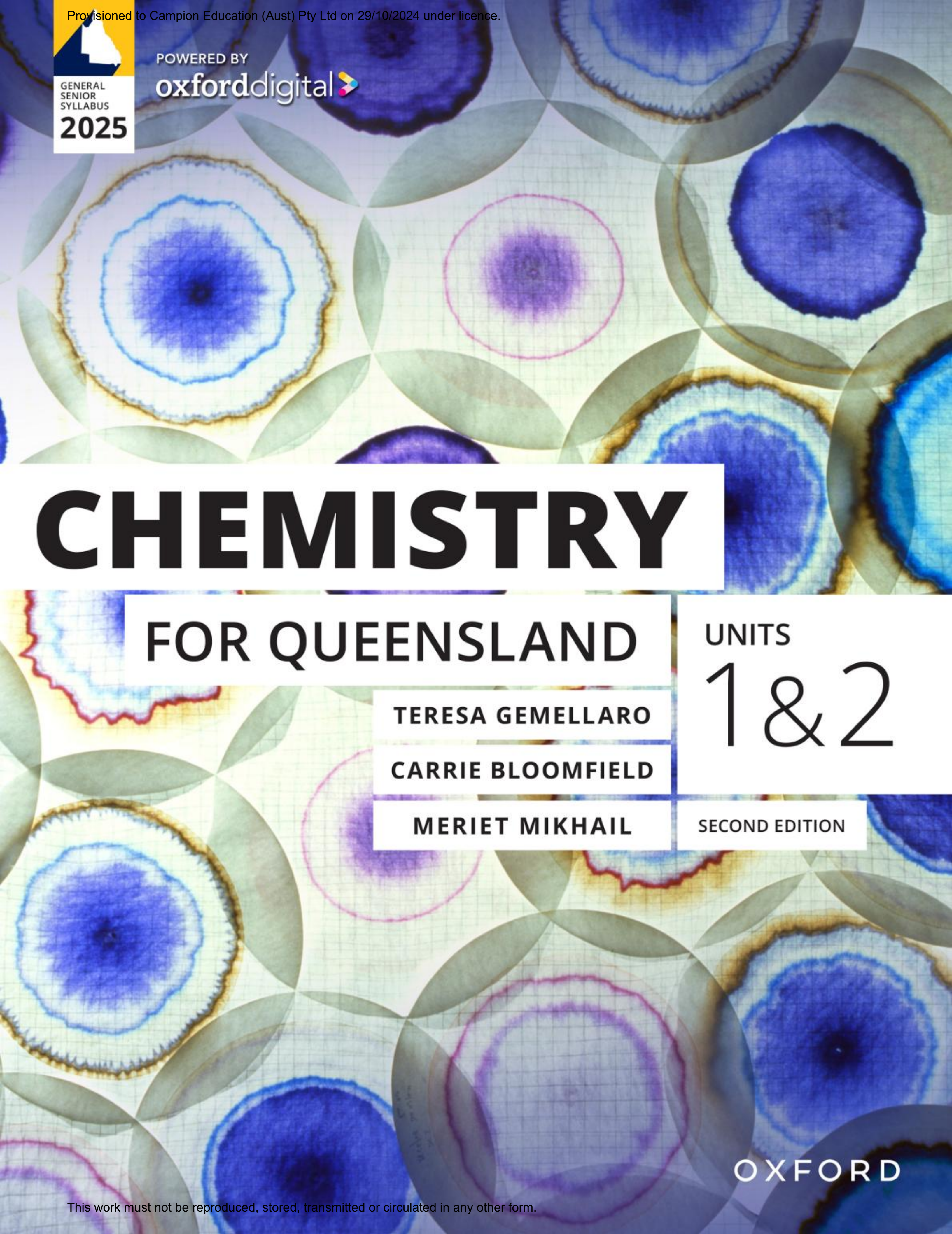

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CHEMISTRY

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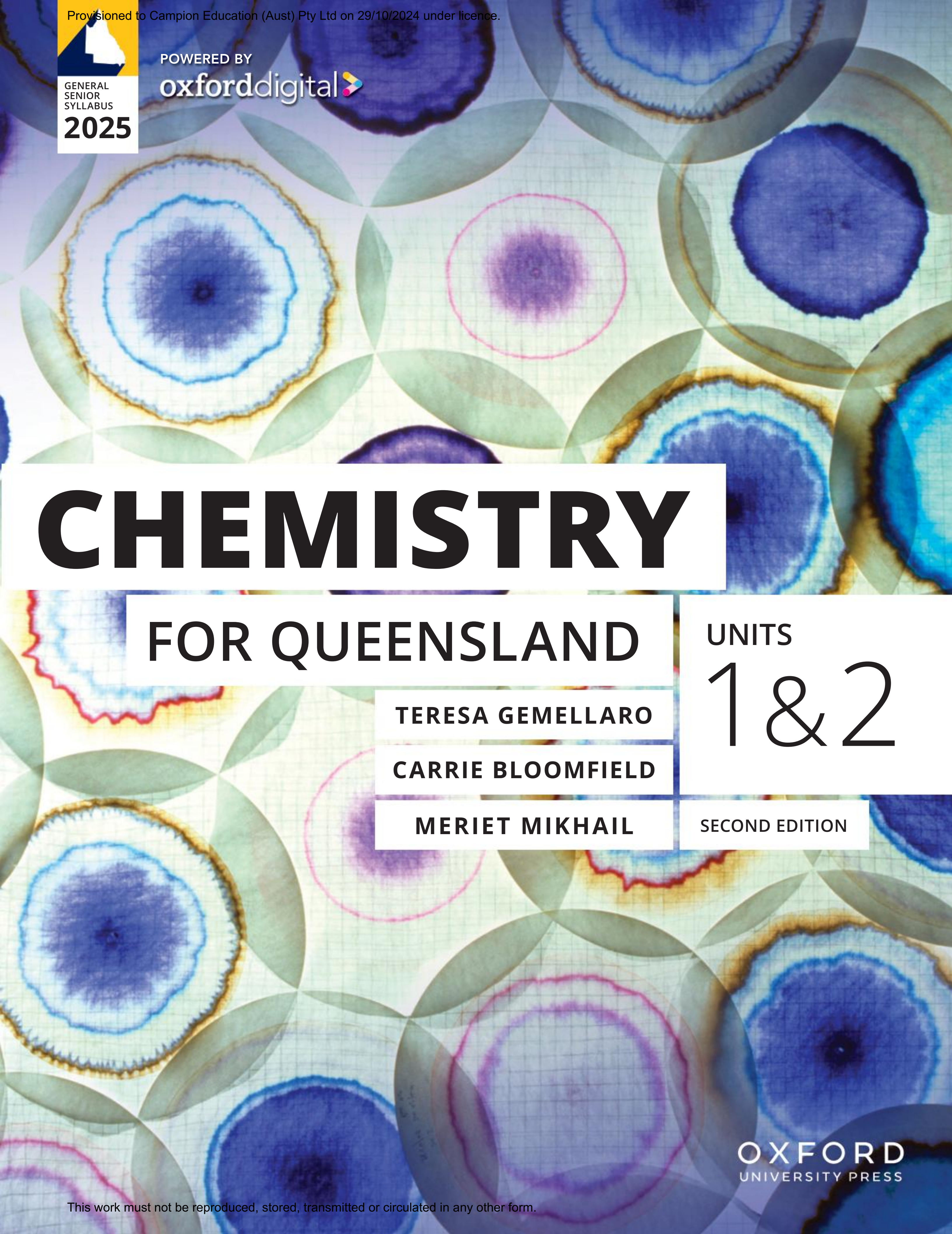

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Contents

▶ Lessons available on Oxford Digital.

Lessons available on Oxford Digital.

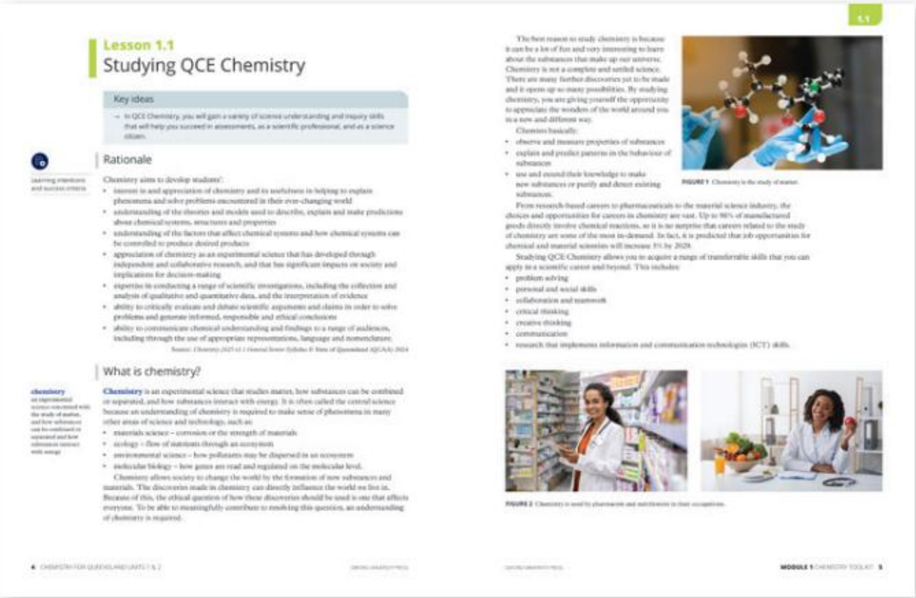
Introducing *Chemistry for Queensland Units 1 & 2* (Second edition)

Congratulations on choosing *Chemistry for Queensland Units 1 & 2* as part of your studies this year!

Chemistry for Queensland Units 1 & 2 has been purpose-written to meet the requirements of the QCAA Chemistry 2025 General senior syllabus. It includes a range of flexible print and digital products to suit your school and incorporates a wide variety of features designed to make learning fun, purposeful and accessible for all students!

Key features of the Student Books

The Chemistry toolkit module provides an overview of the syllabus, student-friendly guidance for every science inquiry skill and tips for success on assessment tasks.



Lesson 1.1

Studying QCE Chemistry

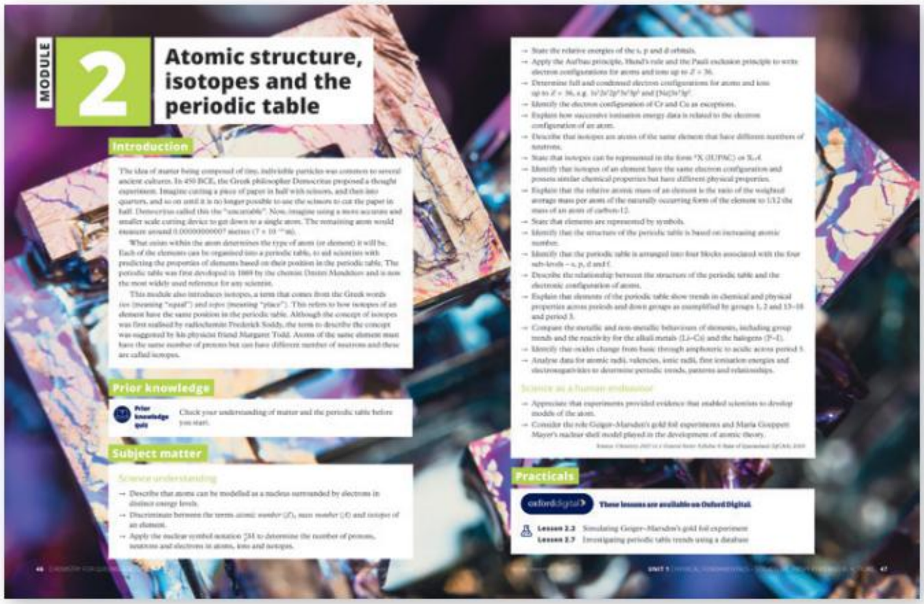
Key ideas

Rationale

What is chemistry?

Each module begins with a module opener that includes:

- QCAA subject matter
- reference to a supporting prior knowledge quiz that assesses and informs student understanding of pre-requisite concepts
- a list of practical lessons that support science inquiry.



MODULE 2

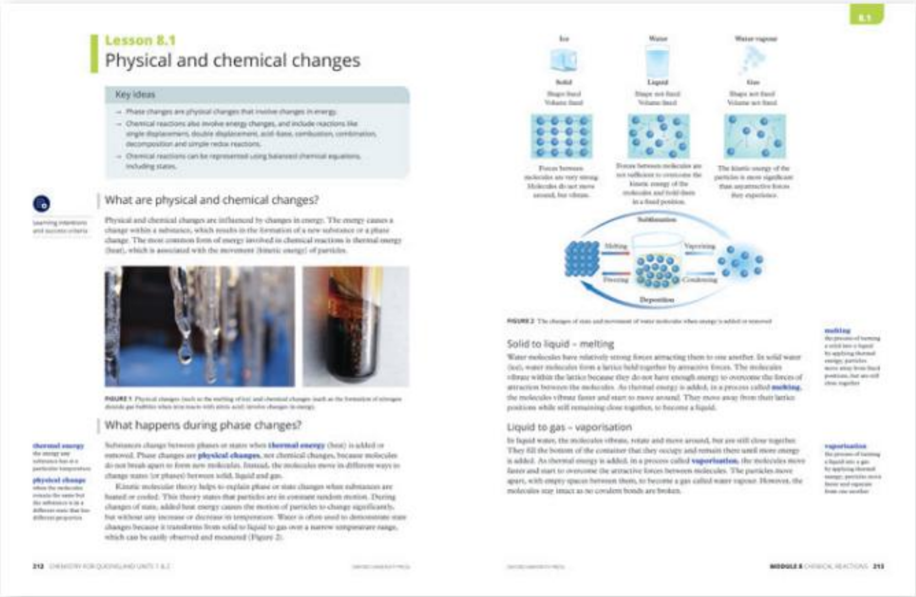
Atomic structure, isotopes and the periodic table

Introduction

Prior knowledge

Subject matter

Practicals



Lesson 8.1

Physical and chemical changes

Key ideas

What are physical and chemical changes?

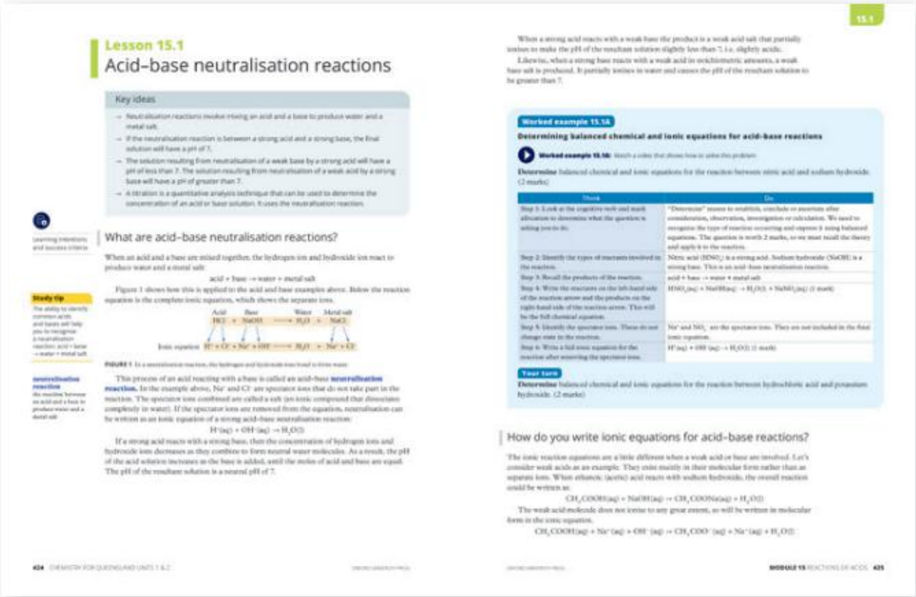
What happens during phase changes?

Solid to liquid – melting

Liquid to gas – vaporisation

Each lesson includes:

- **learning intentions and success criteria**
- clearly structured content written in clear, concise language
- definitions for all key terms on the page
- engaging, relevant and informative images and illustrations
- a range of tips and features designed to bring course content to life including **study tips, worked examples, skill drills** and examples of **real-world science applications**
- references to supporting **digital resources**
- **Check your learning** activities organised according to **Marzano and Kendall's taxonomy** and incorporating **cognitive verbs**.



Lesson 15.1

Acid–base neutralisation reactions

Key ideas

What are acid–base neutralisation reactions?

How do you write ionic equations for acid–base reactions?

Each module contains a range of practical activities designed to meet the requirements of science understanding and science inquiry subject matter and develop science inquiry skills.

Find out more

For a complete overview of all the features and benefits of this Student Book – as well as helpful videos showing you how to get the most out of the series:

> activate your digital access (using the instructions on the inside front cover of this book) and click on "Introducing *Chemistry for Queensland Units 1 & 2*" in the Course menu

Key features of Oxford Digital

Oxford Digital has been designed in consultation with Australian teachers for Australian classrooms. The new platform delivers fully accessible, reflowable course content with videos, auto-marking activities, interactives and more embedded right where you need them.

There's also a range of unique features designed to improve learning outcomes.

Key features for students

As a student, you can:

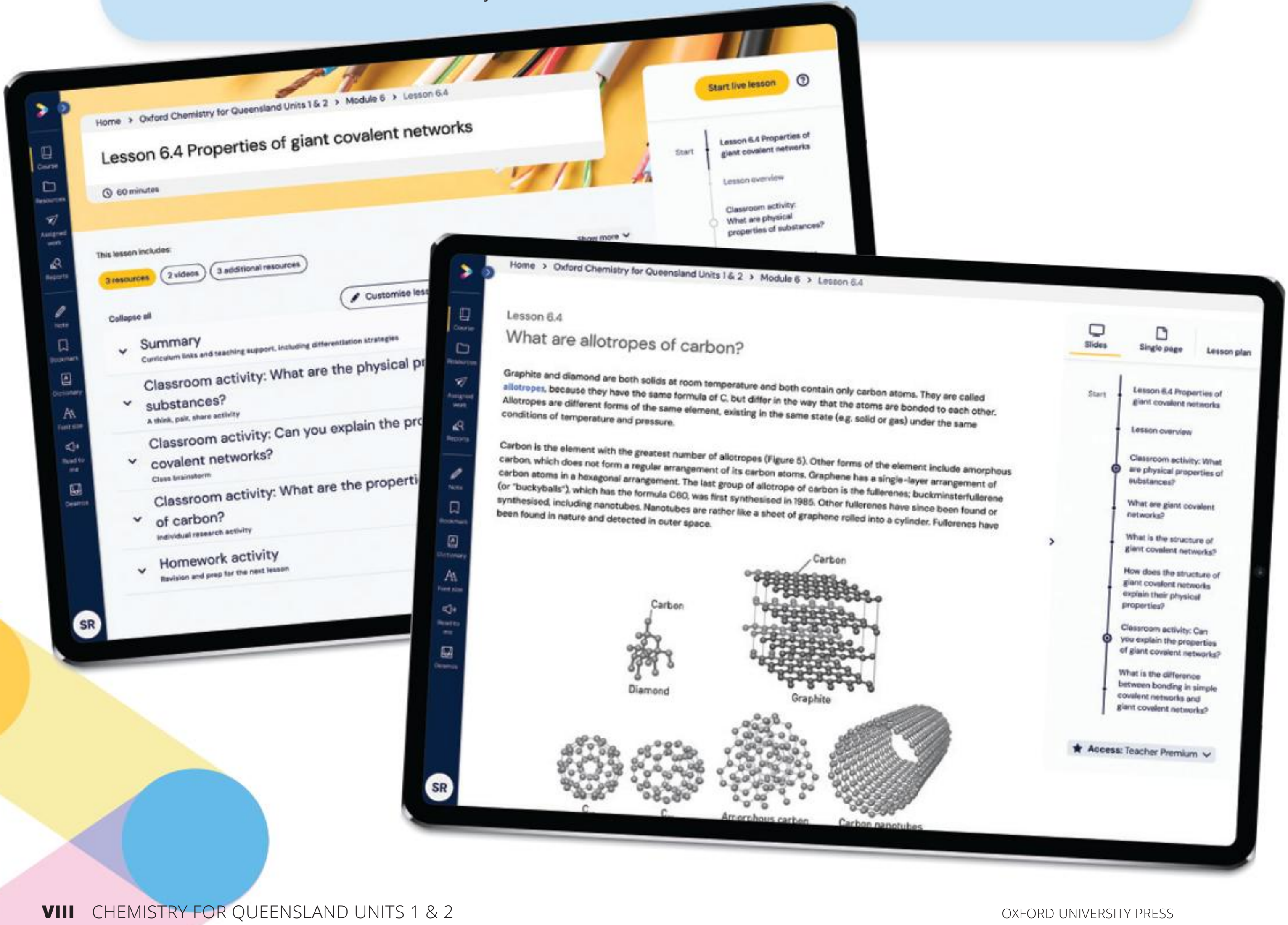
> **view** all course content in a **fully accessible**, reflowable format that's delivered in bite-sized chunks so you can work at you own pace

> use the "Read to me" button to have any part of the course read aloud to you

> **highlight, take notes, bookmark pages**, or define words with the built-in **Australian Oxford Dictionary**

> **watch** short **content videos, worked example videos** and **practical demonstration videos** to help you revise anything you don't understand, catch up on things you've missed, or help you with your homework

> **complete** hundreds of interactive questions and quizzes as you work through the content and get the answers and results sent to you.



Lesson 6.4 Properties of giant covalent networks

60 minutes

This lesson includes:

Summary
Curriculum links and teaching support, including differentiation strategies

Classroom activity: What are the physical pr
substances?
A think, pair, share activity

Classroom activity: Can you explain the pro
covalent networks?
Class brainstorm

Classroom activity: What are the properti
of carbon?
Individual research activity

Homework activity
Revision and prep for the next lesson

Lesson 6.4

What are allotropes of carbon?

Graphite and diamond are both solids at room temperature and both contain only carbon atoms. They are called allotropes, because they have the same formula of C, but differ in the way that the atoms are bonded to each other. Allotropes are different forms of the same element, existing in the same state (e.g. solid or gas) under the same conditions of temperature and pressure.

Carbon is the element with the greatest number of allotropes (Figure 5). Other forms of the element include amorphous carbon, which does not form a regular arrangement of its carbon atoms. Graphene has a single-layer arrangement of carbon atoms in a hexagonal arrangement. The last group of allotrope of carbon is the fullerenes; buckminsterfullerene (or "buckyballs"), which has the formula C60, was first synthesised in 1985. Other fullerenes have since been found or synthesised, including nanotubes. Nanotubes are rather like a sheet of graphene rolled into a cylinder. Fullerenes have been found in nature and detected in outer space.

