 477; withGod, p. 65c; Feng Yu, p. 33l; Zelenskaya, p. 14tl.
TarraWarra Estate: p. 487
University of Idaho Coeur d' Alene: Marie Pengilly, p. 526r.
University of NSW: Centre for Quantum Computation and Communication Technology, p. 190.
Volvo Car Corporation: p. 216.
WolframAlpha: p. 129l.

Disclaimer
Some of the images used in *Pearson Chemistry 11 Queensland* might have associations with deceased Indigenous Australians. Please be aware that these images might cause sadness or distress in Aboriginal or Torres Strait Islander communities.

Groups

ATOMIC NUMBER → 12
SYMBOL → **Mg**
ELECTRON STRUCTURE → [Ne]$3s^2$
RELATIVE ATOMIC MASS ( ) indicates most stable isotope → 24.31
BOILING POINT °C → 1090
MELTING POINT °C → 650
ELECTRONEGATIVITY → 1.3
Magnesium

s BLOCK

d BLOCK

p BLOCK

Periods	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1									1 1.008 **H** -252.9 -259 2.2 $1s^1$ Hydrogen									2 4.003 **He** -268.9 -272 – $1s^2$ Helium
2	3 6.941 **Li** 1342 180 1.0 $1s^22s^1$ Lithium	4 9.012 **Be** 2468 1287 1.6 $1s^22s^2$ Beryllium											5 10.81 **B** 4000 2077 2.0 $1s^22s^22p^1$ Boron	6 12.01 **C** 4827 3500 2.6 $1s^22s^22p^2$ Carbon	7 14.01 **N** -195.8 -210 3.0 $1s^22s^22p^3$ Nitrogen	8 16.00 **O** -183 -219 3.4 $1s^22s^22p^4$ Oxygen	9 19.00 **F** -188.1 -220 4.0 $1s^22s^22p^5$ Fluorine	10 20.18 **Ne** -246 -249 – $1s^22s^22p^6$ Neon
3	11 22.99 **Na** 882.9 97.8 0.9 [Ne]$3s^1$ Sodium	12 24.31 **Mg** 1090 650 1.3 [Ne]$3s^2$ Magnesium											13 26.98 **Al** 2519 660.3 1.6 [Ne]$3s^23p^1$ Aluminium	14 28.09 **Si** 3265 1414 1.9 [Ne]$3s^23p^2$ Silicon	15 30.97 **P** 280.5 44.2 2.2 [Ne]$3s^23p^3$ Phosphorus	16 32.07 **S** 444.6 115.2 2.6 [Ne]$3s^23p^4$ Sulfur	17 35.45 **Cl** -34 -102 3.2 [Ne]$3s^23p^5$ Chlorine	18 39.95 **Ar** -185.8 -189 – [Ne]$3s^23p^6$ Argon
4	19 39.10 **K** 758.8 63.4 0.8 [Ar]$4s^1$ Potassium	20 40.08 **Ca** 1484 842 1.0 [Ar]$4s^2$ Calcium	21 44.96 **Sc** 2836 1541 1.4 [Ar]$3d^14s^2$ Scandium	22 47.87 **Ti** 3287 1670 1.5 [Ar]$3d^24s^2$ Titanium	23 50.94 **V** 3407 1910 1.6 [Ar]$3d^34s^2$ Vanadium	24 52.00 **Cr** 2671 1907 1.7 [Ar]$3d^54s^1$ Chromium	25 54.94 **Mn** 2061 1246 1.6 [Ar]$3d^54s^2$ Manganese	26 55.85 **Fe** 2861 1538 1.8 [Ar]$3d^64s^2$ Iron	27 58.93 **Co** 2927 1495 1.9 [Ar]$3d^74s^2$ Cobalt	28 58.69 **Ni** 2913 1455 1.9 [Ar]$3d^84s^2$ Nickel	29 63.55 **Cu** 2560 1085 1.9 [Ar]$3d^{10}4s^1$ Copper	30 65.38 **Zn** 907 419.5 1.6 [Ar]$3d^{10}4s^2$ Zinc	31 69.72 **Ga** 2229 27.8 1.8 [Ar]$3d^{10}4s^24p^1$ Gallium	32 72.64 **Ge** 2833 938.2 2.0 [Ar]$3d^{10}4s^24p^2$ Germanium	33 74.92 **As** 613 816.8 2.2 [Ar]$3d^{10}4s^24p^3$ Arsenic	34 78.96 **Se** 684.8 220.8 2.6 [Ar]$3d^{10}4s^24p^4$ Selenium	35 79.90 **Br** 58.8 -7.1 3.0 [Ar]$3d^{10}4s^24p^5$ Bromine	36 83.80 **Kr** -153.4 -157 – [Ar]$3d^{10}4s^24p^6$ Krypton