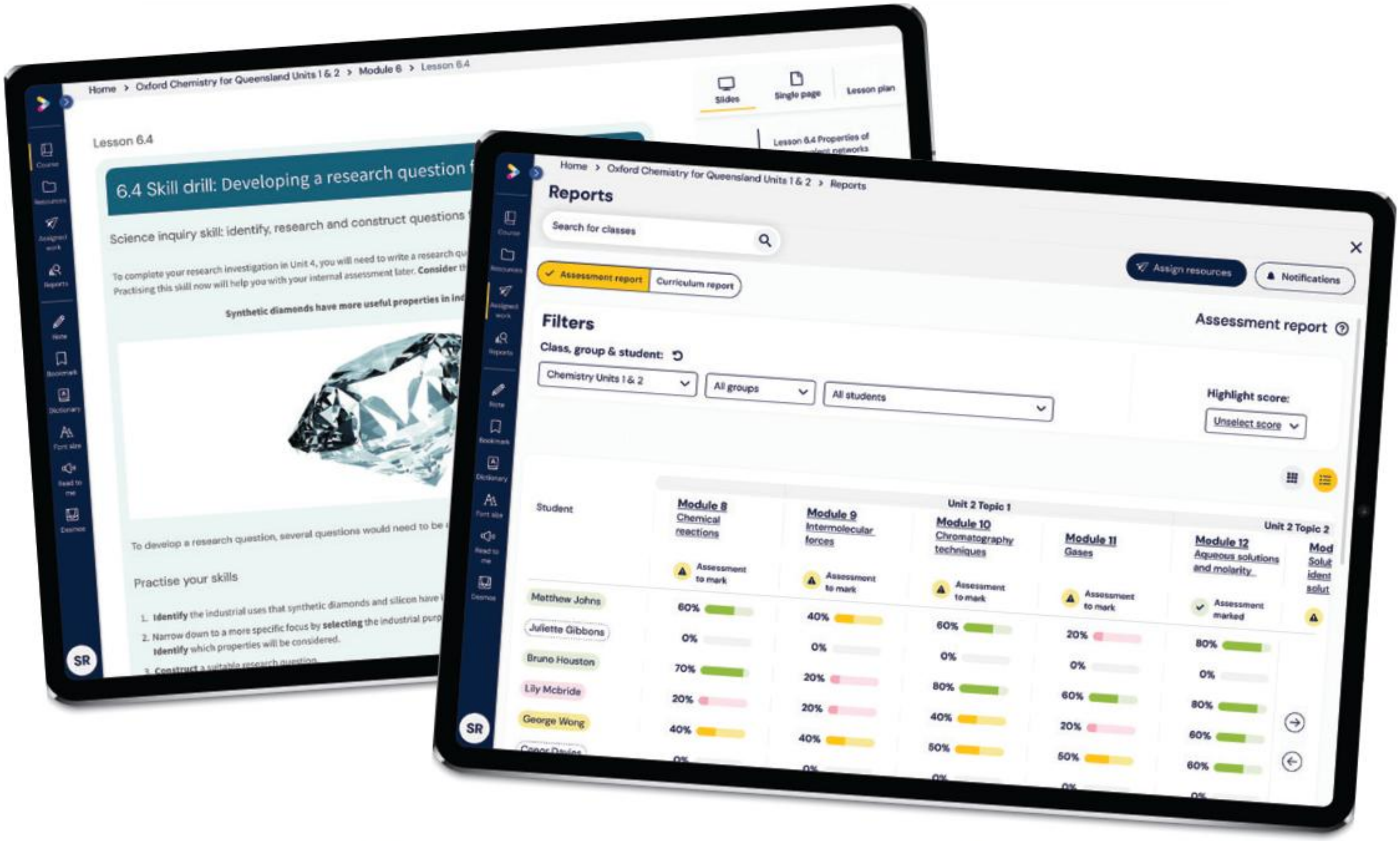
Key features for teachers

As a teacher, you can:

> **elevate** your teaching and **reduce planning and preparation time** with **Live Lesson mode**. This is an Australian first that lets you upgrade from traditional print-based lesson plans to **fully interactive, perfectly sequenced and timed interactive lessons complete with classroom activities** that are ready to go

> **personalise** learning for every student and **differentiate** content based on student strengths and weaknesses. Assign support or extension resources to any student using a range of differentiation resources

> **revolutionise** your planning, marking and reporting with powerful analytics on student performance and progress.



Lesson 6.4

6.4 Skill drill: Developing a research question f

Science inquiry skill: identify, research and construct questions f

To complete your research investigation in Unit 4, you will need to write a research qu
Practising this skill now will help you with your internal assessment later. **Consider** th

Synthetic diamonds have more useful properties in ind

To develop a research question, several questions would need to be

Practise your skills

1. **Identify** the industrial uses that synthetic diamonds and silicon have
2. Narrow down to a more specific focus by **selecting** the industrial purp
Identify which properties will be considered.

		Unit 2 Topic 1			Unit 2 Topic 2
Student	Module 8 Chemical reactions	Module 9 Intermolecular forces	Module 10 Chromatography techniques	Module 11 Gases	Module 12 Aqueous solutions and molarity
	Assessment to mark	Assessment to mark	Assessment to mark	Assessment to mark	✓ Assessment marked
Matthew Johns	60%	40%	60%	20%	80%
Juliette Gibbons	0%	0%	0%	0%	0%
Bruno Houston	70%	20%	80%	0%	80%
Lily Mcbride	20%	20%	40%	60%	60%
George Wong	40%	40%	50%	20%	60%

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Meet the authors & reviewers

Teresa Gemellaro

Author

Teresa holds a Bachelor of Science and Graduate Diploma in Education from the University of Queensland. She has taught chemistry, mathematics and science since 1980, with the majority of her career spent at Ayr State High School in North Queensland. Teresa is an experienced chemistry text book author, and has worked for the QCAA and its predecessor in many different roles – both chemistry-related and other positions – since the mid 1980s.

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Author

Carrie Bloomfield has been teaching VCE Chemistry and science in secondary schools for the past 11 years and is currently a leading teacher at Mount Eliza Secondary College. She is a VCAA assessor for Chemistry and has presented on science inquiry skills at previous workshops for the Oxford Science series. Carrie co-authored the Oxford University Press's first editions of the *Chemistry for Queensland Units 1–4* and *Chemistry for VCE Units 1–4 series*.

Dr Meriet Mikhail

Author

Meriet Mikhail was awarded a PhD in Medicine from the University of Sydney in 2005, and has been the recipient of various awards and grants, presenting at both national and international conferences for vaccine development and disease epidemiology. She is currently a QCE and IB senior chemistry teacher, with over 12 years of experience teaching QCE Chemistry. Meriet is currently the Head of Chemistry at Anglican Church Grammar School, and is a QCAA Subject Matter Expert for Chemistry.

We would also like to acknowledge and thank the following authors for their contributions: Catherine Millar, Geoffrey Giles, Hamilton Wright, Helen Silvester, Jennie Nash, Krystle Kuipers, Paul Devlin, Paul Keillor and Philip Sharpe.

Bernice Zaro

First Nations reviewer

Bernice is a proud Aboriginal and Torres Strait Islander Woman with a strong passion for educational greatness through culturally-inclusive learning. With family cultural heritage connection to the Gubbi Gubbi and Bundjalung of South East Queensland area and Kemer Kemer Meriam Nation and Maluligal Nation of the Torres Strait, she is inspired to share and learn continuously. Bernice along with her husband Aicey Zaro, a recognised traditional Artist, have been educating schools and communities through Cultural Awareness art workshops for over 15 years during their time managing the Zaro Cultural Gallery in the Burdekin region. Bernice has a passion for learning through her ongoing studies in Community Development, Child Wellbeing, Cultural Diversity and also sharing personal experiences through family, community and business, which inspire her to take on new opportunities.

Malcolm Corney

Reviewer

Malcolm holds a Bachelor of Applied Science (Applied Chemistry), a Graduate Diploma of Computer Science, a Masters of Information Technology (Research) and a Graduate Diploma of Education (Secondary). He has worked as an analytical chemistry in the fields of wastewater, pharmaceuticals and magnesium metal production for 20 years. He taught computer science and computer forensics at QUT for 15 years and has published in the fields of computer forensics and computer science education. He has taught STEM and senior chemistry at Kelvin Grove State College for eight years, where he leads the design of teaching and learning programs for these subjects. He has been involved as a QCAA marker and writer and hopes that he will soon be able to play bridge more than once a week.

Dr Philip Sharpe

Reviewer

Philip is a lecturer in Chemistry at the University of Queensland, teaching introductory, general, organic and biological inorganic chemistry since 2008 and is the immediate past Director of First Year Chemistry. He has received School and Faculty awards for his teaching and a national award for University Teaching as part of the First Year Chemistry teaching team. He has worked in Australia and the UK in the areas of macrocyclic coordination chemistry, electropolymers containing metal centres as nitrite sensors, amyloid formation in artificial proteins and iron-binding drugs for cancer treatment. He is a long-term member of the Royal Australian Chemical Institute (RACI) Queensland Chemistry Education group, which organises the long-standing and popular titration competition in Queensland.

MODULE 1

Chemistry toolkit

Introduction

Chemistry requires research skills to experiment and test theories in order to gain further knowledge. A true experiment always aims to discover natural laws of cause and effect – how do changes in one property cause changes in another?

This module will become a useful reference throughout Units 1 and 2 of Chemistry. It teaches you the principles that underpin how we approach investigating chemical phenomena and discovering new information about our natural world. It is only logical that it should come first!

This module is set out in a way that makes each piece of information easy to access. It is not meant to be read from beginning to end. Rather, it's like a toolbox – you dip your hand into it, get the tool you need and then apply it.

Prior knowledge

Prior knowledge quiz

Check your understanding of the science inquiry skills before you start.

oxforddigital Full lessons for this module are available on Oxford Digital.

Online-only Lessons

Lesson 1.1

Studying QCE Chemistry

Key ideas

→ In QCE Chemistry, you will gain a variety of science understanding and inquiry skills that will help you succeed in assessments, as a scientific professional, and as a science citizen.

Learning intentions and success criteria

Rationale

Chemistry aims to develop students':

- interest in and appreciation of chemistry and its usefulness in helping to explain phenomena and solve problems encountered in their ever-changing world
- understanding of the theories and models used to describe, explain and make predictions about chemical systems, structures and properties
- understanding of the factors that affect chemical systems and how chemical systems can be controlled to produce desired products
- appreciation of chemistry as an experimental science that has developed through independent and collaborative research, and that has significant impacts on society and implications for decision-making
- expertise in conducting a range of scientific investigations, including the collection and analysis of qualitative and quantitative data, and the interpretation of evidence
- ability to critically evaluate and debate scientific arguments and claims in order to solve problems and generate informed, responsible and ethical conclusions
- ability to communicate chemical understanding and findings to a range of audiences, including through the use of appropriate representations, language and nomenclature.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

What is chemistry?

chemistry
an experimental science concerned with the study of matter, and how substances can be combined or separated and how substances interact with energy

Chemistry is an experimental science that studies matter, how substances can be combined or separated, and how substances interact with energy. It is often called the central science because an understanding of chemistry is required to make sense of phenomena in many other areas of science and technology, such as:

- materials science – corrosion or the strength of materials
- ecology – flow of nutrients through an ecosystem
- environmental science – how pollutants may be dispersed in an ecosystem
- molecular biology – how genes are read and regulated on the molecular level.

Chemistry allows society to change the world by the formation of new substances and materials. The discoveries made in chemistry can directly influence the world we live in. Because of this, the ethical question of how these discoveries should be used is one that affects everyone. To be able to meaningfully contribute to resolving this question, an understanding of chemistry is required.

The best reason to study chemistry is because it can be a lot of fun and very interesting to learn about the substances that make up our universe. Chemistry is not a complete and settled science. There are many further discoveries yet to be made and it opens up so many possibilities. By studying chemistry, you are giving yourself the opportunity to appreciate the wonders of the world around you in a new and different way.

Chemists basically:

- observe and measure properties of substances
- explain and predict patterns in the behaviour of substances
- use and extend their knowledge to make new substances or purify and detect existing substances.

FIGURE 1 Chemistry is the study of matter.

From research-based careers to pharmaceuticals to the material science industry, the choices and opportunities for careers in chemistry are vast. Up to 96% of manufactured goods directly involve chemical reactions, so it is no surprise that careers related to the study of chemistry are some of the most in-demand. In fact, it is predicted that job opportunities for chemical and material scientists will increase 5% by 2029.

Studying QCE Chemistry allows you to acquire a range of transferrable skills that you can apply in a scientific career and beyond. This includes:

- problem solving
- personal and social skills
- collaboration and teamwork
- critical thinking
- creative thinking
- communication
- research that implements information and communication technologies (ICT) skills.

FIGURE 2 Chemistry is used by pharmacists and nutritionists in their occupations.

How is the QCE Chemistry course structured?

Studying QCE Chemistry provides you with the opportunity to engage in a range of inquiry tasks and develop science inquiry skills.

Syllabus objectives

Like the other senior sciences, there are six syllabus objectives in QCE Chemistry.

1 **Describe ideas and findings.**
Students use scientific representations and language in appropriate genres to give a detailed account of scientific phenomena, concepts, theories, models and systems.

2 **Apply understanding.**
Students use scientific concepts, theories, models and systems within their limitations. They use algebraic, visual and graphical representations of scientific relationships and data to determine unknown scientific quantities or features. They explain phenomena, concepts, theories, models, systems and modifications to methodologies.

3 **Analyse data.**
Students consider scientific information from primary and secondary sources to identify trends, patterns, relationships, limitations and uncertainty. In qualitative data, they identify the essential elements, features or components. In quantitative data, they use mathematical processes and algorithms. They identify data to support ideas, conclusions or decisions.

4 **Interpret evidence.**
Students use their understanding of scientific concepts, theories, models and systems and their limitations to draw conclusions and develop scientific arguments. They compare, deduce, extrapolate, infer, justify and make predictions based on their analysis of data.

5 **Evaluate conclusions, claims and processes.**
Students critically reflect on the available evidence and make judgements about its application to research questions. They extrapolate findings to support or refute claims. They use the quality of the evidence to evaluate the validity and reliability of inquiry processes and suggest improvements and extensions for further investigation.

6 **Investigate phenomena.**
Students develop rationales and research questions for experiments and investigations. They modify methodologies to collect primary data and select secondary sources. They manage risks, environmental and ethical issues and acknowledge sources of information.

Study tip

Your assessments in QCE Chemistry are mapped to these syllabus objectives. The data test (Lesson 1.10) assesses objectives 2 to 4. The student experiment (Lesson 1.11) and research investigation (Lesson 1.12) assess objectives 1 to 6. The external exam (Lesson 1.13) assesses objectives 1 to 4.

Subject matter

The structure of the QCE Chemistry course is laid out in the Chemistry General Senior Syllabus. The course consists of four units. Units 1 and 2 are completed in the first year of the QCE Chemistry course and Units 3 and 4 in the second year. Each unit is divided into topics and each topic can include science understanding, science as a human endeavour and science inquiry subject matter. You should be familiar with these categories of understanding from your studies in years 7 to 10.

An overview of the QCE Chemistry units is shown in Figure 3 and Units 1 and 2 are summarised in Table 1. Each unit has its own specific objectives, which are outlined in the Unit 1 opener and Unit 2 opener.

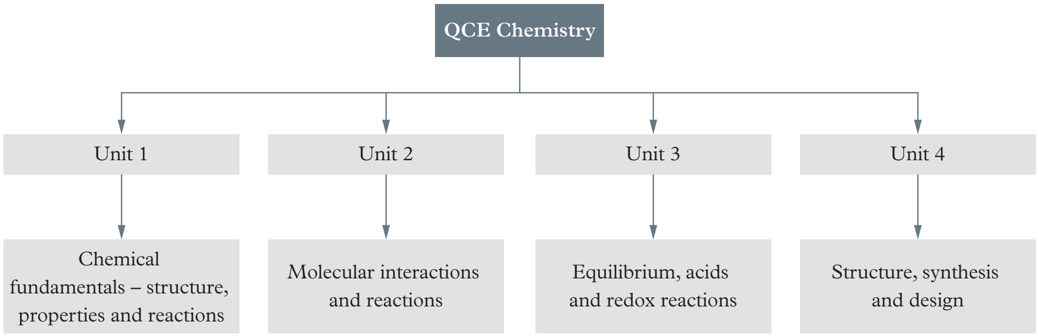

FIGURE 3 The structure of the QCE Chemistry course

TABLE 1 Topics in Units 1 and 2 Chemistry

Unit 1 Chemical fundamentals – structure, properties and reactions	
Topic	**Description**
1. Properties and structure of atoms	In this topic, you will learn about: • atomic structure and isotopes • analytical techniques • the periodic table and trends • the basics of bonding
2. Properties and structure of materials	In this topic, you will learn about: • compounds and mixtures • bonding and properties
3. Chemical reactions – reactants, products and energy change	In this topic, you will learn about: • chemical reactions • exothermic and endothermic reactions • the mole concept and law of conservation of mass
Unit 2 Molecular interactions and reactions	
Topic	**Description**
1. Intermolecular forces and gases	In this topic, you will learn about: • intermolecular forces • chromatography techniques • gases
2. Aqueous solutions and acidity	In this topic, you will learn about: • aqueous solutions and molarity • identifying ions in solution • solubility • pH and reactions of acids
3. Rates of chemical reactions	In this topic, you will learn about: • rates of reactions

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Assessment in QCE Chemistry

In Units 1 and 2, you can be assessed in several ways across the different topics. The syllabus requires that you:

- complete at least two, but no more than four assessments
- complete at least one assessment for each unit
- are assessed on each unit objective at least once.

Many schools assess students studying Units 1 and 2 as they would for students studying Units 3 and 4. This means that you will likely complete three assessment pieces and an end-of-year examination or examinations. One possible structure of your assessment is outlined in Table 2. Keep in mind that your school might choose to conduct the data test, student experiment and research investigation in any of Units 1 or 2.

TABLE 2 Units 1 and 2 assessments

Unit and assessment type	Assessment description	Assessment objectives
Unit 1 Chemical fundamentals – structure, properties and reactions: Data test	Students respond to items using qualitative data and/or quantitative data derived from practicals, activities or simulations from Unit 1	2. Apply understanding of properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change. 3. Analyse data about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change. 4. Interpret evidence about properties and structure of atoms and materials, and chemical reactants, products and energy change.
Unit 1 Chemical fundamentals – structure, properties and reactions: Student experiment	Students modify (i.e. refine, extend or redirect) an experiment relevant to Unit 1 subject matter to address their own related hypothesis or question. This assessment provides opportunities to assess science inquiry skills.	1. Describe ideas and findings about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change. 2. Apply understanding of properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change. 3. Analyse data about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change. 4. Interpret evidence about properties and structure of atoms and materials, and chemical reactants, products and energy change. 5. Evaluate processes, claims and conclusions about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change. 6. Investigate phenomena associated with properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.
Unit 2 Molecular interactions and reactions: Research investigation	Students gather evidence related to a research question to evaluate a claim relevant to Unit 2 subject matter. This assessment provides opportunities to assess science inquiry skills and science as a human endeavour (SHE) subject matter.	1. Describe ideas and findings about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions. 2. Apply understanding of intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions. 3. Analyse data about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions. 4. Interpret evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions. 5. Evaluate processes, claims and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions. 6. Investigate phenomena associated with intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
Units 1 and 2 examination/s		

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You can use Lesson 1.10 Preparing for your data test, Lesson 1.11 Conducting your student experiment, Lesson 1.12 Conducting your research investigation and Lesson 1.13 Preparing for your exams to guide you through these assessments. Note that Science as a human endeavour content will not be directly assessed in your examinations.

What are the science inquiry skills?

In addition to developing your science understanding in Chemistry (which we will cover in Modules 2 to 16), the QCE course requires you to develop and apply a range of science inquiry skills. These skills are specified in the QCE Chemistry General Senior Syllabus and are listed on the opening pages of this module. This module will help you develop these skills.

The science inquiry skills are applicable to all areas of study in Units 1 to 4 of the QCE Chemistry course. They are especially important for preparing and planning for your data test, student experiment and research investigation assessment tasks.

FIGURE 4 Having good science inquiry skills prepares you for success in QCE Chemistry (and beyond!)

Check your learning 1.1

Check your learning 1.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define** the term “chemistry” in 10 words or less. (1 mark)

2 **Recall** three transferrable skills that you can learn during your QCE Chemistry studies. (1 mark)

Analytical processes

3 **Judge** whether this is true: “Chemistry is said to be an experimental science, so all chemistry theories have to come from experiments”. (2 marks)

Knowledge utilisation

4 **Investigate** how chemistry would be relevant to:
  - **a** a veterinarian working on finding a cure for chlamydia in koalas (1 mark)
  - **b** a technician monitoring water quality for Queensland Urban Utilities at the South Pine dam (1 mark)
  - **c** an art conservator working in the Queensland Museum (1 mark)
  - **d** a wheat farmer from the Darling Downs (1 mark)
  - **e** a marine biologist working on the Great Barrier Reef (1 mark)
  - **f** an environmental toxicologist working in Mt Isa (1 mark)
  - **g** an architect designing shade structures for a park in Townsville (1 mark)
  - **h** a perfume maker from Noosa (1 mark)
  - **i** a council worker controlling mosquitoes in Cairns. (1 mark)

5 **Propose** how chemistry might be used in physics and biology; and how chemistry relies on mathematics. (3 marks)

Lesson 1.2

Considering First Nations perspectives in Chemistry

Key ideas

→ First Nations peoples have longstanding scientific knowledge.

→ First Nations peoples have developed knowledge about the world by observing using all the senses, predicting and hypothesising, testing (trial and error) and making generalisations within specific contexts such as the use of food, natural materials, navigation and sustainability of the environment.

→ Correctly acknowledging cultural and/or language groups, avoiding Eurocentrism and critically evaluating sources of information can help you to respectfully engage with First Nations perspectives in QCE Chemistry.

Learning intentions and success criteria

Science inquiry skills

This lesson provides support for the following science inquiry skills:

- identify and implement strategies to manage risks, ethics and environmental impact, e.g.
    - cultural guidelines, protocols for working with the knowledges of First Nations peoples

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

oxforddigital **This lesson is available on Oxford Digital.**

Lesson 1.3

Understanding the scientific method

Key ideas

→ The scientific method is a circular process that involves making observations, formulating a hypothesis which will often lead to performing experiments or simulations, developing models or theorems, retesting and trailing.

→ Research questions define the scope of an investigation. They can be used to develop hypotheses which predict the outcome of the investigation.

Science inquiry skills

Learning intentions and success criteria

This lesson provides support for the following science inquiry skills:

- identify, research and construct questions for investigation
- propose hypotheses and/or predict possible outcomes

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Lesson 1.4

Planning investigations

Key ideas

→ A method outlines the steps followed in an experiment and lists all of the materials and equipment used.

→ Valid and reliable measurements can be obtained by carefully designing your investigation to collect sufficient data and minimise errors.

→ All measurements include errors or uncertainties, either systematic or random. It is important to consider these and implement strategies to minimise their effects when you plan your experiments.

Science inquiry skills

Learning intentions and success criteria

This lesson provides support for the following science inquiry skills:

- design investigations, including the procedure/s to be followed, the materials required, and the type and amount of primary and/or secondary data required to obtain valid and reliable evidence, e.g.
  - consider replicates, number of data points and quality of sources
  - identify the types of errors, extraneous variables or confounding factors that are likely to influence results and implement strategies to minimise systematic and random error
- use appropriate equipment, techniques, procedures and sources to systematically and safely collect primary and secondary data, e.g.
  - laboratory and field techniques: measurement, and equipment calibration
  - ICTs, scientific texts, databases, simulations, online sources
- suggest improvements and extensions to minimise uncertainty, address limitations and improve the overall quality of evidence, e.g.
  - analyse the impact of random error/measurement uncertainties and systematic errors in experimental work and determine how these errors/measurement uncertainties can be reduced
  - discriminate between random and systematic errors

  - identify that experimental design and procedure usually leads to systematic errors in measurement, which causes a deviation in a direction and that repeated trials and measurements will reduce random error but not systematic error

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Lesson 1.5

Considering safety and ethics

Key ideas

→ Good laboratory practices and lab safety allows you to identify risks and take measures to control them so that you and those sharing the lab space with you stay safe during scientific investigations.

→ Conducting experiments ethically involves considering the impacts of your investigation beyond the laboratory.

Science inquiry skills

Learning intentions and success criteria

This lesson provides support for the following science inquiry skills:

- identify and implement strategies to manage risks, ethics and environmental impact, e.g.
  - material safety data sheets
  - workplace health and safety guidelines
  - appropriate disposal methods
  - standard operating procedures

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Lesson 1.6

Collecting data

Key ideas

→ Single experimental measurements are reported using best estimates, indicators of measurement uncertainty and units.

→ Scientific notation is used to easily express extremely large or small values.

→ Significant figures are digits in a number that are known with certainty plus the first digit that is uncertain.

→ All measurements, information and observations should be recorded in your logbook.

Science inquiry skills

Learning intentions and success criteria

This lesson provides support for the following science inquiry skills:

- use scientific language and representations to systematically record information, observations, data and measurement error, e.g.
  - symbols, units and prefixes
  - tables, graphs and diagrams
  - logbooks
- translate information between graphical, numerical and/or algebraic forms, e.g.
  - units and measurement conversions
  - ratios and percentages
  - symbols and notation

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Worked examples

This lesson is supported by the following Worked examples:

- **Worked example 1.6A** Writing numbers using scientific notation
- **Worked example 1.6B** Rounding answers to the correct number of significant figures
- **Worked example 1.6C** Converting between units for physical quantities

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Lesson 1.7

Processing and analysing data

Key ideas

→ Data is processed and analysed to identify trends, patterns, relationships, limitations and uncertainty.

→ Absolute and percentage uncertainty give an idea of the precision of measurements.

→ Percentage error gives an idea of the accuracy of measurements.

→ Data can be summarised in a variety of ways, such as in a table, a scatterplot or a scientific diagram.

Learning intentions and success criteria

Science inquiry skills

This lesson provides support for the following science inquiry skills:

- use scientific language and representations to systematically record information, observations, data and measurement error, e.g.
  - indicators of measurement uncertainty and state measurement uncertainties as a range (±) to an appropriate precision, e.g. when adding or subtracting, the final answer should be given to the least number of decimal places, when multiplying or dividing, the final answer should be given to the least number of significant figures
  - identify that concentration can be represented in a variety of ways including, but not limited to, mol L^{-1}, g L^{-1} and ppm and that square brackets can be used to denote concentration
- use mathematical techniques to summarise data in a way that allows for identification of relevant trends, patterns, relationships, limitations and uncertainty, e.g.
  - mean
  - gradient analysis
  - scatterplots (with maximum and minimum trendlines and R^2)
  - propagate random error in data processing to show the impact of measurement uncertainties on the final result
  - apply simple treatment of error analysis, e.g. for functions such as addition and subtraction, absolute uncertainties should be added, for multiplication, division and powers, percentage uncertainties should be added
  - calculate the measurement uncertainties for processed data, including the use of absolute uncertainties of the mean (Formula: $\Delta\bar{x} = \frac{\text{absolute uncertainty}}{\text{measurement}} \times 100\%$)
  - calculate the percentage error, when the experimental result can be compared with a theoretical or accepted result (value)
    (Formula: percentage error (%) $= \left|\frac{\text{measured value} - \text{true value}}{\text{true value}}\right| \times 100\%$)
  - discriminate between absolute uncertainty and percentage error

- select and construct appropriate representations to present data and communicate findings, e.g.
  - summary tables
  - apply appropriate graphical representations to analyse data and draw conclusions
  - analyse data to identify trends, patterns and relationships; recognising error, uncertainty and limitations of evidence
  - interpret graphs in terms of the relationship between dependent and independent variables
- extrapolate findings to determine unknown values, predict outcomes and evaluate claims

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What information does data analysis give us?

Once we have obtained our raw data, we are ready to process and analyse it. The goal of data analysis is to find patterns or trends, whereas data interpretation involves explaining those patterns. This typically involves completing calculations and constructing scatterplots. By transforming the data into more meaningful forms, we can then analyse it to quantify uncertainty, identify limitations, construct evidence-based arguments and draw conclusions.

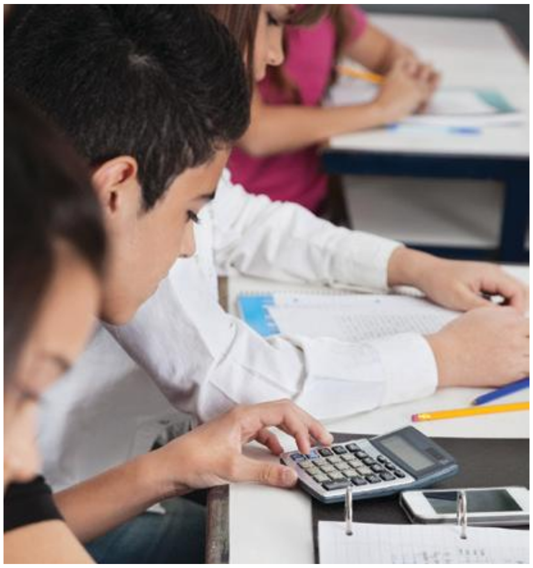


FIGURE 1 Analysing data, such as via mathematical techniques, helps us to draw conclusions from an investigation.

How do you use the measures of central tendency?

In the previous lesson, we talked about expressing measurements using a best estimate. This is straightforward when we are taking single measurements, but when there are a number of repeated measurements of the same value, there are some calculations we need to perform. These calculations allow us to better describe the measurement value and uncertainty.

To do this, we use measures of **central tendency**. When measurements involve random errors, the values tend to cluster around a central value. This is because random errors are just as likely to result in underestimates and overestimates of the "true" value. It is useful to discuss how close the values are to a central value using measures of central tendency.

central tendency
the tendency for repeated measurements of the same value to be grouped around the mean, mode or median

The normal distribution

Many natural variables tend to have a normal distribution. This is also sometimes referred to as a bell curve (Figure 2). On a normal distribution, the most likely value is the central value, which is the arithmetic **mean** or average of the results. The curve is symmetrical, with the probability of a particular value becoming lower as that value gets further away from the mean, in either direction. The curve can be more-or-less stretched out. This variation or spread in the results can be quantified by a value known as the **standard deviation**.

mean
the average of multiple values

standard deviation
a statistical value which expresses how spread out a group of values are or by how much they differ from the mean value for the group

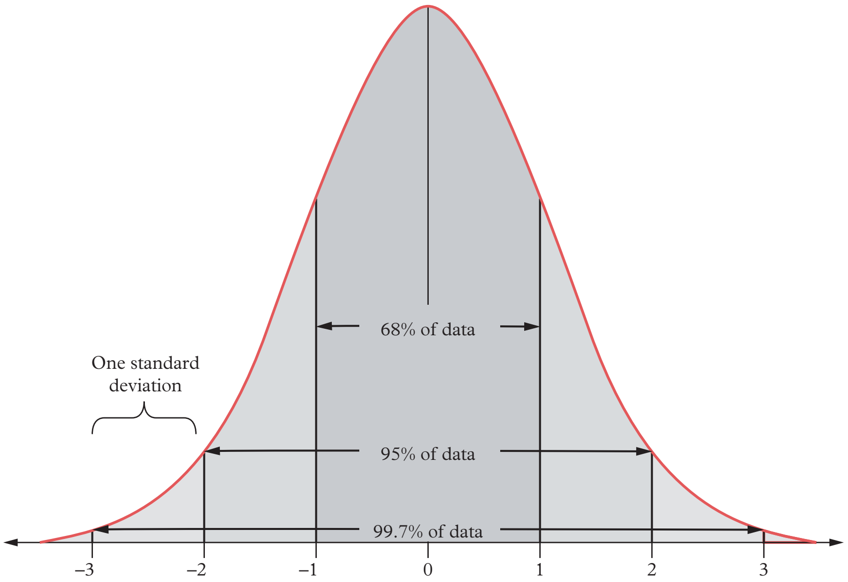


FIGURE 2 The normal distribution is also known as a bell curve.

For example, Table 1 shows the volume of aqueous hydrochloric acid needed to react with 20.0 mL of aqueous sodium hydroxide. Only six values are listed, but if the measurements were repeated several hundred times, then it is likely that a normal distribution of the values would be found around the mean.

TABLE 1 Volume of HCl(aq) required for reaction with 20.0 mL NaOH solution

Trial number	Volume of HCl(aq) (mL)
1	19.92
2	20.04
3	19.86
4	19.98
5	20.08
6	19.86

Calculating mean

The arithmetic mean of a group of numbers is the sum of the numbers divided by the number of measurements. This can be expressed mathematically:

$$\overline{x} = \frac{\Sigma x}{n}$$

where $\overline{x}$ is the mean, x is a measured value, and n is the number of measurements (replicates) and Σ (capital sigma) means to add all the values (x) together.

This formula can be applied to the data in Table 1. The mean here should be reported to two decimal places, matching the precision in the original measurements.

$$\overline{x} = \frac{\Sigma x}{n} = \frac{19.92\,\text{mL} + 20.04\,\text{mL} + 19.86\,\text{mL} + 19.98\,\text{mL} + 20.08\,\text{mL} + 19.86\,\text{mL}}{6}$$
$$= 19.96\,\text{mL}$$

outlier
a value that is much smaller or larger than most of the other values in a set of data; it is greater than three standard deviations away from the mean

One problem with the mean value is that it is sensitive to **outlier** values (those that are much larger or smaller than the other values). If there was an additional measurement of 40.00 mL, it would shift the average to 22.82 mL, which is quite far from the other values. In practice, a value of 40.00 mL would probably be rejected as an outlier.

Outliers are typically attributed to mistakes or random errors. Once identified, they should be omitted from data analysis.

Mode

The **mode** is another way of describing a group of values. The mode is the value that appears most often in a set of measurements. For the data in Table 1, this is 19.86 mL, because this value appears twice, but all the other values appear only once.

mode
the value that appears the most often in a dataset

Median

The **median** value is the value that appears in the middle of a sorted list of numbers, if there is an odd number of items. Therefore, there are an equal number of values above and below the median value. If there is an even number of measurements, then the median is the average value of the middle pair of values. For the data in Table 1 (19.86, 19.86, 19.92, 19.98, 20.04, 20.08), these are 19.92 and 19.98 mL, and the median value is their average, 19.95 mL.

median
the value that appears in the middle when the dataset is sorted from smallest to largest value

The median value can be very useful to report, particularly when there are extreme outliers in the data.

Standard deviation

The other parameter that describes a normal distribution is the standard deviation, which is a measure of how spread out the values are (Figure 3). Steeper bell curves have smaller standard deviations and a smaller spread of values. Flatter bell curves have a larger standard deviation and a larger spread of values.

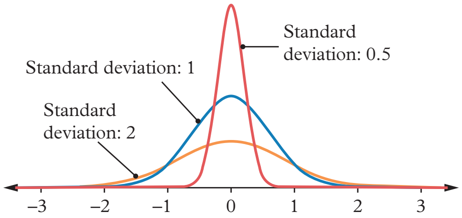


FIGURE 3 The effect of changing the standard deviation on the shape of normal curves

The experimental standard deviation s is defined in mathematical terms by the formula:

$$s = \sqrt{\frac{\Sigma(x - \bar{x})^2}{n - 1}}$$

where $\bar{x}$ is the mean, x is a measured value, n is the number of measurements and Σ (capital sigma) means to add all the values $(x_\mathrm{i} - \bar{x})^2$ together.

Table 2 shows the results of these calculations on the experimental data given in Table 1.

TABLE 2 Steps for calculating the standard deviation of hydrochloric acid volumes

Trial number	Volume of aqueous hydrochloric acid solution (mL)	$x_i - \bar{x}$ (mL)	$(x_i - \bar{x})^2$ (mL^2)
1	19.92	19.92 – 19.957 = –0.037	0.001344
2	20.04	20.04 – 19.957 = 0.083	0.006944
3	19.86	19.86 – 19.957 = –0.0967	0.009344
4	19.98	19.98 – 19.957 = 0.0233	0.000544
5	20.08	20.08 – 19.957 = 0.12337	0.015211
6	19.86	19.86 – 19.957 = –0.0967	0.009344

$$\begin{aligned} s &= \sqrt{\frac{0.001344 + 0.006944 + 0.009344 + 0.000544 + 0.015211 + 0.009344}{6 - 1}} \\ &= \sqrt{\frac{0.042733}{5}} \\ &= \sqrt{8.5466 \times 10^{-3}} \\ &= 0.09245 \text{ mL} \end{aligned}$$

This means that, on average, each data value deviates from the mean of 19.96 mL by 0.09 mL. The full working out has been shown here, but many calculators and spreadsheet programs also have statistical functions that will calculate these values more quickly and conveniently.

Standard deviation is a useful measurement to help you identify errors, outliers and evaluate the reliability of your results.

How do you calculate absolute uncertainty?

absolute uncertainty
the exact magnitude of difference between the mean and the range of measurements; an indicator of the precision of measurements

Using measures of central tendency helps you to be more certain of your data, but the reality is that you introduce another source of uncertainty through replication. When you report your result, you must therefore consider absolute or percentage uncertainty.

Absolute uncertainty ($\Delta\overline{x}$) quantifies how values deviate around a mean. It is the half-range of the measurements, meaning that it indicates the highest measured value and the lowest measured value. It is calculated using:

$$\Delta\overline{x} = \pm\frac{(x_{\text{max}} - x_{\text{min}})}{2}$$

where x_{max} is the largest value and x_{min} is the smallest value. It is expressed to the same number of decimal places as the value with the fewest significant figures (s.f.).

For single measurements, uncertainty is reported based on the uncertainty of the measurement instrument. What if this value is larger than the absolute uncertainty? Which value should you use?

The rule is: the uncertainty value that you report should be the largest one. This is because we don't want to underestimate the uncertainty of the result. This would make the result appear more precise and reliable than is justified or warranted.

Worked example 1.7A

Calculating the mean and absolute uncertainty

Worked example 1.7A: Watch a video that shows how to solve this problem.

An acid–base titration was conducted using a burette with an uncertainty of ±0.02 mL. The titre volumes obtained were: 5.80 mL, 6.00 mL, 6.05 mL. **Calculate** the mean and the absolute uncertainty and express your answer correctly. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We are being asked to calculate the mean and absolute uncertainty. The question is worth 2 marks, so we must correctly apply the formulas to complete the calculations and express the answers correctly.

Think	Do
Step 2: Select the appropriate formulas and gather any data required.	$\bar{x} = \frac{\Sigma x}{n}$ $\Delta\bar{x} = \pm\frac{(x_{\text{max}} - x_{\text{min}})}{2}$ $x_{\text{max}} = 6.05\,\text{mL}$; $x_{\text{min}} = 5.80\,\text{mL}$; $n = 3$
Step 3: Substitute the known values into the formulas and solve for the mean.	$\bar{x} = \frac{5.80\,\text{mL} + 6.00\,\text{mL} + 6.05\,\text{mL}}{3}$ $= 5.95\,\text{mL}$
Step 4: Substitute the known values into the formulas and solve for the absolute uncertainty. Compare the absolute uncertainty to the uncertainty of the instrument. The larger uncertainty is the one to be reported in the final answer.	$\Delta\bar{x} = \pm\frac{(6.05\,\text{mL} - 5.80\,\text{mL})}{2}$ $= \pm\,0.125\,\text{mL}$ The absolute uncertainty of 0.125 mL is larger than the instrument uncertainty of 0.02 mL.
Step 5: Finalise your answer. Report it as the best estimate (mean), uncertainty, and the correct units. Make sure you use the correct number of significant figures in the mean and the same number of decimal places in the uncertainty.	The measurements are to 3 s.f. The mean is currently expressed also as 3 s.f., so we do not need to round it up or down. Absolute uncertainty must be presented to the same number of decimal places (d.p.), i.e. 2. Thus, 0.125 rounds up to 0.13. The final answer is: 5.95 ± 0.13 mL (1 mark for correct mean; 1 mark for correct absolute uncertainty)

Your turn

The mass of a solid substance was measured using an electronic balance with an uncertainty of ±0.05 grams. The recorded masses were 25.20 g, 25.20 g and 25.15 g. **Calculate** the mean and the absolute uncertainty and express your answer correctly. (2 marks)

How do you calculate percentage uncertainty?

After data is collected, you may need to transform it in some way or apply a formula to it calculate other quantities. For example, if you gather mass, amount (in moles) or volume, you can calculate the number of particles, concentration or molar mass, and compare these calculated values with accepted (true) values.

If you add, subtract, multiply or divide the data, you can propagate uncertainty through your calculations. Often, this can involve different units and measurement instruments, so absolute uncertainty may no longer be useful. You will need to make adjustments to take into account the calculations you have performed, and the value you use may instead be percentage uncertainty.

Percentage uncertainty is calculated by dividing the absolute uncertainty by the observed measurement and multiplying the result by 100 to give a percentage:

$$\text{percentage uncertainty }(\%) = \frac{\text{absolute uncertainty }(\Delta\bar{x})}{\text{measurement }(x)} \times 100\%$$
$$= \frac{\Delta\bar{x}}{x} \times 100\%$$

percentage uncertainty
an indicator of uncertainty in which the range of values for a measurement result (the absolute uncertainty) is expressed as a percentage of the measurement

Worked example 1.7B

Calculating the percentage uncertainty

Worked example 1.7B: Watch a video that shows how to solve this problem.

Calculate the percentage uncertainty of a 0.10 M solution of NaOH with a titre volume of 26.18 mL. The calibrated burette used in the titration had an absolute uncertainty of 0.02 mL. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We are being asked to calculate the percentage uncertainty. The question is worth 1 mark, so we must correctly apply the formula to complete the calculations.
Step 2: Select the appropriate formulas and gather any data required.	percentage uncertainty (%) = $\frac{\Delta\bar{x}}{x} \times 100\%$ $\Delta\bar{x} = \pm\, 0.02\,\text{mL}$, $x = 26.18\,\text{mL}$
Step 3: Substitute the known values into the formula and solve for the percentage uncertainty.	percentage uncertainty (%) = $\frac{0.02\,\text{mL}}{26.18\,\text{mL}} \times 100\%$ $= 0.07639\%$
Step 4: Finalise your answer.	± 0.08% (1 mark)

Your turn

A calibrated burette used in titration has an absolute uncertainty of ± 0.02 mL. It is used in an acid–base titration between sodium hydroxide and acetic acid. The titre volume was 13.25 mL. **Calculate** the percentage uncertainty. (1 mark)

How is uncertainty propagated in calculations?

Study tip

For addition and subtraction, add the **absolute** uncertainties. For multiplication, division and powers, add the **percentage** uncertainties.

When completing calculations involving two or more quantities, the uncertainties of each measurement must be combined. This is called propagation of uncertainty. There are some simple rules that can be used to calculate the final uncertainty:

1 If the measured values are added or subtracted, the absolute uncertainties must be added. For example:

$$5.0 \pm 0.2\,\text{g} - 2.3 \pm 0.2\,\text{g} = 2.7 \pm 0.4\,\text{g}$$

2 If the measured value is multiplied or divided, the percentage uncertainties must be added together. For example:

$$3\,\text{g} + 4\% \times 5.2\,\text{g} + 2\% = 15.6\,\text{g} + 6\%$$

3 If powers are applied to the measured value, the percentage uncertainties must be added together. For example:

$$(5.0\,\text{g} \pm 2\%)^3 = (5.0\,\text{g} \pm 2\%) \times (5.0\,\text{g} \pm 2\%) \times (5.0\,\text{g} \pm 2\%) = 125\,\text{g} \pm 6\%$$

Worked example 1.7C

Calculating absolute uncertainty involving subtraction or addition

Worked example 1.7C: Watch a video that shows how to solve this problem.

During a titration, the following titres were recorded for a 0.10 M solution of HCl:

Initial titre = 6.05 ± 0.02 mL
Final titre = 23.35 ± 0.02 mL

Calculate the difference in volume and the absolute uncertainty, and report it correctly. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We are being asked to calculate the difference in volume and absolute uncertainty. The question is worth 2 marks, so we must complete the calculations and express the answer correctly.
Step 2: Find the difference in volume.	difference in volume = 23.35 mL − 6.05 mL = 17.30 mL
Step 3: Determine how to propagate the uncertainties.	We are dealing with subtraction, so we need to add the absolute uncertainties.
Step 4: Solve for the final uncertainty by adding the absolute uncertainties.	0.02 mL + 0.02 mL = 0.04 mL
Step 5: Finalise your answer. Report it as the best estimate, uncertainty and the correct units. Make sure you use the correct number of significant figures in the mean and the same number of decimal places in the uncertainty.	17.30 ± 0.04 mL (1 mark for correct volume; 1 mark for correct uncertainty)

Your turn

A chemistry student prepared three solutions using variable pieces of glassware to measure the required quantities and reported their volumes:

Solution A: 25.00 ± 0.02 mL
Solution B: 36.1 ± 0.1 mL
Solution C: 48.2 ± 0.2 mL

Calculate the total volume of the three solutions and **determine** the absolute uncertainty in the total volume, reporting it correctly. (2 marks)

Worked example 1.7D

Calculating percentage uncertainty involving division

Worked example 1.7D: Watch a video that shows how to solve this problem.

Calculate the concentration in g L^{-1} when 15.7 ± 0.1 g is fully dissolved in a flask of water to make 2.0 ± 0.1 L of solution. Report your answer correctly. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We are being asked to calculate the concentration and percentage uncertainty. The question is worth 2 marks, so we must complete the calculations and express the answer correctly.
Step 2: Find the concentration using $c = \frac{m}{V}$.	$c = \frac{15.7\,\text{g}}{2.0\,\text{L}}$ $= 7.85\,\text{g L}^{-1}$
Step 3: Determine how to propagate the uncertainties.	We are dealing with division, so we need to add the percentage uncertainties. percentage uncertainty (%) = $\frac{\Delta\bar{x}}{x} \times 100\%$ For the mass: $\bar{x} = 0.1$ g, $x = 15.7$ g For the volume: $\bar{x} = 0.1$ L, $x = 2.0$ L

Think	Do
Step 4: Substitute the known values into the formula and solve for the percentage uncertainties.	For the mass: percentage uncertainty (%) $= \frac{0.1\,\text{g}}{15.7\,\text{g}} \times 100\%$ $= 0.6369\%$ For the volume: percentage uncertainty (%) $= \frac{0.1\,\text{L}}{2.0\,\text{L}} \times 100\%$ $= 5.0\%$
Step 5: Solve for the final uncertainty by adding the percentage uncertainties.	$0.6369 + 5.0 = 5.6369\%$ This can also be converted back into an absolute uncertainty, by multiplying the percentage by the calculated concentration. $5.6369\,\% \times 7.85\,\text{g L}^{-1} = 0.4425\,\text{g L}^{-1}$
Step 4: Finalise your answer. Report it as the best estimate, uncertainty and the correct units. Make sure you use the correct number of significant figures in the mean and the same number of decimal places in the uncertainty.	$7.9 \pm 0.4\,\text{g L}^{-1}$ or $7.9\,\text{g L}^{-1} \pm 5.6\%$ (1 mark for correct concentration; 1 mark for correct uncertainty)

Your turn

Calculate the total concentration in grams per litre (g L^{-1}) when 28.4 ± 0.2 g of solute is fully dissolved in a flask of water to make 3.5 ± 0.1 L of solution. Report your answer correctly. (2 marks)

How do you calculate percentage error?

percentage error
the percentage difference between the accepted (true or theoretical) value and the measured (experimental) value

So far, we have talked about uncertainties, which are a way of expressing the level of precision of measurements. To indicate the accuracy of experimental results, we use **percentage error**. This is the percentage difference between the measured experimental value and the known true (or theoretical) value:

$$\text{percentage error (\%)} = \left|\frac{\text{measured value} - \text{true value}}{\text{true value}}\right| \times 100\%$$

The straight lines (|, modulus signs) in the equation indicate the "absolute value", which means the sign (+/–) of the answer is ignored. Percentage error is reported to one significant figure.