5	37 85.47 **Rb** 687.8 39 0.8 [Kr]$5s^1$ Rubidium	38 87.61 **Sr** 1377 769 1.0 [Kr]$5s^2$ Strontium	39 88.91 **Y** 3345 1522 1.2 [Kr]$4d^15s^2$ Yttrium	40 91.22 **Zr** 4406 1854 1.3 [Kr]$4d^25s^2$ Zirconium	41 92.91 **Nb** 4741 2477 1.6 [Kr]$4d^45s^1$ Niobium	42 95.96 **Mo** 4639 2622 2.2 [Kr]$4d^55s^1$ Molybdenum	43 (98) **Tc** 4262 2157 2.1 [Kr]$4d^55s^2$ Technetium	44 101.1 **Ru** 4147 2333 2.2 [Kr]$4d^75s^1$ Ruthenium	45 102.9 **Rh** 3695 1963 2.3 [Kr]$4d^85s^1$ Rhodium	46 106.4 **Pd** 2963 1555 2.2 [Kr]$4d^{10}5s^0$ Palladium	47 107.9 **Ag** 2162 961.8 1.9 [Kr]$4d^{10}5s^1$ Silver	48 112.4 **Cd** 766.8 321.1 1.7 [Kr]$4d^{10}5s^2$ Cadmium	49 114.8 **In** 2072 156.6 1.8 [Kr]$4d^{10}5s^25p^1$ Indium	50 118.7 **Sn** 2602 231.9 2.0 [Kr]$4d^{10}5s^25p^2$ Tin	51 121.8 **Sb** 1587 630.6 2.0 [Kr]$4d^{10}5s^25p^3$ Antimony	52 127.6 **Te** 987.8 449.5 2.1 [Kr]$4d^{10}5s^25p^4$ Tellurium	53 126.9 **I** 184.4 113.7 2.7 [Kr]$4d^{10}5s^25p^5$ Iodine	54 131.3 **Xe** -108.1 -112 2.6 [Kr]$4d^{10}5s^25p^6$ Xenon
6	55 132.9 **Cs** 670.8 28.5 0.8 [Xe]$6s^1$ Caesium	56 137.3 **Ba** 1845 727 0.9 [Xe]$6s^2$ Barium	57 138.9 **La** 2464 920 1.1 [Xe]$5d^16s^2$ Lanthanum	72 178.5 **Hf** 4600 2233 1.3 [Xe]$4f^{14}5d^26s^2$ Hafnium	73 180.9 **Ta** 5455 3017 1.5 [Xe]$4f^{14}5d^36s^2$ Tantalum	74 183.9 **W** 5555 3414 1.7 [Xe]$4f^{14}5d^46s^2$ Tungsten	75 186.2 **Re** 5596 3454 1.9 [Xe]$4f^{14}5d^56s^2$ Rhenium	76 190.2 **Os** 5008 3033 2.2 [Xe]$4f^{14}5d^66s^2$ Osmium	77 192.2 **Ir** 4428 2446 2.2 [Xe]$4f^{14}5d^76s^2$ Iridium	78 195.1 **Pt** 3825 1768 2.2 [Xe]$4f^{14}5d^96s^1$ Platinum	79 197.0 **Au** 2836 1064 2.4 [Xe]$4f^{14}5d^{10}6s^1$ Gold	80 200.6 **Hg** 356.6 -38.8 1.9 [Xe]$4f^{14}5d^{10}6s^2$ Mercury	81 204.4 **Tl** 1473 303.8 1.8 [Xe]$4f^{14}5d^{10}6s^26p^1$ Thallium	82 207.2 **Pb** 1749 327.5 1.8 [Xe]$4f^{14}5d^{10}6s^26p^2$ Lead	83 209.0 **Bi** 1564 271.4 1.9 [Xe]$4f^{14}5d^{10}6s^26p^3$ Bismuth	84 (210) **Po** 962 253.8 2.0 [Xe]$4f^{14}5d^{10}6s^26p^4$ Polonium	85 (210) **At** 366.8 301.8 2.2 [Xe]$4f^{14}5d^{10}6s^26p^5$ Astatine	86 (222) **Rn** -61.8 -71.2 – [Xe]$4f^{14}5d^{10}6s^26p^6$ Radon
7	87 (223) **Fr** 676.8 27 0.7 [Rn]$7s^1$ Francium	88 (226) **Ra** 1140 699.8 0.9 [Rn]$7s^2$ Radium	89 (227) **Ac** 3200 1050 1.1 [Rn]$6d^17s^2$ Actinium	104 (261) **Rf** – – – [Rn]$5f^{14}6d^27s^2$ Rutherfordium	105 (262) **Db** – – – [Rn]$5f^{14}6d^37s^2$ Dubnium	106 (266) **Sg** – – – [Rn]$5f^{14}6d^47s^2$ Seaborgium	107 (264) **Bh** – – – [Rn]$5f^{14}6d^57s^2$ Bohrium	108 (267) **Hs** – – – [Rn]$5f^{14}6d^67s^2$ Hassium	109 (268) **Mt** – – – [Rn]$5f^{14}6d^77s^2$ Meitnerium	110 (271) **Ds** – – – [Rn]$5f^{14}6d^87s^2$ Darmstadtium	111 (272) **Rg** – – – [Rn]$5f^{14}6d^97s^2$ Roentgenium	112 (285) **Cn** – – – [Rn]$5f^{14}6d^{10}7s^2$ Copernicium	113 (284) **Nh** – – – [Rn]$5f^{14}6d^{10}7s^17p^1$ Nihonium	114 (289) **Fl** – – – [Rn]$5f^{14}6d^{10}7s^27p^2$ Flerovium	115 (289) **Mc** – – – [Rn]$5f^{14}6d^{10}7s^27p^3$ Moscovium	116 (292) **Lv** – – – [Rn]$5f^{14}6d^{10}7s^27p^4$ Livermorium	117 (294) **Ts** – – – [Rn]$5f^{14}6d^{10}7s^27p^5$ Tennessine	118 (294) **Og** – – – [Rn]$5f^{14}6d^{10}7s^27p^6$ Oganesson