Study tip

Absolute and percentage uncertainty are indicators of precision of measurements, whereas percentage error is an indicator of accuracy.

Study tip

Percentage error is reported to one significant figure.

Worked example 1.7E

Calculating percentage error

Worked example 1.7E: Watch a video that shows how to solve this problem.

In an experiment, the enthalpy of combustion for propanol was determined to be $-1{,}978\,\text{kJ mol}^{-1}$ but the theoretical enthalpy of combustion is $-2{,}021\,\text{kJ mol}^{-1}$. **Calculate** the percentage error for the enthalpy of combustion. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We are being asked to calculate the percentage error. The question is worth 1 mark, so we must correctly apply the formula and complete the calculation.

Think	Do
Step 2: Select the appropriate formulas and gather any data required.	percentage error (%) = $\left\lvert\frac{\text{measured value} - \text{true value}}{\text{true value}}\right\rvert \times 100\%$ measured value = $-1{,}978\,\text{kJ}\,\text{mol}^{-1}$ true value = $-2{,}021\,\text{kJ}\,\text{mol}^{-1}$
Step 3: Substitute the known values into the formulas and solve for the percentage error.	percentage error (%) = $\left\lvert\frac{-1{,}978\,\text{kJ}\,\text{mol}^{-1} - (-2{,}021\,\text{kJ}\,\text{mol}^{-1})}{-2{,}021\,\text{kJ}\,\text{mol}^{-1}}\right\rvert \times 100\%$ $= \left\lvert\frac{43\,\text{kJ}\,\text{mol}^{-1}}{-2{,}021\,\text{kJ}\,\text{mol}^{-1}}\right\rvert \times 100\%$ = 2.128%
Step 4: Finalise your answer. Percentage error is rounded to one significant figure.	2% (1 mark)

Your turn

A student measures the density of an object to be 8.63 g mL^{-1}, but the accepted value is 8.96 g mL^{-1}. **Calculate** the student's percentage error. (1 mark)

Study tip

Generally, we allow a 5% threshold for percentage uncertainty and percentage error. If percentage uncertainty is 5% or lower, we consider the result to be precise. If percentage error is 5% or lower, we consider the result to be accurate.

How can you present your results?

After measurements are collected and data is processed, these pieces must be presented in a useful way. It is important to consider the best way to display the data. The information needs to be clear and easy to read, and relationships between variables should be straightforward to identify. Common types of data representation used in chemistry include summary tables and scatterplots.

Summary tables

Tables are often used to summarise and organise data. While a data table states what has been measured and presents the raw data, a summary table can bring together processed data. A summary table should have a heading, and column and row headings. Where quantities are being reported, appropriate units should appear in the headings.

Tables are best where there is a relatively small amount of data and the values are individually important, as shown in Table 3.

TABLE 3 Change in average reaction temperature with time

Time (min)	Average temperature (°C)
5	43.5
10	37.8
15	35.7
20	24.2

Scatterplots

Graphs are often more useful than summary tables when investigating trends in data. They are visual displays of how one variable changes either over time or in response to changes in another variable. x–y scatterplots are often used where there is a continuous trend in a measured value as one value is altered in a systematic way.

For example, the change in temperature of honey as it is heated over time can be shown in a scatterplot.

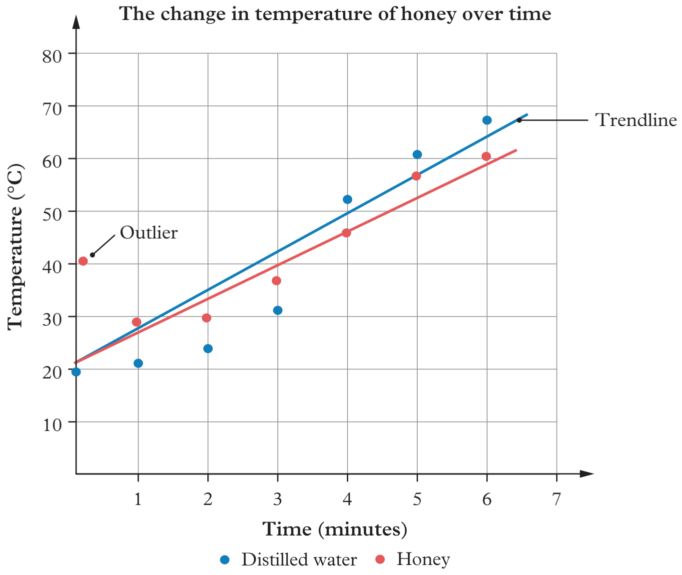


FIGURE 4 An *x*–*y* scatterplot showing how the temperature of a reaction (the dependent variable) changes with time (the independent variable).

All graphs should include the following:

- a graph title that succinctly describes what the graph is showing (typically includes the independent variable (IV) and dependent variable (DV))
- clearly labelled axes, including units of measurement
- equally spaced units of measurement along the axes (scaling)
- axes that start at zero (where possible)
- data plotted within the confines of each axis
- distinguishing symbols, colours or keys when more than one dataset is plotted on a single graph
- measurement uncertainty (where relevant) as error bars (see below)
- a trendline that fits between the error bars (see below).

By analysing the shape of your graphs, you can construct arguments about the relationship between variables.

Study tip

The trendline can allow you to predict values beyond and between the set of data points collected. This is called extrapolation and interpolation, respectfully. You will learn about this later.

How do you graph linear relationships?

Graphing the (x, y) pairs of values can visually reveal the relationship between the variables. Where there is a linear relationship, a **trendline** (or line of best fit) can be added to the graph. This is a line drawn through the middle of the plotted data so that the points are evenly spread on either side of the line. It does not connect all the data points on a graph. You can see two trendlines in Figure 4.

It is rare for the data to be perfectly linear, so drawing a trendline allows you to identify the proportionality relationship between the two variables. A straight trendline can be mathematically represented by a linear equation in the general form of:

$$y = mx + c$$

where x and y are the IV and DV, respectfully, m is the **gradient** or the slope of the line and c is the y-intercept or the point at which the line cuts the y-axis at $x = 0$.

trendline (line of best fit)
a line drawn on a graph joining as many points as possible and showing the general direction of the data; should be drawn with, approximately, an equal number of points above and below the line

gradient
the slope of a graph

Gradient

The slope of any straight line is defined as the change in y divided by the change in x:

$$\begin{aligned} m &= \frac{\text{change in } y}{\text{change in } x} \\ &= \frac{\Delta y}{\Delta x} \\ &= \frac{y_2 - y_1}{x_2 - x_1} \end{aligned}$$

The slope of the trendline can therefore be calculated by choosing two points on the line, (x_1, y_1) and (x_2, y_2). Let's consider the trendline in Figure 5. The two points are (0, 1) and (2, 4). We can calculate m:

$$\begin{aligned} m &= \frac{4 - 1}{2 - 0} \\ &= \frac{3}{2} \end{aligned}$$

Study tip

You may find it easier to remember the gradient as "rise over run", where "rise" refers to the y-axis and "run" refers to the x-axis.

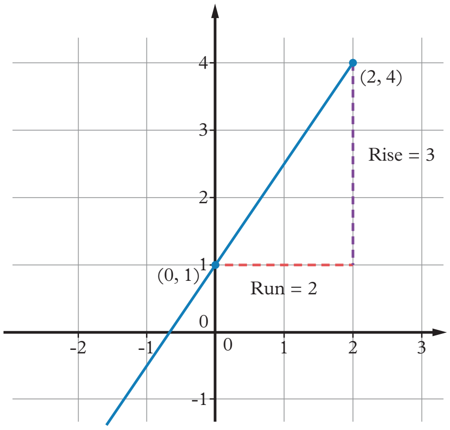


FIGURE 5 An example of linear data

There are three important points to remember:

- don't force the line to go through the origin; it should go through the middle of the data points, with as many points above it as below it
- you need to use the x and y coordinates of points on the trendline to calculate the gradient
- use points as widely separated as possible to get the most accurate value for the gradient.

The direction in which the gradient slopes tells you about the relationship between the variables. When there is an incline in the positive direction of the x-axis, the gradient is a positive value and this indicates that as the IV increases, the DV also increases. The opposite is also true: where there is a decline in the positive direction of the x-axis, the gradient of the line will have a negative value and this indicates that as the IV increases, the DV decreases.

This can also be described using **correlations**. Variables can be positively or negatively correlated. Positive correlation is when the value of one variable increases as the other variable increases (gradient slopes upwards). Negative correlation is when the value of one variable decreases, as the other variable increases (gradient slopes downwards). Figure 5 displays a positive correlation.

correlation
a link between a change in the independent variable and a change in the dependent variable; this does not mean that the changing independent variable caused the change in the dependent variable

y-intercept

The y-intercept is the point on the graph where the trendline intercepts or cuts the y-axis at $x = 0$. This data point can be found in two ways:

1 Extending of the trendline and reading the y-intercept from the graph. Let's consider Figure 5 again. We have found that the gradient m is $\frac{3}{2}$. We can already tell that the y-intercept is +1, so $y = \frac{3}{2}x + 1$.

2 Calculating the y-intercept using the equation for the trendline and using algebra to find c. From Figure 5, we have the equation $y = \frac{3}{2}x + c$. To find c, we can select any point along the line. Let's use (2, 4). If we substitute these x and y coordinates into the equation for the trendline, we get $4 = \frac{3}{2}(2)+c$ or $4 - 3 = c$. Therefore, $c = 1$.

How can you infer data by extrapolation and interpolation?

If the equation for a trendline is known, it is possible to determine unknown values. Extending and reading a graph beyond the last plotted point is called **extrapolation**. Inferring a reading between plotted points is called **interpolation**.

extrapolation
the prediction of values beyond the range of data points by extending the trendline

interpolation
the prediction of values between data points using a trendline

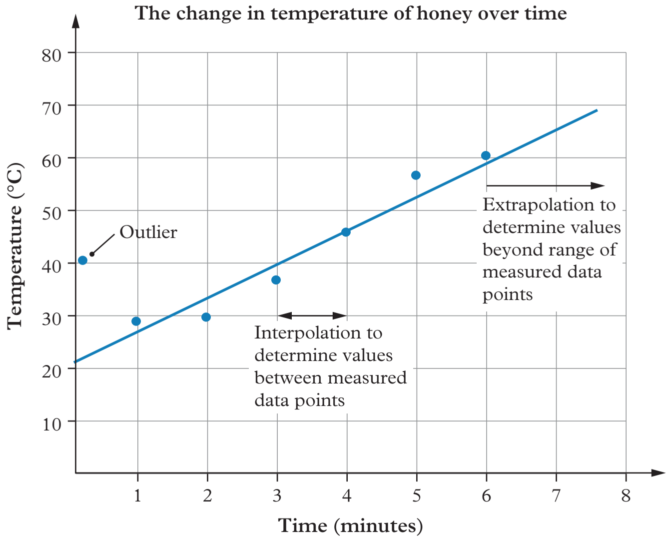


FIGURE 6 Extrapolation and interpolation of data to determine unknown values

In Figure 6, the data points are only at 0, 1, 2, 3, 4, 5 and 6 minutes. You can extrapolate the trendline past 6 minutes to 7 minutes and read off the temperature as 65°C. You can also interpolate the data to read off values between points. For example, at 1.5 minutes, the temperature is 30°C.

In QCE Chemistry, you will use standard curves to find unknown values. This involves interpolation. Worked example 1.7F shows you how to construct a standard curve (scatterplot with a trendline) for an atomic absorption spectroscopy experiment, and then use interpolation to find the concentration of an unknown sample.

Worked example 1.7F

Constructing and analysing linear scatterplots

Worked example 1.7F: Watch a video that shows how to solve this problem.

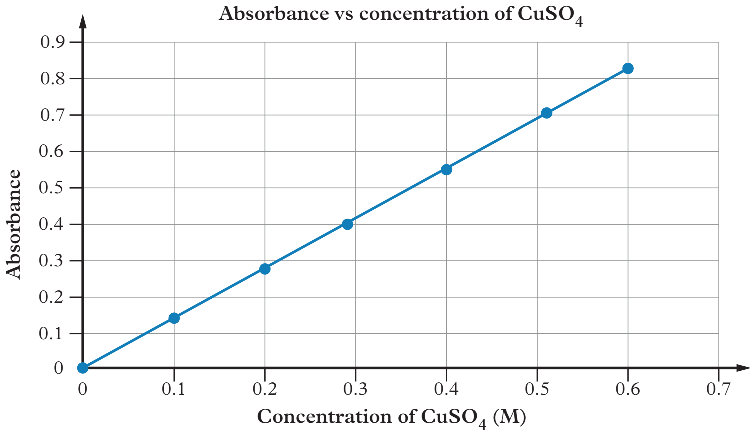
A student wants to determine the concentration of a copper sulfate solution. They use atomic absorption spectroscopy to measure the concentration of various copper sulfate standard solutions, as well as their sample.

TABLE 4 Data obtained from the experiment

Concentration of copper sulfate (M)	Absorbance
0	0.000
0.1007	0.157

Concentration of copper sulfate (M)	Absorbance
0.2009	0.290
0.2902	0.400
0.3999	0.550
0.5094	0.705
0.5999	0.825
unknown	0.460

a **Sketch** a scatterplot for the copper sulfate standard solutions. (3 marks)
b **Calculate** the gradient, *m*, of the trendline. (1 mark)
c **Derive** the equation for the trendline. (2 marks)
d **Calculate** the concentration, in M, of the unknown solution of copper sulfate solution. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Sketch" means to represent using a diagram or graph. "Calculate" means to determine or find a number or answer by using mathematical processes. "Derive" means manipulate a mathematical relationship to give a new equation. These questions are worth a variety of marks. We must graph the data, find the equation for the trendline and use it to find the unknown value.
Step 2: For part **a,** sketch the graph by drawing and labelling the axes with the IV (concentration) on the *x*-axis and DV (absorbance) on the *y*-axis – this includes units! Plot each data point and draw a trendline.	 (1 mark for correct axes and labels; 1 mark for correct data points; 1 mark for trendline)
Step 3: For part **b**, find the gradient by selecting the appropriate equation. Determine the change in *y* and change in *x* for two widely separated points on the line.	$m = \frac{y_2 - y_1}{x_2 - x_1}$ using the two points (0.4, 0.55) and (0.2, 0.28). $m = \frac{0.55 - 0.28}{0.4\,M - 0.2\,M}$ (1 mark) $= 1.35\,M^{-1}$
Step 4: For part **c**, read the intercept off the *y*-axis and include the units. Alternatively, we can calculate this more accurately using the *m* value and one (*x*, *y*) datapoint.	From the graph, $c = 0$. From calculating, using $m = 1.35$ and (0.4, 0.55). $0.55 = 1.35(0.4) + c$ (1 mark) $c = 0.01$ Putting it altogether: $y = 1.35x + 0.01$ (1 mark)
Step 5: For part **d**, use the equation for the trendline and substitute in the value for absorbance (*y*) to solve for concentration (*x*).	$0.460 = 1.35x + 0.01$ $0.450 = 1.35x$ $x = 0.333\,M$ (1 mark)

Your turn

A student conducted an experiment to investigate the density of copper by measuring the masses of samples of known volume. The results were tabulated. A linear graph can then be used to determine the volume of an irregularly shaped copper ornament with a mass of 113.25 g.

TABLE 5 Results from experiment measuring mass and volume of Cu

Volume (cm^3)	Mass (g)
2	17.83
3	26.84
4	35.64
8	71.66
16	143.11
unknown	113.25

a **Sketch** a scatterplot for the mass of Cu as a function of volume, including a linear trendline. (3 marks)

b **Calculate** the gradient of the trendline. (1 mark)

c **Derive** the equation for the trendline. (2 marks)

d **Calculate** the volume of 113.25 g of Cu metal. (1 mark)

How do you graph non-linear relationships?

linearising
the process of transforming non-linear data by applying a mathematical function to one of the variables so that the relationship between the variables becomes closer to a straight line

Not all relationships between variables are linear. At a glance, these relationships may appear more complex to analyse. However, sometimes a linear relationship can be found by applying a mathematical operation to one or both variables. This is called **linearising** the data.

One example of an experiment in which linearising is necessary is when the rate of reaction is measured in response to a change in concentration of a reactant. This is an exponential relationship given by the equation $y = e^x$.

When x is plotted against y, the graph is non-linear (Figure 7A). To linearise the relationship, the function must be undone. For an exponential relationship, we take the natural logarithm (ln or $\log_e$) of both sides to "undo" the exponential: $\ln y = x$. If the relationship is indeed exponential, plotting x against $\ln y$ should give a linear relationship (Figure 7B).

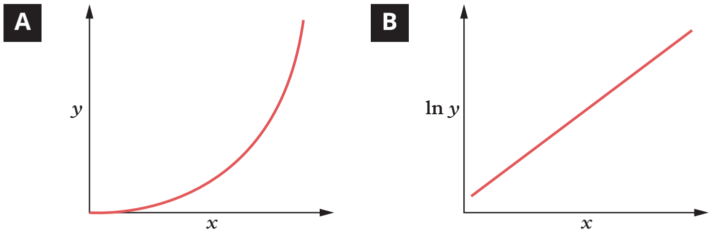
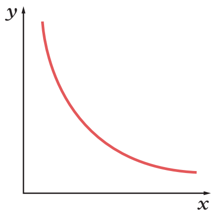
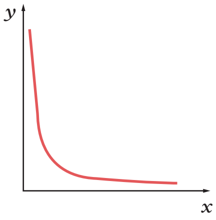
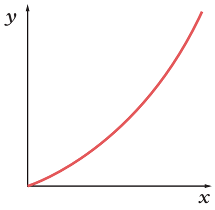
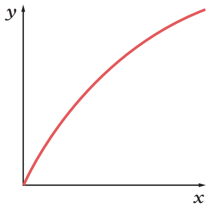
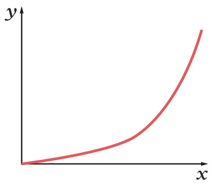
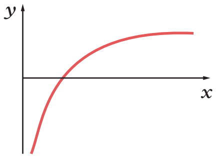


FIGURE 7 (A) The x versus y graph of the data does not give a straight line; (B) The straight line indicates that this is an exponential relationship.

The trendline for the linearised graph can be determined using the methods described earlier. This can enable you to extrapolate or interpolate the data to predict values not captured by direct measurement.

The same strategy can be applied to other non-linear relationships. Table 6 summarises the functions that must be applied to different relationships.

TABLE 6 Mathematical functions to apply to linearise different non-linear relationships

Type of relationship	Relationship	What to graph
Inverse power 	$y = \frac{1}{x}$	y vs $\frac{1}{x}$
Inverse-square 	$y = \frac{1}{x^2}$	y vs $\frac{1}{x^2}$
Square or parabolic 	$y = x^2$	y vs x^2
Square root 	$y = \sqrt{x}$	y vs $\sqrt{x}$
Exponential 	$y = e^x$	$\ln y$ vs x
Logarithmic 	$y = 10^x$	$\log_{10} y$ vs x

How are error bars used to represent uncertainty?

The measurement uncertainty associated with a particular mean value can be easily displayed on a graph using error bars. These are vertical or horizontal lines that are added to each data point that represent the absolute uncertainty at that point. For example, Figure 8 shows a student experiment that investigates the effect of increasing temperature on the solubility of a salt (in grams of salt per 100 g of water). The vertical error bars reflect a small degree of uncertainty for the dependent variable measured which in this case, is the solubility of the salt.

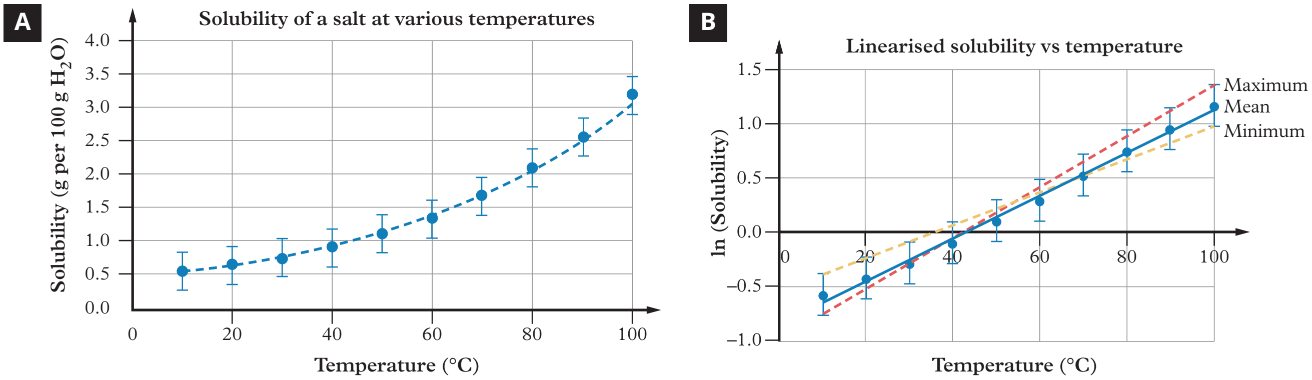


FIGURE 8 (A) Vertical error bars show the zone of uncertainty for the solubilities of salt in water (g per 100 g H_2O). (B) Maximum and minimum trendlines can be added to the graph to show the uncertainty in the gradient for the linearised curve.

maximum trendline
a line of best fit of maximum gradient within the bounds of the error bars

minimum trendline
a line of best fit of minimum gradient within the bounds of the error bars

Figure 8A shows vertical error bars for the measurement of solubility of a salt (y-axis). Figure 8B shows the linearised solubility over time. The gradient of the trendline (or mean trendline) is 0.0195. Two other trendlines have been added.

- The **maximum trendline** has the highest possible gradient. It connects the bottom of the error bar of the first data point (10, –0.6500) with the top of the error bar for the last data point (100, 1.4125) but must not go beyond any of the error bars in between.
- The **minimum trendline** has the smallest possible gradient. It connects the top of the error bar of first data point (10, –0.3500) with the bottom of the error bar for the last data point (100, 0.9625) but must not go beyond any of the error bars in between.

To calculate the maximum and minimum gradients, it is good practice to use points spaced as wide apart as possible. Ideally, try to use the full line between $x = 10°C$ and $x = 100°C$.

- The maximum trendline (red) gradient is: $\frac{(1.4125 - (-0.6500))}{(100 - 10)} = 0.023$.
- The minimum trendline (yellow) gradient is: $\frac{(0.9625 - (0.3500))}{(100 - 10)} = 0.015$.

The uncertainty in the gradient is calculated using the following formula:

$$\begin{aligned}\Delta\bar{x} &= \pm\frac{(x_{\text{max}} - x_{\text{min}})}{2}\\ &= \pm\frac{(0.023 - 0.015)}{2}\\ &= \pm\, 0.0042\,\text{g}\,\text{mL}^{-1}\,°\text{C}^{-1}\end{aligned}$$

Thus, the gradient can be stated using absolute uncertainty as $m = 0.0195 \pm 0.0042$.

Worked example 1.7G

Analysing minimum and maximum trendlines

Worked example 1.7G: Watch a video that shows how to solve this problem.

Calculate the uncertainty of the slope of the trendline using the maximum and minimum best fit lines and express the gradient correctly. (2 marks)

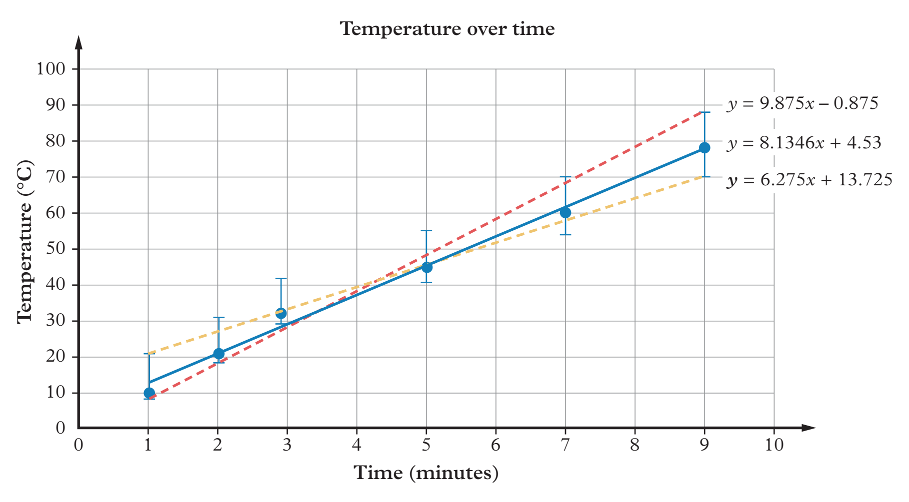


FIGURE 9 Change in temperature over time

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We are being asked to calculate the uncertainty of the gradient. The question is worth 2 marks, so we must correctly apply the formulas to complete the calculations and express the answer correctly.
Step 2: Select the appropriate formulas and gather any data required.	$\Delta\bar{x} = \pm\frac{(x_{max} - x_{min})}{2}$ x_{max} is the maximum trendline gradient = 9.875 x_{min} is the minimum trendline gradient = 6.275
Step 3: Substitute the known values into the formula and solve for the uncertainty of the gradients.	$\Delta\bar{x} = \pm\frac{(9.875\,°C\,min^{-1} - 6.275\,°C\,min^{-1})}{2}$ (1 mark) $= \pm 1.8\,°C\,min^{-1}$
Step 4: Finalise your answer. Report it as the best estimate, uncertainty and the correct units. Make sure you use the correct number of significant figures in the mean and the same number of decimal places in the uncertainty.	$8.135 \pm 1.800\,°C\,min^{-1}$ (1 mark)

Your turn

Given the equations of the mean, maximum, and minimum trendlines below, **calculate** the uncertainty of the gradient and express the gradient correctly. (2 marks)

Mean trendline: $y = 1.98x + 47.9$
Maximum trendline: $y = 2.35x + 53.2$
Minimum trendline: $y = 1.18x + 51.6$

How do you interpret graphs?

Scatterplots allow scientists to describe the relationship between the IV and DV. The description of the graph should include the:

- IV and DV
- type of correlation shown by the graph
- the shape of the graph (type of relationship, e.g. linear, exponential, etc).

causation
when a change in a single variable causes a change in a second variable

When interpreting the data shown on a graph, it is important to remember that correlation does not imply **causation**; you may not be able to make the claim that an IV *is the result of* a change in the DV. Instead, you can describe them as being correlated – a change in the IV *is related to* a change in the DV but does not necessarily cause it to change.

Check your learning 1.7

Check your learning 1.7: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define** the following terms: absolute uncertainty, percentage uncertainty, percentage error. (3 marks)

2 The titres of an acid–base titration were measured in triplicate and found to be: 24.3 mL, 24.7 mL, 23.8 mL. **Calculate** the:

- a average titre volume (1 mark)
- b absolute uncertainty of the mean titre value (1 mark)
- c percentage uncertainty of the titre value. (1 mark)

3 A student obtained a value of 3.74 mL s^{-1} for the rate of hydrogen production in a metal–acid displacement reaction, whereas the accepted value for the same reaction is 3.85 mL s^{-1}. **Calculate** the percentage error. (1 mark)

4 **Explain** what it means to “linearise” a relationship. (2 marks)

Analytical processes

5 **Determine** the mean, mode and median of the following dataset. (3 marks)

12.5 mg, 12.7 mg, 13.1 mg, 13.4 mg, 13.4 mg, 13.8 mg, 13.9 mg, 14.8 mg

6 **Contrast** positive and negative correlation. **Sketch** two graphs to support your answer. (3 marks)

7 **Contrast** correlation and causation. (2 marks)

8 A student combines two liquid reactants they have measured in a measuring cylinder. Reactant A has a volume of 15.3 ± 0.5 mL and reactant B has a volume of 18.5 ± 0.5 mL.

- a **Determine** the combined volume of the reactants. (1 mark)
- b **Calculate** the absolute uncertainty of the combined volume of reactants. (1 mark)

9 A student wants to determine the volume of a rectangular container. They use a ruler with an absolute uncertainty of ±0.2 cm to obtain the following measurements: length 4.5 cm, width 2.3 cm, height 1.0 cm.

- a **Analyse** the information to **determine** the percentage uncertainty of each measurement. (3 marks)
- b To calculate the volume, the student multiplies the length, width and height of the container. **Determine** the volume. (1 mark)
- c **Determine** the percentage uncertainty of the volume of the container. (1 mark)
- d If the accepted true volume of the container is 10.7 cm^3, **determine** the percentage error of the experiment. (1 mark)

10 A chemistry class sets up an apparatus to measure how changes in temperature affect a sample of air trapped in a long thin glass tube. The air is trapped at the upper end by a droplet of oil which can move up or down as temperature changes. The assembly is immersed in hot water, allowing the air to reach the same temperature as the water. The temperature and the length of the sample of air is measured, and recorded. Length is proportional to volume. The hot water is allowed to cool gradually, and further measurements are made at a number of temperatures as it cools (Table 7).

TABLE 7 Results from experiment measuring length of air sample at various temperatures

Temperature (°C)	70	65	60	55	50	45	40	35	30
Length (cm)	35.3	34.7	34.1	33.8	33.4	32.5	32.1	31.5	31.1

a **Construct** a graph of the results on Excel. Draw a linear trendline, displaying the equation, R^2 and appropriate error bars. (3 marks)

b **Determine** the length of the column of air that would be expected if the assembly could be cooled to –100°C, using the equation for the trendline in part **a**. (1 mark)

11 A student experiment investigated how the pressure of air in a syringe sealed at one end changed as the volume of the syringe was decreased by pushing in the plunger. Three trials were conducted for each volume. The results are shown in Table 8.

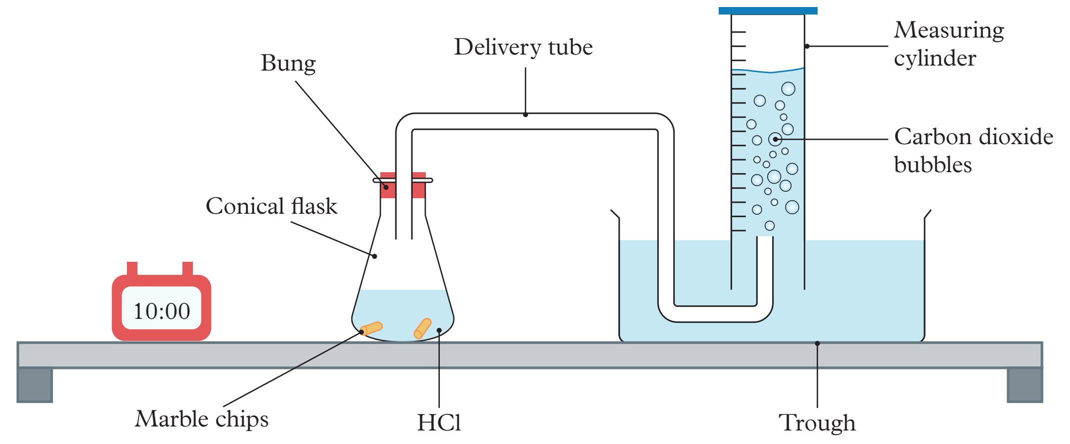


FIGURE 10 Set-up for the student experiment

TABLE 8 Results for the experiment

Volume of air (mL)	Pressure (kPa), trial 1	Pressure (kPa), trial 2	Pressure (kPa), trial 3
60	101	101	101
55	109	111	110
50	121	118	121
45	133	130	133
40	147	143	148
35	167	160	168
30	191	186	193

a **Calculate** the mean pressures for each volume value. (7 marks)

b **Sketch** a scatterplot for the average data using Excel. (2 marks)

c **Compare** the shape of the scatterplot in part **b** to the shapes of various non-linear functions shown in Table 6. **Determine** which type of relationship best matches the shape of the scatterplot, and what variable should be graphed to linearise the data. (2 marks)

d **Calculate** $\frac{1}{V}$ for each volume in the table. (7 marks)

e **Sketch** a scatterplot of P as a function of $\frac{1}{V}$ using Excel, including the trendline, equation for the trendline, and R^2. Set the intercept at (0, 0). (4 marks)

Knowledge utilisation

12 A decomposition reaction of reagent X was carried out and the concentration of reagent over time was monitored and recorded in Table 9.

TABLE 9 Results from the experiment

Time (s)	Concentration of reagent X (M)
4.00×10^2	3.30×10^{-3}
1.00×10^3	2.50×10^{-3}
2.00×10^3	2.00×10^{-3}
3.00×10^3	1.50×10^{-3}
5.00×10^3	5.00×10^{-4}

a **Construct** a suitable scatterplot for the results and **determine** the equation of the trendline that best represents the relationship between the decomposition of X over time. (5 marks)

b **Predict** the initial concentration of reagent X (time = 0 s) and the concentration at 4.0×10^3 s. Include units. (1 mark)

Lesson 1.8

Evaluating evidence

Key ideas

→ Evidence is evaluated to determine its validity and reliability.

→ Goodness of fit (R^2) provides information about the precision of the data.

→ Random and systematic errors can be identified by assessing the trendline of a graph.

→ Not all secondary sources are reliable; they must be evaluated before use.

Science inquiry skills

Learning intentions and success criteria

This lesson provides support for the following science inquiry skills:

- identify and implement strategies to manage risks, ethics and environmental impact, e.g.
  - acknowledgement of sources and referencing
- select and construct appropriate representations to present data and communicate findings, e.g.
  - discriminate between precision and accuracy
  - identify that all measurements have limits to the precision and accuracy that must be considered when evaluating experimental results
  - draw and interpret best-fit lines or curves through data points, including evaluating when it can and cannot be considered as a linear function

  - identify that quantitative data obtained from measurements is associated with random error/measurement uncertainties
- use data and reasoning to discuss and evaluate the validity and reliability of evidence, e.g.
  - discuss ways in which measurement error, instrumental uncertainty, the nature of the methodology or other factors influence uncertainty and limitations in the data
  - evaluate information sources and compare ideas, information and opinions presented within and between texts, considering aspects such as bias, appropriateness and reasonableness
  - compare findings to theoretical models or expected values
  - discriminate between validity and reliability
- appreciate the role of peer review in scientific research.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

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Lesson 1.9

Communicating scientifically

Key ideas

→ Different communication conventions are used depending on the target audience.

→ Evidence-based arguments bring together scientific ideas and primary and/or secondary data obtained from scientific investigation.

→ Sources can be acknowledged using in-text referencing and a bibliography.

Science inquiry skills

Learning intentions and success criteria

This lesson provides support for the following science inquiry skills:

- select, synthesise and use evidence to construct scientific arguments and draw conclusions
- communicate to specific audiences and for specific purposes using appropriate language, nomenclature, genres and modes
- acknowledge sources of information and use standard scientific referencing conventions

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Lesson 1.10

Preparing for your data test

Key ideas

→ The data test is an assessment task where you respond to items using qualitative data and/or quantitative data derived from practicals, activities or case studies.

→ In your data test, you will apply your understanding, analyse data and interpret evidence.

Learning intentions and success criteria

Assessment objectives

This lesson provides support for achieving the assessment outcomes for the data test:

Objective 2. Apply understanding of properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Objective 3. Analyse data about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Objective 4. Interpret evidence about properties and structure of atoms and materials, and chemical reactants, products and energy change.

Source: modified from *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Worked examples

This lesson is supported by the following Worked example:

- **Worked example 1.10A** Responding to the data test

oxforddigital **This lesson is available on Oxford Digital.**

Lesson 1.11

Conducting your student experiment

Key ideas

→ The student experiment is an assessment task where you modify (i.e. refine, extend or redirect) an experiment to address your own related hypothesis or question.

→ You will research and plan your experiment, analyse, interpret and evaluate your evidence, and communicate your findings.

Assessment objectives

Learning intentions and success criteria

This lesson provides support for achieving the assessment outcomes for the student experiment:

Objective 1. Describe ideas and findings about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Objective 2. Apply understanding of properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Objective 3. Analyse data about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Objective 4. Interpret evidence about properties and structure of atoms and materials, and chemical reactants, products and energy change.

Objective 5. Evaluate processes, claims and conclusions about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

Objective 6. Investigate phenomena associated with properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

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Lesson 1.12

Conducting your research investigation

Key ideas

→ The research investigation is an assessment task where you gather evidence related to a research question to evaluate a claim about an issue in chemistry.

Assessment objectives

Learning intentions and success criteria

This lesson provides support for achieving the assessment outcomes for the research investigation:

Objective 1. Describe ideas and findings about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 2. Apply understanding of intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 3. Analyse data about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 4. Interpret evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 5. Evaluate processes, claims and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 6. Investigate phenomena associated with intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

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Lesson 1.13

Preparing for your exams

Key ideas

→ Your external exam requires you to demonstrate your understanding of the QCE Chemistry course by completing multiple-choice and short response questions.

→ Exam questions are typically written using cognitive verbs, which tell you what information you need to provide in your answer to the question.

Learning intentions and success criteria

Assessment objectives

This lesson provides support for achieving the assessment outcomes for the examination:

Objective 1. Describe ideas and findings about properties and structure of atoms and materials, chemical reactions in terms of reactants, products and energy change, intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 2. Apply understanding of properties and structure of atoms and materials, chemical reactions in terms of reactants, products and energy change, intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 3. Analyse data about properties and structure of atoms and materials, chemical reactions in terms of reactants, products and energy change, intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

Objective 4. Interpret evidence about properties and structure of atoms and materials, chemical reactions in terms of reactants, products and energy change, intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

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MODULE 1

Lesson 1.14

Review: Chemistry toolkit

Module summary

1.1
- In QCE Chemistry, you will gain a variety of science understanding and inquiry skills that will help you succeed in assessments, as a scientific professional, and as a science citizen.

1.2
- First Nations peoples have longstanding scientific knowledge.
- First Nations peoples have developed knowledge about the world by observing using all the senses, predicting and hypothesising, testing (trial and error), and making generalisations within specific contexts such as the use of food, natural materials, navigation and sustainability of the environment.
- Correctly acknowledging cultural and/or language groups, avoiding Eurocentrism and critically evaluating sources of information can help you to respectfully engage with First Nations perspectives in QCE Chemistry.

1.3
- The scientific method is a circular process that involves making observations, formulating a hypothesis which will often lead to performing experiments or simulations, developing models or theorems, retesting and trailing.
- Research questions define the scope of an investigation. They can be used to develop hypotheses which predict the outcome of the investigation.

1.4
- A method outlines the steps followed in an experiment and lists all of the materials and equipment used.
- Valid and reliable measurements can be obtained by carefully designing your investigation to collect sufficient data and minimise errors.
- All measurements include errors or uncertainties, either systematic or random. It is important to consider these and implement strategies to minimise their effects when you plan your experiments.

1.5
- Good laboratory practices and lab safety allows you to identify risks and take measures to control them so that you and those sharing the lab space with you stay safe during scientific investigations.
- Conducting experiments ethically involves considering the impacts of your investigation beyond the laboratory.

1.6
- Single experimental measurements are reported using best estimates, indicators of measurement uncertainty and units.
- Scientific notation is used to easily express extremely large or small values.
- Significant figures are digits in a number that are known with certainty plus the first digit that is uncertain.
- All measurements, information and observations should be recorded in your logbook.

1.7
- Data is processed and analysed to identify trends, patterns, relationships, limitations and uncertainty.
- Absolute and percentage uncertainty give an idea of the precision of measurements.
- Percentage error gives an idea of the accuracy of measurements.
- Data can be summarised in a variety of ways, such as in a table, a scatterplot or a scientific diagram.

1.8
- Evidence is evaluated to determine its validity and reliability.
- Goodness of fit (R^2) provides information about the precision of the data.
- Random and systematic errors can be identified by assessing the trendline of a graph.
- Not all secondary sources are reliable; they must be evaluated before use.

1.9
- Different communication conventions are used depending on the target audience.
- Evidence-based arguments bring together scientific ideas and primary and/or secondary data obtained from scientific investigation.
- Sources can be acknowledged using in-text referencing and a bibliography.

1.10
- The data test is an assessment task where you respond to items using qualitative data and/or quantitative data derived from practicals, activities or case studies.
- In your data test, you will apply your understanding, analyse data and interpret evidence.

1.11
- Your student experiment requires you to modify an experiment you have conducted in Units 1 and 2 to address your own related hypothesis or question.
- You will research and plan your experiment, analyse, interpret and evaluate your evidence and communicate your findings.

1.12
- The research investigation is an assessment task where you gather evidence related to a research question to evaluate a claim about an issue in chemistry.

1.13
- Your external exam requires you to demonstrate your understanding of the QCE Chemistry course by completing multiple-choice and short response questions.
- Exam questions are typically written using cognitive verbs, which tell you what information you need to provide in your answer to the question.

Review questions 1.14A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 How many significant figures are in 0.0034 g?

A 4

B 2

C 5

D 3

2 How close a measurement is to its true value is called

A reliability.

B accuracy.

C precision.

D validity.

3 Two students experimentally determined the concentration of a copper(II) sulfate solution. They repeated their experiment three times. Their results are shown in Table 1.

TABLE 1 Concentrations of copper(II) sulfate solution

Student 1	Student 2
0.893	0.884
0.897	0.882
0.889	0.883

The true value of the copper(II) sulfate solution was 0.893 M. Which student(s) produced results showing a systematic error?

A Student 1

B Student 2

C Both students

D Neither student

4 Consider the research question: "How does concentration of acid affect the rate of hydrogen gas production?" The independent variable is

A concentration of acid.

B time taken to produce gas.

C rate of gas production.

D volume of gas produced.

5 A student adds 4.92 g of sodium chloride to 7.587 g of copper metal. What is the final total mass of the two substances correct to 2 s.f.?

A 13.0 g

B 12.5 g

C 12.50 g

D 12.51 g

6 A piece of metallic magnesium with a mass of 6.431 g was found to have a volume of 3.70 cm^3. A student carried out the following calculation to determine the density: density = $\frac{6.431\text{ g}}{3.70\text{ cm}^3}$. What is the best estimate the student could report for the density of magnesium?

A 1.738 g cm^{-3}

B 1.7 g cm^{-3}

C 1.74 g cm^{-3}

D 1.7381 g cm^{-3}

7 The document that a seller of chemicals in Australia must supply with a chemical is called a
- **A** risk assessment.
- **B** certificate of analysis.
- **C** fitness for use warranty.
- **D** safety data sheet.

8 Which are likely to be reduced when an experiment is repeated several times?
- **A** Both random and systematic errors
- **B** Neither random or systematic errors
- **C** Random errors only
- **D** Systematic errors only

9 Which term is defined as “a value closest to the true value, usually found by taking repeated measurements and averaging”?
- **A** Average
- **B** Accepted value
- **C** Uncertainty
- **D** Best estimate

10 Students are using a poorly calibrated electronic balance to measure the mass of a watch glass. They repeat the measurement three times to be sure of the answer. Their measurements could be described as having
- **A** high accuracy and high precision.
- **B** high accuracy and low precision.
- **C** low accuracy and high precision.
- **D** low accuracy and low precision.

11 How close a measurement is to other replicates in the same experiment is referred to as
- **A** accuracy.
- **B** relevance.
- **C** precision.
- **D** validity.

12 Which of the following expresses the answer to the calculation using the correct number of significant figures?

$$\frac{1.47 \times 10^{-5}}{6.538 \times 10^{-3}}$$

- **A** 2.25×10^{-3}
- **B** 2.225×10^{-3}
- **C** 2.3×10^{-3}
- **D** 2.2248×10^{-3}

13 Which of the following expresses the answer to the calculation using the correct number of significant figures?

$$(0.335 \pm 0.001) + (0.279 \pm 0.002)$$

- **A** 0.614 ± 0.003
- **B** 0.614 ± 0.001
- **C** 0.614 ± 0.002
- **D** 0.614

14 Convert 237.8 metres to millimetres.
- **A** 237,800 mm or 2.378×10^5 mm
- **B** 238,000 mm or 2.38×10^5 mm
- **C** 240,000 mm or 2.4×10^5 mm
- **D** 200,000 mm or 2×10^5 mm

Review questions 1.14B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

15 **Explain** the difference between significant figures and scientific notation. (2 marks)

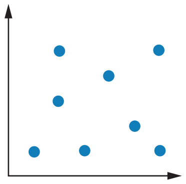
16 **Explain** how a “systematic error” differs from a “random error”. (2 marks)

17 **Explain** what is meant by “propagation of errors” when performing a calculation. (1 mark)

18 **Explain** the reasons for making repeated measurements in an experiment. (2 marks)

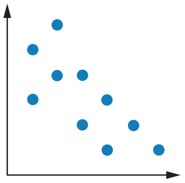
19 **Describe** the correlation (if any) in the following graphs.

a



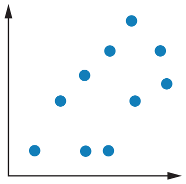
(1 mark)

b



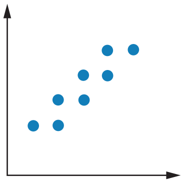
(1 mark)

c



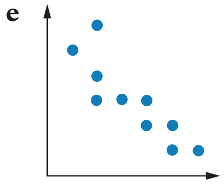
(1 mark)

d



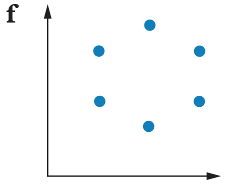
(1 mark)

e



(1 mark)

f



(1 mark)

20 **Calculate** the following to the correct number of significant figures.