f BLOCK

58 140.1 **Ce** 3443 799 1.1 [Xe]$4f^25d^06s^2$ Cerium	59 140.9 **Pr** 3520 931 1.1 [Xe]$4f^35d^06s^2$ Praseodymium	60 144.2 **Nd** 3074 1016 1.1 [Xe]$4f^45d^06s^2$ Neodymium	61 (145) **Pm** 3000 1042 – [Xe]$4f^55d^06s^2$ Promethium	62 150.4 **Sm** 1794 1072 1.1 [Xe]$4f^65d^06s^2$ Samarium	63 152.0 **Eu** 1794 1072 – [Xe]$4f^75d^06s^2$ Europium	64 157.3 **Gd** 3273 1313 1.2 [Xe]$4f^75d^16s^2$ Gadolinium	65 158.9 **Tb** 3230 1359 – [Xe]$4f^95d^06s^2$ Terbium	66 162.5 **Dy** 2567 1412 1.2 [Xe]$4f^{10}5d^06s^2$ Dysprosium	67 164.9 **Ho** 2700 1472 1.2 [Xe]$4f^{11}5d^06s^2$ Holmium	68 167.3 **Er** 2868 1529 1.2 [Xe]$4f^{12}5d^06s^2$ Erbium	69 168.9 **Tm** 1950 1545 1.3 [Xe]$4f^{13}5d^06s^2$ Thulium	70 173.1 **Yb** 1196 824 – [Xe]$4f^{14}5d^06s^2$ Ytterbium	71 175.0 **Lu** 3402 1663 1.0 [Xe]$4f^{14}5d^16s^2$ Lutetium
90 232.0 **Th** 4875 1750 1.3 [Rn]$5f^06d^27s^2$ Thorium	91 231.0 **Pa** – 1572 1.5 [Rn]$5f^26d^17s^2$ Protactinium	92 238.0 **U** 4131 1135 1.7 [Rn]$5f^36d^17s^2$ Uranium	93 (237) **Np** 3902 644 1.3 [Rn]$5f^46d^17s^2$ Neptunium	94 (244) **Pu** 3228 640 1.3 [Rn]$5f^66d^07s^2$ Plutonium	95 (243) **Am** 2011 1176 – [Rn]$5f^76d^07s^2$ Americium	96 (247) **Cm** 3110 1340 – [Rn]$5f^76d^17s^2$ Curium	97 (247) **Bk** – 986 – [Rn]$5f^96d^07s^2$ Berkelium	98 (251) **Cf** – 900 – [Rn]$5f^{10}6d^07s^2$ Californium	99 (252) **Es** – – – [Rn]$5f^{11}6d^07s^2$ Einsteinium	100 (257) **Fm** – – – [Rn]$5f^{12}6d^07s^2$ Fermium	101 (258) **Md** – 827 – [Rn]$5f^{13}6d^07s^2$ Mendelevium	102 (259) **No** – – – [Rn]$5f^{14}6d^07s^2$ Nobelium	103 (262) **Lr** – – – [Rn]$5f^{14}6d^17s^2$ Lawrencium