- **a** 12.39 + 1.4 (1 mark)
- **b** 0.4329 ÷ 0.031 (1 mark)
- **c** 0.043 – 0.03 (1 mark)
- **d** 340 × 0.3402. (1 mark)

21 Four students weigh a standard 10 g mass on the same balance. All four students achieve a measurement of 8.9 g for the mass of the standard. **Describe** the accuracy and precision of the measurement. (2 marks)

22 A chemist decides to increase the number of replicates in their experiment. **Identify** the type of error they are trying to reduce. **Explain** your answer. (2 marks)

23 In an acid–metal reaction, $46.9 \pm 0.2\,\text{cm}^3$ of hydrogen gas was produced in 20 ± 1 seconds. **Calculate** the rate of reaction in $\text{cm}^3\,\text{s}^{-1}$ and express your answer with an uncertainty correct to the appropriate number of significant figures. (3 marks)

Analytical processes

24 **Determine** the diameter of the bauble shown. **Calculate** the absolute uncertainty of your measurement. Provide reasoning to support the uncertainty you have shown. (3 marks)

FIGURE 1 The width of a bauble measured using a ruler

25 A scientist used an electronic balance that consistently measured mass as 0.05 g above the true value. **Identify** an error that would result from using this electronic balance. **Consider** how the scientist could minimise this error in the future. (2 marks)

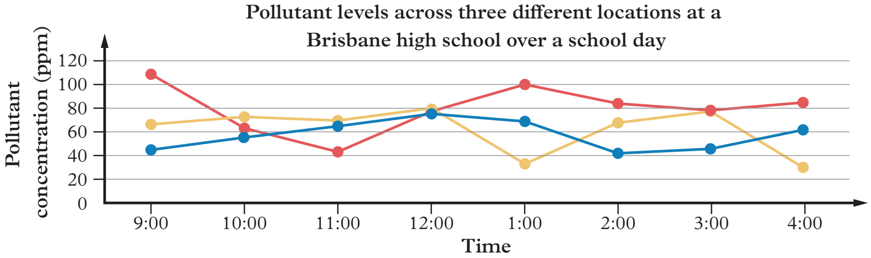
Knowledge utilisation

26 The accepted boiling point of water is 100°C. The following boiling point values were measured by a student during an experiment under standard laboratory conditions:

89°C, 90°C, 91°C, 88°C, 91°C, 95°C, 97°C, 93°C, 91°C, 90°C

Comment on the accuracy and precision of the data obtained. (4 marks)

27 As part of a Year 11 science project, the concentration of air pollutants at different locations around a school located in midtown Brisbane was investigated. Air quality sensors were positioned at varying locations and pollutant levels were recorded over a 4-hour period during a weekday. The graph of hourly pollutant levels is shown in the graph.



● **Library**	45	56	64	76	69	43	47	62
● **Playground**	108	65	45	79	100	85	79	84
● **Classroom**	67	73	69	79	34	68	78	32

FIGURE 2 Results from an experiment measuring pollutant levels in the library, playground and classroom at a Brisbane high school

- **a** **Determine** the time during the school day in which pollutants are at their greatest concentration. (1 mark)
- **b** Briefly **describe** the trend in pollution over the course of a school day and the relationship between pollution and location. (2 marks)
- **c** Provide a hypothesis explaining the trend in location and pollution over the course of the day. (3 marks)
- **d** **Propose** a possible experiment that can test your hypothesis. (4 marks)

28 The density of an aluminium block was experimentally determined to be $2.53\,g\,cm^{-3}$. The true density is $2.70\,g\,cm^{-3}$. **Calculate** the percentage error and **evaluate** the validity of the experimental results obtained if 5% is the cut-off. (3 marks)

29 As part of a student investigation, 12 copper coins were weighed with the recorded masses displayed in Table 2.

TABLE 2 Mass of copper coins

Mass of copper coin (g)		
3.113	3.109	3.049
2.456	2.786	3.053
3.101	3.121	3.024
2.789	3.063	3.098

After further investigation into the dates on each of the coins, it was revealed that:

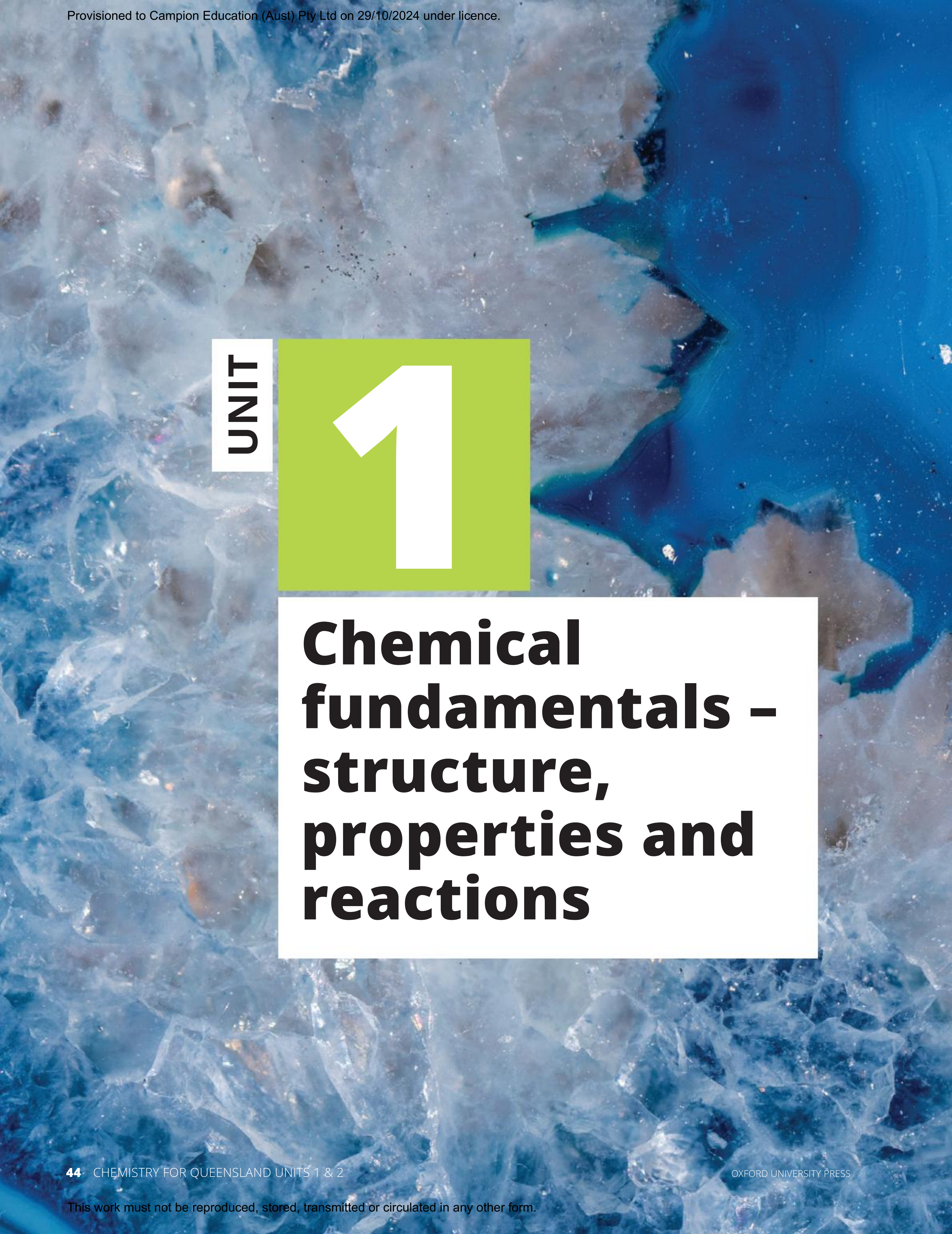
- two of the lightest coins were minted between 1950 and 1966
- two of the heaviest coins were minted between 1970 and 1980.

- **a** **Propose** a possible explanation for why coins appear to be lighter with time. (1 mark)
- **b** The student calculated the average mass of the 12 coins and expressed her answer as: 2.9802 g ± 0.3325 g. **Explain** why this is incorrect, then rewrite her answer to fix the error. (2 marks)

30 **Use** two different online chemistry sources to **investigate** basic information about caffeine (including its melting point, proper chemical name and type of elements it contains). **Consider** which database was easier to use and why. **Discuss** which is more reliable or trustworthy. (2 marks)

Module 1 checklist: Chemistry toolkit

Quizlet: Revise key terms online to test your understanding



UNIT 1

Chemical fundamentals – structure, properties and reactions

Unit 1 overview

All living and non-living objects on Earth are constructed from matter and energy that is neither created nor destroyed. All matter has unique physical and chemical properties. Molecules can break apart and re-form, building whole new structures with new sets of properties.

The ability to model the breaking and re-forming of bonds has allowed chemists to explain how living organisms from the past became combustible fuels. Burning these fuels releases excess energy in an exothermic reaction. This yields new molecules with new properties. During this process, molecules absorb energy from heat and light, and then release the energy in the spectrum of colours used to distinguish and identify unique matter.

Close examination of the properties of matter has revealed a series of patterns, where elements periodically group together in a manner predicted by their subatomic particles. These particles determine how atoms react and bond by sharing or donating electrons to achieve stability with a full outer shell. Measuring such small particles or amounts of chemicals can be difficult, and requires the use of mass, molar mass and moles of a substance.

Unit objectives

1. Describe ideas and findings about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.
2. Apply understanding of the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.
3. Analyse data about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.
4. Interpret evidence about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.
5. 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.
6. Investigate phenomena associated with properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

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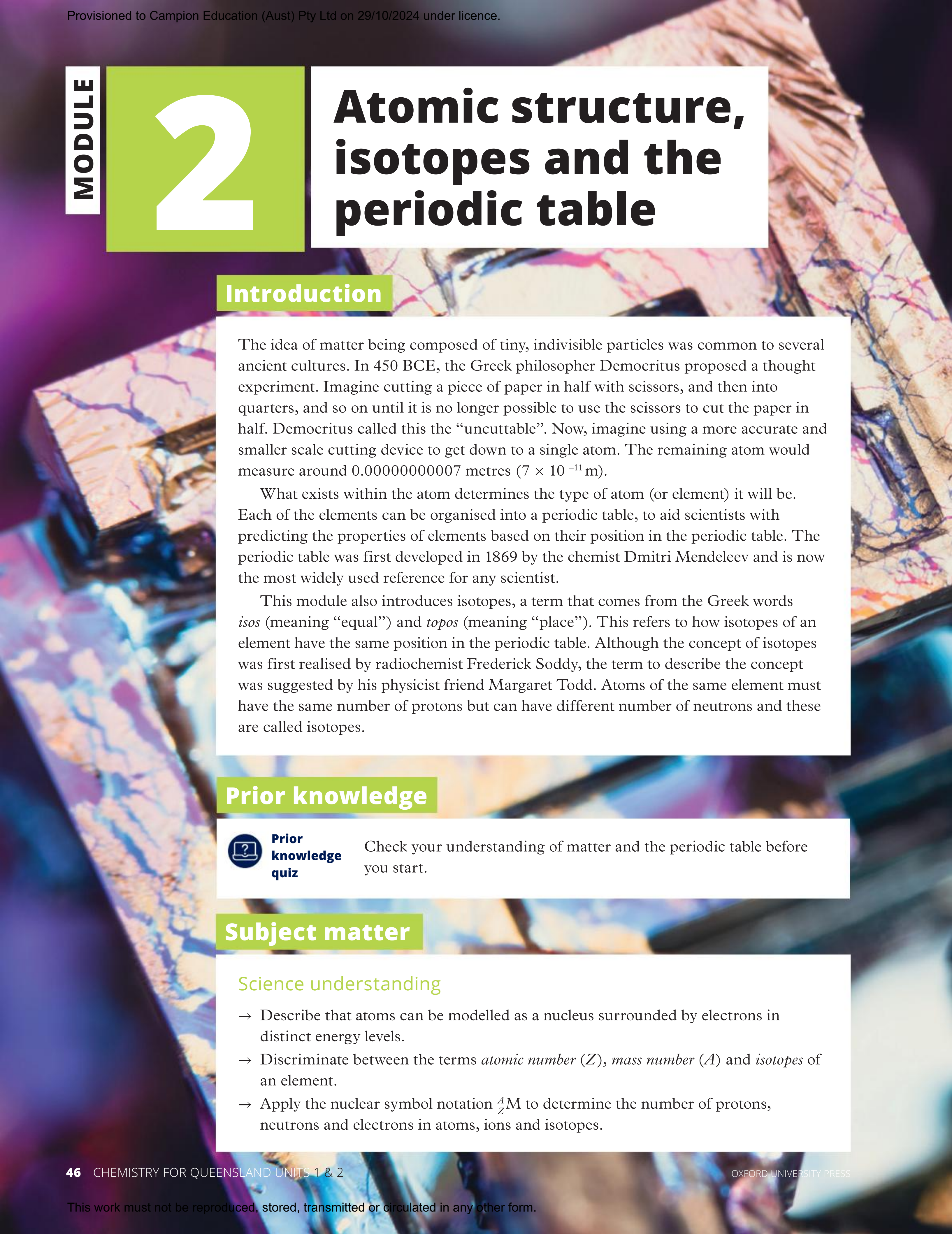
Unit 1 Topics

Topic	Module	
Topic 1 Properties and structure of atoms	**Module 2**	Atomic structure, isotopes and the periodic table
	Module 3	Introduction to bonding
	Module 4	Analytical techniques
Topic 2 Properties and structure of materials	**Module 5**	Compounds and mixtures
	Module 6	Properties and structure of materials
Topic 3 Chemical reactions – reactants, products and energy change	**Module 7**	Mole concept and law of conservation of mass
	Module 8	Chemical reactions



MODULE 2

Atomic structure, isotopes and the periodic table

Introduction

The idea of matter being composed of tiny, indivisible particles was common to several ancient cultures. In 450 BCE, the Greek philosopher Democritus proposed a thought experiment. Imagine cutting a piece of paper in half with scissors, and then into quarters, and so on until it is no longer possible to use the scissors to cut the paper in half. Democritus called this the "uncuttable". Now, imagine using a more accurate and smaller scale cutting device to get down to a single atom. The remaining atom would measure around 0.00000000007 metres (7×10^{-11} m).

What exists within the atom determines the type of atom (or element) it will be. Each of the elements can be organised into a periodic table, to aid scientists with predicting the properties of elements based on their position in the periodic table. The periodic table was first developed in 1869 by the chemist Dmitri Mendeleev and is now the most widely used reference for any scientist.

This module also introduces isotopes, a term that comes from the Greek words *isos* (meaning "equal") and *topos* (meaning "place"). This refers to how isotopes of an element have the same position in the periodic table. Although the concept of isotopes was first realised by radiochemist Frederick Soddy, the term to describe the concept was suggested by his physicist friend Margaret Todd. Atoms of the same element must have the same number of protons but can have different number of neutrons and these are called isotopes.

Prior knowledge

Prior knowledge quiz

Check your understanding of matter and the periodic table before you start.

Subject matter

Science understanding

→ Describe that atoms can be modelled as a nucleus surrounded by electrons in distinct energy levels.

→ Discriminate between the terms *atomic number* (Z), *mass number* (A) and *isotopes* of an element.

→ Apply the nuclear symbol notation $^{A}_{Z}\mathrm{M}$ to determine the number of protons, neutrons and electrons in atoms, ions and isotopes.

- → State the relative energies of the s, p and d orbitals.
- → Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to $Z = 36$.
- → Determine full and condensed electron configurations for atoms and ions up to $Z = 36$, e.g. $1s^22s^22p^63s^23p^5$ and $[Ne]3s^23p^5$.
- → Identify the electron configuration of Cr and Cu as exceptions.
- → Explain how successive ionisation energy data is related to the electron configuration of an atom.
- → Describe that isotopes are atoms of the same element that have different numbers of neutrons.
- → State that isotopes can be represented in the form ^{A}X (IUPAC) or X-A.
- → Identify that isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties.
- → Explain 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 1/12 the mass of an atom of carbon-12.
- → State that elements are represented by symbols.
- → Identify that the structure of the periodic table is based on increasing atomic number.
- → Identify that the periodic table is arranged into four blocks associated with the four sub-levels – s, p, d and f.
- → Describe the relationship between the structure of the periodic table and the electronic configuration of atoms.
- → Explain that elements of the periodic table show trends in chemical and physical properties across periods and down groups as exemplified by groups 1, 2 and 13–18 and period 3.
- → Compare 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).
- → Identify that oxides change from basic through amphoteric to acidic across period 3.
- → Analyse data for atomic radii, valencies, ionic radii, first ionisation energies and electronegativities to determine periodic trends, patterns and relationships.

Science as a human endeavour

- → Appreciate that experiments provided evidence that enabled scientists to develop models of the atom.
- → Consider the role Geiger–Marsden's gold foil experiments and Maria Goeppert Mayer's nuclear shell model played in the development of atomic theory.

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Practicals

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Lesson 2.2 Simulating Geiger–Marsden's gold foil experiment
Lesson 2.7 Investigating periodic table trends using a database

Lesson 2.1
The atomic model

Key ideas

- → An atom can be modelled as a nucleus surrounded by negatively charged electrons in distinct energy levels (and sublevels) held together by electrostatic forces of attraction between the nucleus and electrons.
- → Inside the nucleus are the positively charged protons and neutral neutrons.
- → Nuclear symbol notation ${}^{A}_{Z}M$ summarises the number of subatomic particles in atoms, ions and isotopes.

What's inside an atom?

Learning intentions and success criteria

An atom is an individual unit that makes up all **matter**. It was once thought that the atom was the smallest unit in the universe – the term comes from the Greek word *atomos*, meaning "indivisible". However, it is now known that the atom is made up of even smaller particles.

Scientists have developed a number of models to describe an atom's structure.

The current model of the atom includes a number of components. The **nucleus** is a significant component of the atom. It contains the **protons** and **neutrons**, which are two of the atom's **subatomic particles**. The nucleus contains nearly all of an atom's mass, yet it can be as much as 100,000 times smaller in diameter than the atom itself.

The protons in the nucleus each have a positive charge and mass of 1.67×10^{-24} g. The neutrons do not have a charge and have the same mass as the protons. Because of the positive charge of protons, an atomic nucleus has an overall positive charge.

The third subatomic particle is the **electron**. Electrons have a negative charge and are insignificant in mass compared to the protons and neutrons (9.11×10^{-28} g). The electrons are located outside of the nucleus. They remain close to the nucleus because of the **electrostatic forces** of attraction between the positive charge of the nucleus and the surrounding negatively charged electrons.

matter
physical substances that have mass

nucleus
the dense structure within the atom

proton
a positively charged subatomic particle that exists inside the nucleus

neutron
a neutral subatomic particle that exists inside the nucleus

subatomic particle
a component that makes up, and exists within, the atom

electron
a negatively charged subatomic particle that exists outside of the nucleus

electrostatic force
an attraction between objects of opposing charges

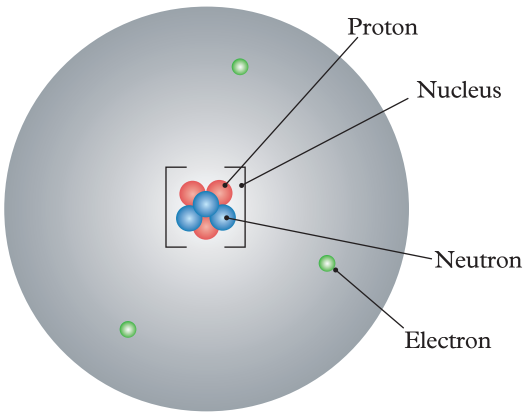


FIGURE 1 A basic model of an atom showing the subatomic particles

Challenge

Magnifying an atom: how tiny is the nucleus?

An atom has a radius of between 10,000 and 100,000 times the radius of the nucleus. **Consider** a scale model of an atom placed on an athletics oval with a radius of 100 m. **Discuss** how large the nucleus would be in this scale model. (2 marks)

How are atoms represented?

The **nuclear symbol** summarises the subatomic particles within an atom (Figure 2).

- *Z* is the atomic number. It is used for identification because it is the number of protons, which does not change for atoms of the same element.
- *A* is the mass number. This represents the total number of subatomic particles in the nucleus (the protons and neutrons). The mass number is sometimes represented as a whole number on the **periodic table**, but it can include decimal points.
- M is the **element** symbol used for each unique element in the periodic table, consisting of one or two letters.

FIGURE 2 The nuclear symbol; for example, $^{4}_{2}He$

Applying the nuclear symbol

An atom, in its original state, has a neutral charge because the positive charge from protons cancels out the negative charge from the electrons. If the protons and electrons are not balanced, then the atom has an ionic charge (either positive or negative). The charged atomic species is called a **monatomic ion**.

Isotopes of an element have different numbers of neutrons but the same number of protons and electrons. The concept of ions and isotopes is covered in more detail in Lesson 2.3.

Table 1 outlines how to determine the number of subatomic particles present. This applies to neutral atoms, ions and isotopes.

TABLE 1 Determining the number of subatomic particles in different species

Subatomic particle	Neutral atom	Ion	Isotope
Proton (p^+)	Use the atomic number from the periodic table. $p^+ = Z$	Use the atomic number from the periodic table. $p^+ = Z$	Use the atomic number from the periodic table. $p^+ = Z$
Neutron (n^0)	Subtract the atomic number from the mass number given in the periodic table. $n^0 = A - Z$	Subtract the atomic number from the mass number given in the periodic table. $n^0 = A - Z$	Subtract the atomic number from the atomic mass of the specific isotope. $n^0 = \text{atomic mass} - Z$
Electron (e^-)	Use the atomic number from the periodic table. $e^- = Z$	Subtract the ionic charge from the atomic number. $e^- = Z - \text{ionic charge}$	Use the atomic number from the periodic table. $e^- = Z$

nuclear symbol
the $^{A}_{Z}$M notation used to represent atoms, where *A* is the mass number, *Z* is the atomic number and M is the element symbol

periodic table
an organised presentation of elements by increasing atomic number

element
a type of atom that is defined by its atomic number

Study tip

The first letter of a chemical symbol is always capitalised. If there is a second letter, it is lower case.

monatomic ion
an electrically charged species (positive or negative) consisting of a single atom

isotope
an atom of an element that contains a specific number of neutrons

Study tip

Not every periodic table is presented in the same manner, but there will be a legend to help interpret the information. In the QCAA Chemistry data and formula book, the legend is in the middle, with hydrogen as an example.

Worked example 2.1A

Determining the number of subatomic particles in neutral atoms and charged ions

Worked example 2.1A: Watch a video that shows how to solve this problem.

Determine:

a the number of neutrons in the neutral atom ${}^{19}_{9}F$ (1 mark)

b the number of electrons in the negative ion ${}^{14}_{7}N^{3-}$. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to look at the nuclear symbol notation for each species and find the value that is being asked for. The questions are worth 1 mark each, so we must analyse the notations presented to us and then complete any calculations.
Step 2: Identify the mass number, atomic number and ionic charge from the nuclear symbols of each atom or ion.	**a** Mass number (A) = 19 Atomic number (Z) = 9 Ionic charge = 0 **b** Mass number (A) = 14 Atomic number (Z) = 7 Ionic charge = −3
Step 3: Calculate the number of neutrons using the formula: $n^0 = A - Z$ or the number of electrons using the formula $e^- = Z -$ ionic charge.	**a** $n^0 = 19 - 9$ $= 10$ neutrons (1 mark) **b** $e^- = 7 - (-3)$ $= 10$ electrons (1 mark)

Your turn

Determine the number of neutrons and electrons in the positive ion ${}^{39}_{19}K^{+}$. (2 marks)

Real-world chemistry

Maria Goeppert Mayer and the nuclear shell model of the atom

Although the modern atomic theory dates back to the early 1800s, our understandings of the atom continue to be refined. In the late 1940s, German-born American physicist Maria Goeppert Mayer and the German physicist J. Hans D. Jensen independently developed the nuclear shell model, which later won them the Nobel Prize (1963).

At the time, much of the existing understanding of atoms focused on the location of electrons within an atom. Little was known about how protons and neutrons were arranged within the atomic nucleus. To address this, Goeppert Mayer and Jensen proposed the nuclear shell model, in which nuclear subatomic particles exist in shells with discrete energy levels.

In their model, neutrons and protons distribute into the shells until they are full. They proposed that when the number of nuclear particles was a "magic number", the nuclei were stable. The magic numbers identified by Goeppert Mayer and Jensen were 2, 8, 20, 28, 50 82 and 126. When the number of neutrons or protons was not a magic number, the nucleus was considered unstable. For example, the nucleus of a calcium atom ($Z = 20$) containing 20 protons and 20 neutrons. Both of these are magic numbers so the calcium nucleus is stable.

Goeppert Mayer's model helps us to understand how the nucleus within an atom behaves and its effect on properties like stability. Her work contributed majorly to our understanding of the model of the atom – but, sadly much of Maria's career went without status, recognition or a paid tenured position; at the time she was considered to be studying chemistry "just for the fun of it".

FIGURE 3 Maria Goeppert Mayer

Apply your understanding

1 **Identify** two examples of atoms that would not be considered stable. (2 marks)

2 The advancement of science often occurs over many, many years and with the combined efforts of many, many scientists. The nuclear shell model is an example of how our understanding of chemistry continues to be refined. **Discuss** the importance of collaboration, with reference to the article. (2 marks)

Check your learning 2.1

Check your learning 2.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the properties of the three subatomic particles in an atom. (3 marks)

Analytical processes

2 **Contrast** the terms "atomic number", "mass number", and "isotopie of an element", with reference to the symbols used to represent them. (3 marks)

3 **Use** the periodic table to **derive** the nuclear symbol for
  - a the titanium isotope with 27 neutrons (1 mark)
  - b the bromine isotope with 44 neutrons. (1 mark)

4 **Using** the nuclear symbol $^{23}_{11}\text{Na}$, **determine** the number of protons, neutrons and electrons present in the neutral atom. (3 marks)

5 Using the nuclear symbol $^{27}_{13}\text{Al}^{3+}$, **determine** the number of protons, neutrons and electrons present in the ion. (3 marks)

6 **Determine** how many electrons are equivalent to the mass of one proton. (1 mark)

Practical

Lesson 2.2

Simulating Geiger–Marsden's gold foil experiment

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Lesson 2.3

Isotopes

Key ideas

→ Isotopes are atoms of the same element that have different numbers of neutrons.

→ Isotopes have similar chemical properties but different physical properties.

→ Relative atomic mass is the ratio of the weighted average of the atomic masses of all the naturally occurring isotopes of an element (taking into account their abundances on Earth) to one-twelfth the atomic mass of the ^{12}C isotope.

Learning intentions and success criteria

What are isotopes?

Isotopes are atoms of an element that have the same number of protons but different numbers of neutrons. As a result, all isotopes of an element have the same atomic number but different mass numbers.

For example, hydrogen has three naturally occurring isotopes: protium, deuterium and tritium. Every hydrogen atom contains one proton, which is why the atomic number is 1. The mass number varies between 1 and 3 because the number of neutrons contained in a hydrogen atom varies between 0 and 2 (Figure 1).

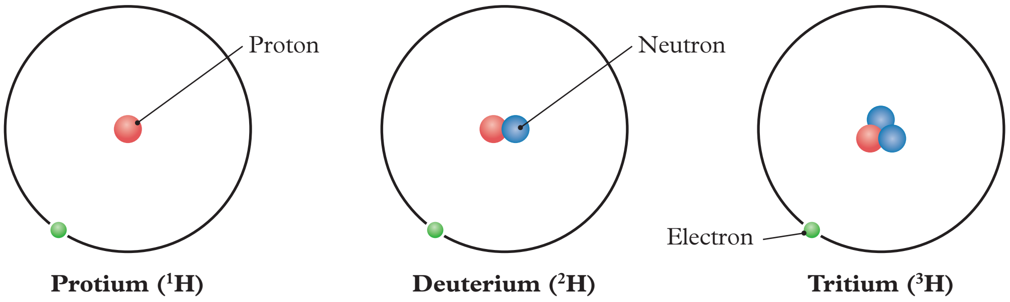


FIGURE 1 The three isotopes of hydrogen

How are isotopes represented?

Study tip

The QCE Chemistry syllabus uses both M and X to represent the elemental symbol.

Isotopes can be identified by using two forms of representation (Table 1): ^{A}X or X-*A*. In both methods:

- *A* represents the mass number
- X represents the elemental symbol (or elemental name) used to identify the number of protons contained in the atom.

TABLE 1 Representations of selected isotopes

^{A}X	X-*A*
^{1}H	H-1 (hydrogen-1)
^{12}C	C-12 (carbon-12)
^{6}Li	Li-6 (lithium-6)

Worked example 2.3A

Representing isotopes of the same element

Worked example 2.3A: Watch a video that shows how to solve this problem.

Determine the nuclear symbols for the two naturally occurring isotopes of boron. The first isotope consists of 5 protons, 5 electrons and 5 neutrons. The second isotope consists of 5 protons, 5 electrons and 6 neutrons. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to look at the subatomic particles for each isotope and represent them using nuclear symbols. The question is worth 2 marks, so we must assess the values provided and fit them to the notation.
Step 2: Identify the mass number (protons + neutrons) and elemental symbol for each isotope.	First isotope: Mass number (A) = 5 + 5 = 10 Elemental symbol (X) = B Second isotope: Mass number (A) = 5 + 6 = 11 Elemental symbol (X) = B
Step 3: Combine A and X using nuclear symbol notation for isotopes. This will be either: ${}^{A}X$ or X-A.	First isotope: ${}^{10}B$ or boron-10 (1 mark) Second isotope: ${}^{11}B$ or boron-11 (1 mark)

Your turn

Determine the nuclear symbol for the four naturally occurring isotopes of iron, which have 28, 30, 31 and 32 neutrons, respectively. Each isotope has 26 protons and 26 electrons. (4 marks)

Do isotopes have different properties?

Isotopes of the same element have the same electron configuration when neutrally charged. For this reason, the chemical properties of isotopes of an element are similar. Remember that the chemical identity of an element depends on its atomic structure.

However, physical properties are determined by the mass number. The varying number of neutrons causes isotopes to have different physical properties. For example, the physical properties of water change depending on which isotope of hydrogen is contained in the water (Table 2). H_2O contains two atoms of ${}^{1}H$, and D_2O contains two atoms of ${}^{2}H$ (where D represents deuterium).

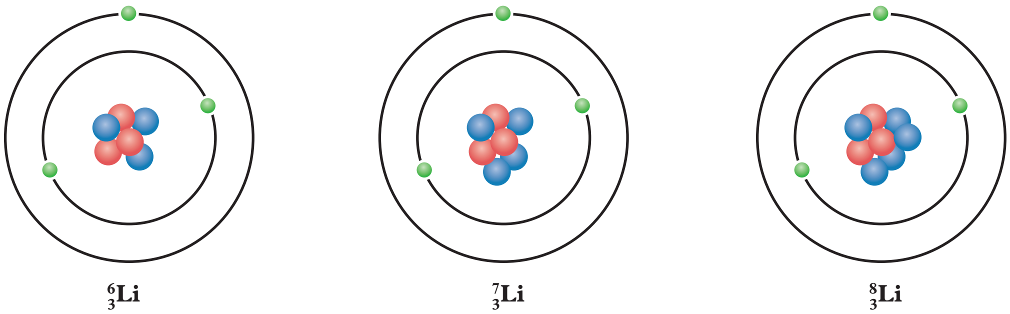


FIGURE 2 The number of neutrons differs between lithium isotopes.

TABLE 2 Physical properties of different types of water

Physical property	D_2O (heavy water)	H_2O (light water)
Melting point (K)	276.97	273.15
Boiling point (K)	374.55	373.15
Density ($g\,mL^{-1}$)	1.1056	0.9982
Surface tension ($N\,m^{-1}$)	0.07187	0.07198
pH	7.44	7.0
Refractive index (at 589 μm)	1.32844	1.33335

Study tip

The unit used for atomic mass is amu, which stands for atomic mass unit and refers to one-twelfth of the atomic mass of the ^{12}C isotope. Relative atomic mass does not have a unit because it is a ratio.

How is relative atomic mass calculated?

Most elements have two or more naturally occurring isotopes, yet only one number is included in the periodic table; this number is based on an average of the atomic masses of all the naturally occurring isotopes for an element. This number is called **relative atomic mass (RAM)** because it has been quantified relative to one-twelfth the atomic mass of the carbon-12 isotope.

Calculating the relative atomic mass of an element uses the atomic masses of all of the naturally occurring isotopes of the element and the natural abundance (usually represented as a **percentage abundance**) of each of the isotopes:

$$\text{RAM} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100} + \frac{(\%_3 \times \text{AM}_3)}{100} \ldots$$

where $\%_1$, $\%_2$ and $\%_3$, are the percentage abundances of isotopes 1, 2 and 3, and AM_1, AM_2 and AM_3, are the atomic masses of isotopes 1, 2 and 3. For elements with more than three isotopes, extra terms are needed. Synthetically prepared isotopes are not used when calculating an element's relative atomic mass.

Boron exists naturally in two forms (or as two different isotopes): boron-10 and boron-11. Boron-11 is the predominant isotope and has a percentage abundance of 80.1%, while boron-10 makes up the remaining 19.9% of boron minerals on Earth. Boron has a RAM of 10.81.

relative atomic mass (RAM)
the mass of an atom once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

percentage abundance
the percentage of each isotope contained within a sample of element; must always add up to 100%

Study tip

The term "mass number" describes the sum of the number of protons and neutrons in an isotope. The term "relative atomic mass" also describes the number of protons and neutrons but as an average number that considers the isotopes (and their natural abundances) of an element, and relative atomic mass appears in the periodic table.

Worked example 2.3B

Calculating RAM using atomic mass and percentage abundance

Worked example 2.3B: Watch a video that shows how to solve this problem.

Calculate the RAM of chlorine using its two naturally occurring isotopes. (2 marks)

TABLE 3 Atomic mass and abundance for isotopes of Cl

Isotope	Atomic mass (amu)	Abundance (%)
^{35}Cl	34.968853	75.77
^{37}Cl	36.965903	24.23

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the RAM of chlorine. The question is worth 1 mark, so we must gather the correct information, correctly apply the formula and complete the calculation.

Think	Do
Step 2: Select the correct formula and gather any data required.	$\text{RAM} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100}$ $\%_1 = 75.77\%$, $\text{AM}_1 = 34.968853$ amu $\%_2 = 24.23\%$, $\text{AM}_2 = 36.965903$ amu
Step 3: Substitute the known values into the formula and find the RAM.	$\text{RAM} = \frac{(34.968853 \times 75.77)}{100} + \frac{(36.965903 \times 24.23)}{100}$ (1 mark) $= 26.4959 + 8.9568$ $= 35.4527$
Step 4: Finalise your answer and make sure you have included the correct units and significant figures or decimal places. RAM does not have units.	Since multiplication and division are used, the answer should be reported to the lowest number of decimal places, i.e. 2. 35.45 (2 d.p.) (1 mark)

Your turn

Calculate the RAM of magnesium using its three naturally occurring isotopes. (2 marks)

TABLE 4 Atomic mass and abundance for isotopes of Mg

Isotope	Atomic mass (amu)	Abundance (%)
^{24}Mg	23.985042	78.99
^{25}Mg	24.985837	10.00
^{26}Mg	25.982593	11.01

Worked example 2.3C

Calculating percentage abundance using RAM

Worked example 2.3C: Watch a video that shows how to solve this problem.

Calculate the percentage abundances of the isotopes of lithium using a RAM of 6.941. ^{6}Li has an atomic mass of 6.015 amu, and ^{7}Li has an atomic mass of 7.016 amu. (3 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the percentage abundance of two lithium isotopes. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required. Since $\%_1$ and $\%_2$ must add up to 100%, we can make one be x and the other $100–x$.	$\text{RAM} = \frac{(\%_1 \times \text{AM}_1)}{100} + \frac{(\%_2 \times \text{AM}_2)}{100}$ $\%_1 = x$, $\text{AM}_1 = 6.015$ amu $\%_2 = 100–x$, $\text{AM}_2 = 7.016$ amu
Step 3: Substitute the known values into the formula and find x.	$\text{RAM} = \frac{x \times 6.015}{100} + \frac{(100 - x)\times 7.016}{100}$ (1 mark) $= 0.06015x + 7.016 - 0.07016x$ $0.01001x = 0.075$ $x = 7.493\%$ $100 - x = 100 - 7.49$ $= 92.51\%$

Think	Do
Step 4: Finalise your answer and make sure you have included the correct units and significant figures or decimal places.	Since multiplication and division are used, the answer is reported to the lowest number of decimal places, i.e. three. % abundance of ^{6}Li (x) is 7.493% (3 d.p.) (1 mark) % abundance of ^{7}Li ($100-x$) is 92.07% (3 d.p.) (1 mark)

Your turn

Calculate the percentage abundances of the isotopes of nitrogen using a RAM of 14.01. ^{14}N has an atomic mass of 14.003074 amu and ^{15}N has an atomic mass of 15.000109 amu. (3 marks)

Check your learning 2.3

Check your learning 2.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the term “isotope” in relation to the number of protons, electrons and neutrons. (1 mark)

2 **Describe** the chemical properties and physical properties of two isotopes of the same element, using an example. (3 marks)

3 **Recall** the two ways in which isotopes are represented. (2 marks)

4 **Explain** what relative atomic mass is. (2 marks)

5 **Calculate** the relative atomic mass of lead from the information of the naturally occurring isotopes of lead listed in Table 5. (2 marks)

TABLE 5 Atomic masses and percentage abundances of the naturally occurring isotopes of lead

Isotopes of lead	Atomic mass (amu)	Percentage abundance (%)
^{204}Pb	203.973037	1.40
^{206}Pb	205.974455	24.1
^{207}Pb	206.975885	22.1
^{208}Pb	207.976641	52.4

6 **Calculate** the percentage abundances of the isotopes of thallium (RAM 204.4). The atomic masses of the isotopes are ^{203}Tl = 202.972 amu and ^{205}Tl = 204.974 amu. (3 marks)

Analytical processes

7 **Deduce** the ^{A}X (elemental symbol) and X-A (element name) forms of two isotopes of copper with atomic masses of 62.929599 amu and 64.927792 amu. (2 marks)

Knowledge utilisation

8 **Infer** the approximate percentage abundances of the following isotopes, by comparing the atomic masses of the isotopes to the relative atomic masses (Table 6). **Justify** your answer for each element. (8 marks)

TABLE 6 Relative atomic masses of selected elements and atomic masses of their respective isotopes

Element	RAM	Atomic mass of isotope 1	Atomic mass of isotope 2
Helium	4.00	3.016	4.002
Carbon	12.01	12.000	13.003
Nitrogen	14.01	14.003	15.000
Bromine	79.90	78.918	80.916

Lesson 2.4

The periodic table

Key ideas

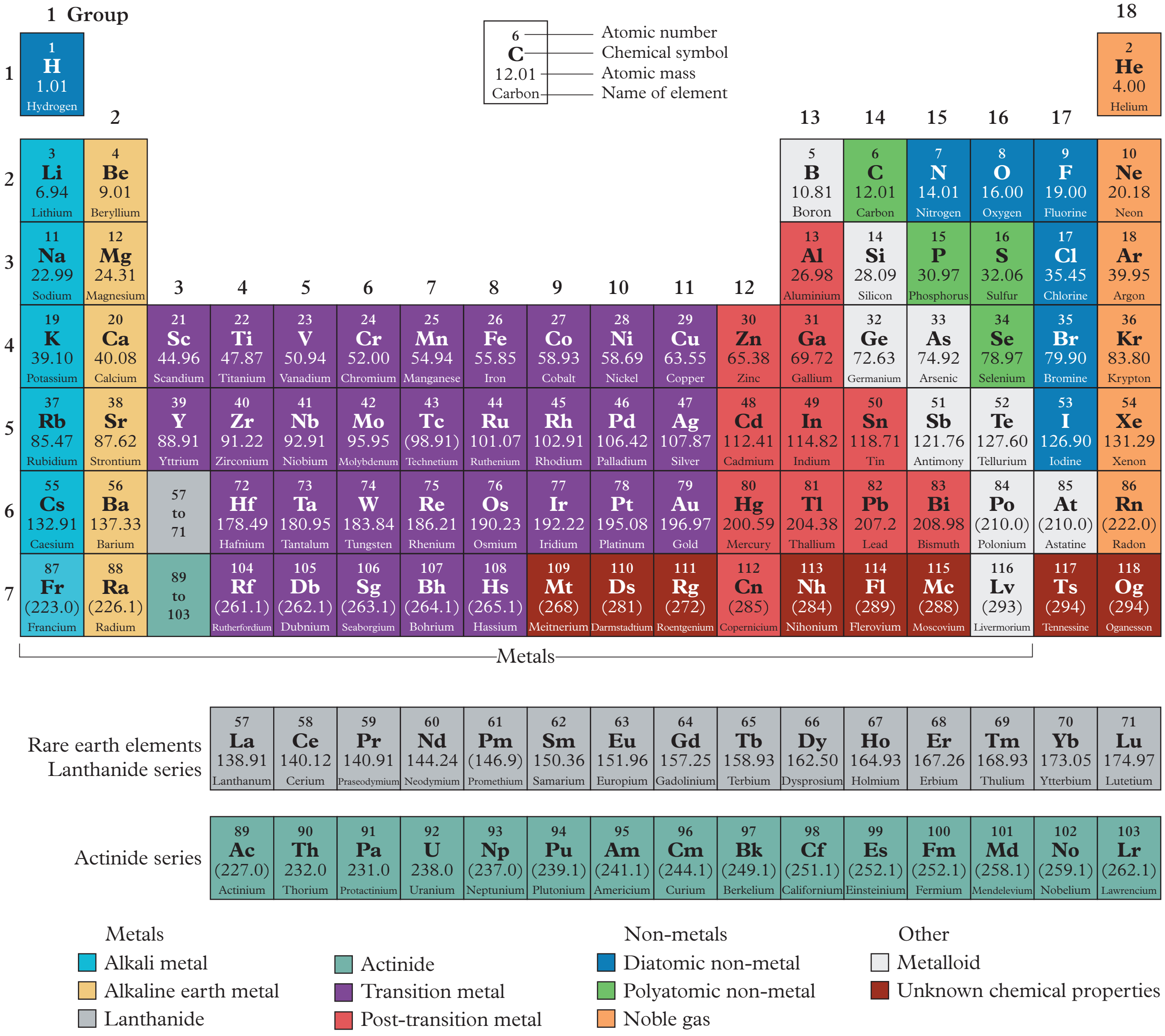
→ Elements on the periodic table are organised into vertical groups and horizontal periods, and into other sections such as metals and non-metals.

→ The location of an element is linked to its properties.

How is the periodic table organised?

Learning intentions and success criteria

Elements are defined by their atomic number; for example, carbon has an atomic number of 6, hydrogen has an atomic number of 1 and iron has an atomic number of 26. The modern periodic table organises elements by increasing order of atomic number (Figure 1).



	1 Group	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01 Hydrogen																	2 He 4.00 Helium
2	3 Li 6.94 Lithium	4 Be 9.01 Beryllium											5 B 10.81 Boron	6 C 12.01 Carbon	7 N 14.01 Nitrogen	8 O 16.00 Oxygen	9 F 19.00 Fluorine	10 Ne 20.18 Neon
3	11 Na 22.99 Sodium	12 Mg 24.31 Magnesium											13 Al 26.98 Aluminium	14 Si 28.09 Silicon	15 P 30.97 Phosphorus	16 S 32.06 Sulfur	17 Cl 35.45 Chlorine	18 Ar 39.95 Argon
4	19 K 39.10 Potassium	20 Ca 40.08 Calcium	21 Sc 44.96 Scandium	22 Ti 47.87 Titanium	23 V 50.94 Vanadium	24 Cr 52.00 Chromium	25 Mn 54.94 Manganese	26 Fe 55.85 Iron	27 Co 58.93 Cobalt	28 Ni 58.69 Nickel	29 Cu 63.55 Copper	30 Zn 65.38 Zinc	31 Ga 69.72 Gallium	32 Ge 72.63 Germanium	33 As 74.92 Arsenic	34 Se 78.97 Selenium	35 Br 79.90 Bromine	36 Kr 83.80 Krypton
5	37 Rb 85.47 Rubidium	38 Sr 87.62 Strontium	39 Y 88.91 Yttrium	40 Zr 91.22 Zirconium	41 Nb 92.91 Niobium	42 Mo 95.95 Molybdenum	43 Tc (98.91) Technetium	44 Ru 101.07 Ruthenium	45 Rh 102.91 Rhodium	46 Pd 106.42 Palladium	47 Ag 107.87 Silver	48 Cd 112.41 Cadmium	49 In 114.82 Indium	50 Sn 118.71 Tin	51 Sb 121.76 Antimony	52 Te 127.60 Tellurium	53 I 126.90 Iodine	54 Xe 131.29 Xenon
6	55 Cs 132.91 Caesium	56 Ba 137.33 Barium	57 to 71	72 Hf 178.49 Hafnium	73 Ta 180.95 Tantalum	74 W 183.84 Tungsten	75 Re 186.21 Rhenium	76 Os 190.23 Osmium	77 Ir 192.22 Iridium	78 Pt 195.08 Platinum	79 Au 196.97 Gold	80 Hg 200.59 Mercury	81 Tl 204.38 Thallium	82 Pb 207.2 Lead	83 Bi 208.98 Bismuth	84 Po (210.0) Polonium	85 At (210.0) Astatine	86 Rn (222.0) Radon
7	87 Fr (223.0) Francium	88 Ra (226.1) Radium	89 to 103	104 Rf (261.1) Rutherfordium	105 Db (262.1) Dubnium	106 Sg (263.1) Seaborgium	107 Bh (264.1) Bohrium	108 Hs (265.1) Hassium	109 Mt (268) Meitnerium	110 Ds (281) Darmstadtium	111 Rg (272) Roentgenium	112 Cn (285) Copernicium	113 Nh (284) Nihonium	114 Fl (289) Flerovium	115 Mc (288) Moscovium	116 Lv (293) Livermorium	117 Ts (294) Tennessine	118 Og (294) Oganesson

Rare earth elements Lanthanide series	57 La 138.91 Lanthanum	58 Ce 140.12 Cerium	59 Pr 140.91 Praseodymium	60 Nd 144.24 Neodymium	61 Pm (146.9) Promethium	62 Sm 150.36 Samarium	63 Eu 151.96 Europium	64 Gd 157.25 Gadolinium	65 Tb 158.93 Terbium	66 Dy 162.50 Dysprosium	67 Ho 164.93 Holmium	68 Er 167.26 Erbium	69 Tm 168.93 Thulium	70 Yb 173.05 Ytterbium	71 Lu 174.97 Lutetium
Actinide series	89 Ac (227.0) Actinium	90 Th 232.0 Thorium	91 Pa 231.0 Protactinium	92 U 238.0 Uranium	93 Np (237.0) Neptunium	94 Pu (239.1) Plutonium	95 Am (241.1) Americium	96 Cm (244.1) Curium	97 Bk (249.1) Berkelium	98 Cf (251.1) Californium	99 Es (252.1) Einsteinium	100 Fm (252.1) Fermium	101 Md (258.1) Mendelevium	102 No (259.1) Nobelium	103 Lr (262.1) Lawrencium

Metals		Non-metals	Other
🟦 Alkali metal	🟩 Actinide	🟦 Diatomic non-metal	⬜ Metalloid
🟨 Alkaline earth metal	🟪 Transition metal	🟩 Polyatomic non-metal	🟥 Unknown chemical properties
⬜ Lanthanide	🟥 Post-transition metal	🟧 Noble gas	

FIGURE 1 The periodic table organises elements by atomic number.

group
a vertical column of the periodic table; reflective of the number of outermost (valence) electrons

Elements with similar chemical and physical properties are located underneath each other in vertical columns called **groups**. The groups are numbered from 1 to 18 with some groups identified by a common name to emphasise their similar chemical and physical properties (Table 1).

TABLE 1 Main groups of the periodic table

Group	Given name	Common properties
1	Alkali metals	Silver coloured, soft metals with a low density. They easily react with water and the halogens. Their reactivity increases down the group (Li–Cs), with caesium being the most reactive.
2	Alkaline earth metals	Silver-coloured, hard metals with a slightly higher density than alkali metals. They react slowly with water, sometimes requiring hot water or even steam.
17	Halogens	Non-metallic. Exist as **diatomic molecules** in their natural state. They are highly reactive with alkali metals and can be harmful to the environment. The halogens decrease in reactivity down the group (F–At) with fluorine being the most reactive.
18	Noble gases	Mostly unreactive.

diatomic molecule
a molecule containing two atoms of the same element or of different elements bonded together

period
a horizontal row of the periodic table; reflective of the number of energy levels (shells) within an atom

The horizontal rows of a periodic table are termed **periods** and are numbered from 1 to 7. The period number corresponds to the highest energy level occupied by electrons of that element. For example, magnesium (Mg) is located in period 3 so the electrons in magnesium are in the first three energy levels. The structure of the periodic table therefore provides information about an element simply by its location. You will learn more about trends in the periodic table in Lesson 2.6.

What categories are elements organised into?

The periodic table consists mainly of main group elements which are further broken down into metals and non-metals. There are also the metalloids, transition metals, basic metals and rare earth elements.

Metals and non-metals

melting point
the temperature at which a solid becomes a liquid

density
how closely packed atoms are arranged; the mass of a substance per unit of volume

malleable
can be hammered into other shapes

ductile
can be drawn out into wires

boiling point
the temperature at which a liquid becomes a gas

Metals are usually shiny solids with high **melting points**, heat and electrical conductivities and **densities**. Metals are also very **malleable** and **ductile**. The exception is mercury, which is a liquid at room temperature. Typically, metals lose electrons in chemical reactions, while non-metals gain electrons. Metals are elements found on the left side and in the middle section of the periodic table. The elements in group 1 excluding hydrogen are called alkali metals and the metals in group 2 are called the alkaline earth metals.

Non-metals commonly have low melting and **boiling points**, do not conduct heat and electricity, and do not have a shiny surface. Examples of non-metals are oxygen, nitrogen, carbon, phosphorus, sulfur, chlorine, neon, helium and hydrogen (even though it is not usually located with the others). Non-metals are elements found on the right side of the periodic table, separated from the metals by a diagonal group of metalloids, also called semi-metals.

FIGURE 2 Neon is a noble gas used in LED lights.

Metalloids

Where the metals and non-metals overlap are the **metalloids**, which display both metallic and non-metallic properties. For example, silicon is a very poor conductor of electricity at room temperature, but a better electrical conductor when heated. Other metalloids are boron, germanium, arsenic, antimony, tellurium, polonium and astatine. Previously, metalloids were known as semi-metals.

metalloid
an element that displays both metallic and non-metallic properties

Study tip

Not every periodic table indicates the metalloids. To help remember them, start at silicon and move down one, then one to the right, until reaching astatine. Don't forget boron up the top.

Transition metals

Transition metals are elements in groups 3–12 of the periodic table which exhibit the same characteristics as the main group metals, but are not as reactive. They often form colourful compounds; for example, vanadium compounds can be yellow, blue, green and purple.

FIGURE 3 Vanadium is a transition metal that forms compounds of different coloured solutions.

Study tip

There is some confusion about the inclusion of group 12 into the transitional metal category. The International Union of Pure and Applied Chemistry (IUPAC) defines them as not being transition metals, despite them sharing many properties with the neighbouring transition metals in group 11. For simplicity, we will consider group 12 as transition metals in QCE Chemistry.

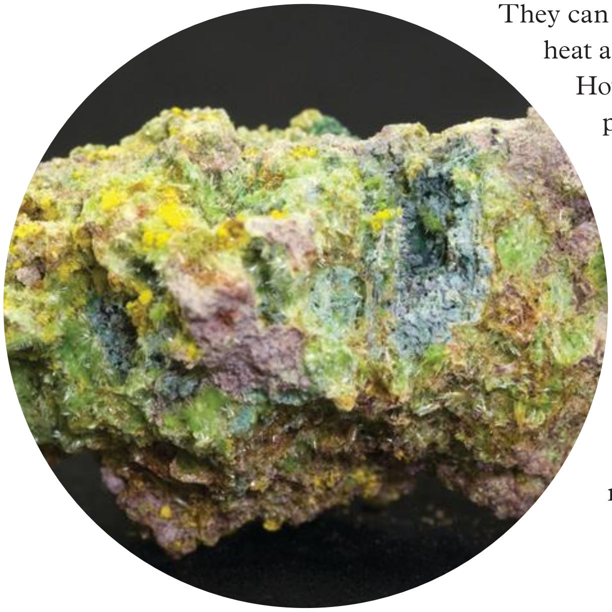
Challenge

The authority for naming elements

The International Union of Pure and Applied Chemistry (IUPAC) is an organisation that has represented chemists in individual countries since 1919. It is the recognised world authority for the naming of elements and compounds. **Discuss** the significance of an organisation such as IUPAC, and why it is necessary for the advancement of science. (3 marks)

Basic metals

The basic metals are located underneath the non-metals and metalloids but to the right of the transition metals. Examples of basic metals are aluminium, gallium, indium, tin, thallium, lead, and bismuth.



They can be recognised as metals because they are good conductors of heat and electricity, and are lustrous, dense, malleable and ductile. However, the basic metals also have some similar non-metallic properties as the elements above them. Lead and gallium are soft, and some of the basic metals have lower melting points than is expected of a metal.

Rare earth elements

Known as the rare earth elements or inner transition metals, the two rows displayed underneath the periodic table consist of the lanthanoids (or lanthanides) (first row) and the actinoids (or actinides) (second row). They are considered special transition metals that are very dense and can be radioactive. Some are found naturally, and others are synthetic.

FIGURE 4 Uranium is a rare earth element.

Real-world chemistry

Are our elements becoming extinct?

Geoscience Australia has identified 50 naturally sourced minerals as "critical minerals". More than 50% of these are elements. They are considered "critical" because their limited supply can significantly affect large industries and economies. Critical elements and minerals are at a high risk of depletion if their recovery or extraction is not appropriately managed.

Helium

Despite being the second most abundant element in the universe, helium is one of the most critically endangered elements on the critical element list. Supplies of helium are finite, meaning that it cannot be artificially reproduced or extracted from the atmosphere safely. Additionally, replenishment of helium supplies will take millions of years. These factors combined with the high use of helium have placed the element in a critically endangered position.

The use of helium goes well beyond filling balloons and aircrafts. Helium's primary use is to cool substances to extremely low temperatures and create controlled environments. This is needed to run superconductors, MRI (magnetic resonance imaging) scanners in hospitals, NMR (nuclear magnetic resonance) spectrometers in chemistry laboratories and several other machines. Helium is also used to detect pipeline leaks because of its small size.

Recovering and recycling helium

Helium used in laboratories and hospitals can be recovered by a re-liquefier that captures and condenses 95% of the helium gas into usable helium. Conserving helium this way, as well as continually finding replacements for helium, can ensure we use this non-renewable resource more sustainably to protect its availability into the long-term future.

Apply your understanding

1 **Conduct** further research to **identify** two critically endangered elements at risk of depletion and **describe** their uses in industry. (2 marks)

2 Recycling materials is a way to manage the use of critical elements and ensure their longevity. **Describe** two factors that could limit the recycling of helium. (2 marks)

3 Hydrogen was once considered as a plausible replacement for helium; however, it was deemed unsuitable due to its flammability. Locate hydrogen and helium on the periodic table and **suggest** one alternative. **Justify** your choice. (2 marks)

Check your learning 2.4

Check your learning 2.4: Complete these questions online or in your workbook.

Note: Use the periodic table in the Appendix to help answer these questions.

Retrieval and comprehension

1 **Describe** how the periodic table is organised. (3 marks)

2 **Identify** the properties of mercury that make it unique compared to most other metals. (2 marks)

Analytical processes

3 **Compare** the terms "atom" and "element". (2 marks)

4 **Categorise** the element in period 4 and group 17 as a metal, non-metal or metalloid. (1 mark)

Lesson 2.5

Electron configurations

Key ideas

→ Electrons are located within energy levels, which are organised further into sublevels and orbitals (s, p, d and f).

→ Electron configurations can be represented in full or condensed formats. They are used to represent the location of the electrons in an atom or ion.

→ The Aufbau principle, Pauli Exclusion principle and Hund's rule are used to write full and condensed electron configurations.

How are electrons organised into energy levels?

Learning intentions and success criteria

Outside of the nucleus, the electrons are located within distinct energy levels at different distances from the nucleus. The electrons closest to the nucleus have the largest electrostatic attraction to the nucleus and the least amount of energy. The energy (n) levels are numbered with integers started from 1 (closest to the nucleus) and are further divided into sublevels.

As the energy levels increase, there is a tendency for electrons in the outer energy levels (higher n values) to spend more time on average further from the nucleus. This means that electrons in higher energy levels experience less attraction to the positively charged nucleus. Consequently, removal of successive electrons causes the charge of a monatomic atom to increase. This means that it becomes harder to remove an electron, i.e. it requires more energy. The energy required to remove the next electron is termed successive ionisation energy. You will learn more about this in Lesson 2.6.

Electrons within the same energy level are organised into four sublevels, in order of increasing energy: s, p, d and f, as shown in Table 1. Not every energy level contains every sublevel. For $n = 1$, only an s sublevel is possible. For $n = 2$, both s and p sublevels can exist. The s, p and d sublevels are present at $n = 3$, and all four are present at $n = 4$.

TABLE 1 Breakdown of sublevels present in each energy level

Energy level	s	p	d	f
1	✓			
2	✓	✓		
3	✓	✓	✓	
4	✓	✓	✓	✓

Each sublevel has its own number of orbitals and distinct orbital shapes that denote the probably locations of its electrons. Each orbital contains a maximum of two electrons.

Where are the s, p, d and f blocks on the periodic table?

The periodic table can be split into four blocks based on the s, p, d and f energy sublevels. Each block represents the subshell where valence electrons or electrons in the outermost shell are likely to be found in an atom. For example, magnesium is located in the s block as its outermost or valence electrons occupy an s subshell.

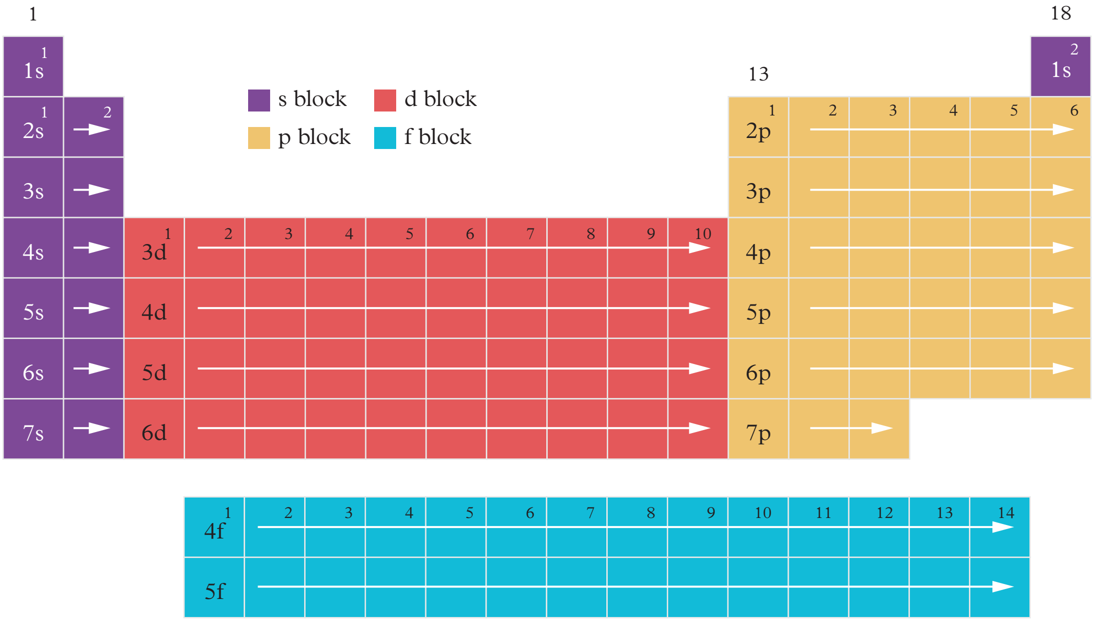


FIGURE 1 Blocks s, p, d and f on the periodic table

What is an electron configuration?

electron configuration
a representation of the electrons in an atom

An **electron configuration** shows where the electrons are located within an atom. Each energy level surrounding an atom is given a letter (s, p, d and f). A breakdown of each energy level capacity is shown in Table 2.

TABLE 2 Breakdown of sublevel capacity

Sublevel	Number of orbitals	Number of electrons
s	1	2
p	3	6
d	5	10
f	7	14

In an electron configuration, the energy level is the number, the sublevel is the letter, and the number of electrons occupying the **orbitals** is the superscript. The order of energy levels and sublevels when creating an electron configuration is as follows:

$$1s^2 \rightarrow 2s^2 \rightarrow 2p^6 \rightarrow 3s^2 \rightarrow 3p^6 \rightarrow 4s^2 \rightarrow 3d^{10} \rightarrow 4p^6$$

Challenge

A pattern for electron configurations

There are many ways to memorise a sequenced list of information. **Analyse** the diagram in Figure 2 to **propose** what mathematical pattern is present, starting at 2p. (1 mark)

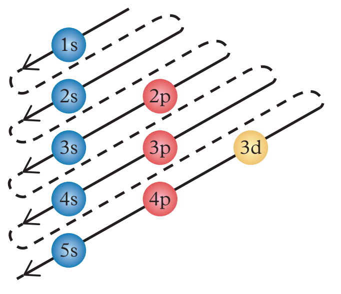


FIGURE 2 The order in which sublevels should be filled for electron configurations

A **full electron configuration** states the location of every electron in the atom, beginning at 1s and continuing until there are no more electrons. For example, nickel has 28 electrons. Its full electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$.

A **condensed electron configuration** is a way of abbreviating the configuration by stating the previous noble gas in square brackets and then stating the location of the remaining outer electrons in the atom. Noble gases are located in the far-right column of the periodic table: helium, neon, argon, krypton, xenon and radon. Using the example of nickel again, the closest noble gas with a smaller atomic number is argon (18 electrons). Beginning at argon, the condensed electron configuration is: $[Ar]4s^2 3d^8$. So, [Ar] + 2 + 8 = 2 electrons or [18] + 10 = 28 electrons. This is shown in Figure 3.

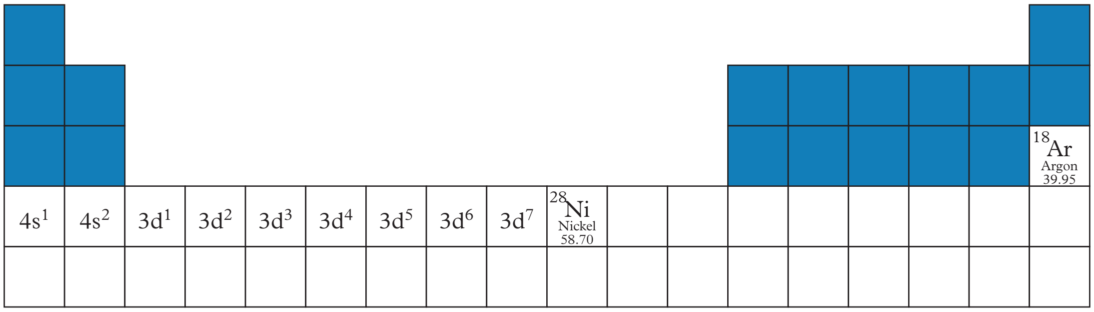


FIGURE 3 Representation of nickel's condensed electron configuration

How are electron sublevels and orbitals filled?

Starting at the top-left corner of the periodic table (hydrogen), electron sublevels are filled from left to right across the periodic table as the atomic number increases by one each time. The energy sublevel filled (or partially filled) last will align with the location of that element within the periodic table.

orbital
the three-dimensional region of space around an atomic nucleus where an electron is most likely to be found

Study tip

You do not need to know the shapes of the orbitals, just their relative energies: the s orbital contains the lowest energy electrons, followed by the p and d orbitals. Electrons in the f orbital have the highest energy.

Study tip

You need to know the order for filling electron configurations up to element 36. It is not chronological, but there is a visual pattern shown in the diagonal diagram (Figure 2), to help remember the order. Understanding the relationship of blocks in the periodic table to orbitals is also useful.

full electron configuration
an electron configuration that states the location of every electron in the atom

condensed electron configuration
an electron configuration that only states electrons in the highest energy levels, based on the nearest noble gas

Study tip

You are required to construct electron configurations for atoms and ions up to *Z* = 36.

For example, phosphorus is located in the p block of the periodic table, specifically within the second row of the p block (which refers to 3p). Therefore, the electron configuration for phosphorus should end at 3p. Phosphorus has 15 electrons. Filling the sublevels in order, the full electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Though often not explicitly examinable, orbital diagrams are handy to understand how electron sublevels and orbitals fill (Figure 4). An orbital diagram helps with visualising the electron configuration. Each square represents an orbital and orbitals in the same sublevel are shown in a line. The filling of electron subshells follows the Aufbau principle, the Pauli exclusion principle and Hund's rule.

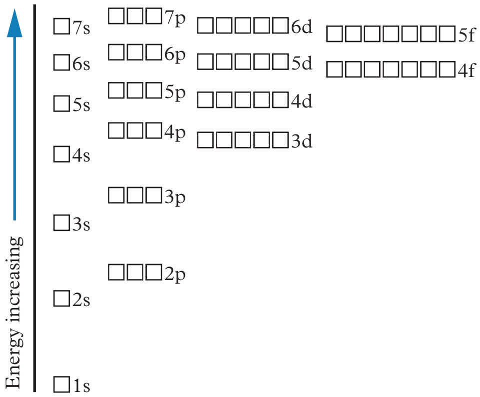


FIGURE 4 The energy levels of atomic orbitals. Each square represents an orbital.

Study tip

Not every periodic table has the same layout, especially for the f block. It is important to note that 4f begins with La and finishes with Yb (14 elements in total). 5d begins with Lu. 5f begins with Ac and then 6d begins with Lr.

Aufbau principle

The Aufbau principle states that electrons fill the orbital of the lowest energy first. Beginning at 1s, the electrons cannot be placed in any orbital outside of this order. However, this strict order starts to break down further down the periodic table as orbitals become closer in energy.

Pauli exclusion principle

The Pauli exclusion principle states that an atomic orbital can contain at most two electrons at any time, but they must have the opposite "spin" represented by the direction of an arrow. This means that an orbital can contain 0, 1 or 2 electrons. The first electron to enter an orbital is drawn as a half-arrow pointing up or down. A second electron must have an opposing direction to the first electron (Figure 5).

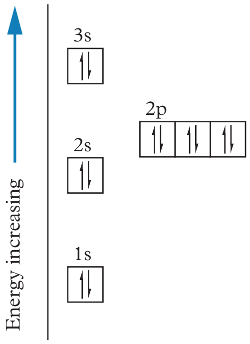


FIGURE 5 An orbital diagram for magnesium, which has 12 electrons. The first electron in each orbital is drawn as a half-arrow pointing up or down, and the second is a half-arrow pointing in the opposite direction.

Hund's rule

Hund's rule states that electrons occupy orbitals of equal energy (p has three orbitals, d has five orbitals and f has seven orbitals). Then, one at a time, electrons go into single orbitals of the same energy, all with the same "spin" (i.e. electron half-arrows pointing the same way) before a second electron with an opposite "spin" (opposite direction of the half-arrow) can enter any of the orbitals (Figure 6).

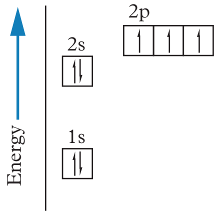


FIGURE 6 An orbital diagram for nitrogen, which has 7 electrons.

Study tip

Orbital diagrams are shown here to help you understand the Aufbau principle, the Pauli exclusion principle and Hund's rule. You do not need to know how to draw them.

Worked example 2.5A

Applying knowledge to write electron configurations

Worked example 2.5A: Watch a video that shows how to solve this problem.

Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write the full and condensed electron configurations for silicon. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Apply" means to use knowledge and understanding in response to a given situation or circumstance; carry out or use a procedure in a given or particular situation. The knowledge and understanding to be applied are the Aufbau principle, Hund's rule and the Pauli exclusion principle. We need to write electron configurations. The question is worth 2 marks, so we must gather the correct information and express our answer in the correct ways.
Step 2: Locate the element on the periodic table and identify the number of electrons.	Silicon (Z = 14) has 14 electrons.
Step 3: Recall the Aufbau order of electron shell and subshells.	1s, 2s, 2p, 3s, 3p, 4s, 4p, 3d, 5s, 4d, 5p
Step 4: Distribute the electrons into their subshells. Remember that s orbitals hold a maximum of two electrons, the p orbitals in each level can hold six and the d orbitals in each level can hold a maximum of 10.	Full electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^2$ (1 mark)
Step 5: Find the closest noble gas with a smaller atomic number.	Neon is the closest (Z = 10), with 10 electrons.
Step 6: Subtract the number of electrons in the closest noble gas from the number of electrons in the element of interest.	14 − 10 = 4 electrons
Step 7: Write the elemental symbol for the closest noble gas using square brackets, followed by the location of the remaining outer electrons in the atom.	Remaining outer electrons: $3s^2 2p^2$ Condensed electron configuration: $[Ne]3s^2 2p^2$ (1 mark)

Your turn

Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write the full and condensed electron configurations for cobalt. (2 marks)

Are there exceptions to the Aufbau principle?

As with many other scientific relationships observed in chemistry, there are exceptions to the Aufbau principle. This is because occupying specific orbitals allows the atom to be more stable overall.

Chromium has 24 electrons and is located in the d block (more specifically, it is the fourth element in the third row of the periodic table). The expected electron configuration of chromium is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$. The actual electron configuration of chromium is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$. The expected electron configuration has 4s full and 3d partially full, whereas the actual electron configuration has both 4s and 3d half-full and is therefore more stable overall. This is because some energy is required to pair up two electrons in the same orbital and the 4s and 3d orbital are very close in energy.

The same situation applies to copper: 4s becomes half-full so that 3d can become completely full. Thus, the actual electron configuration of copper is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$.

TABLE 3 Effect of electron occupancy on stability

Level of stability	Degree sublevel is filled with electron	Example
Large	Completely full	$3d^{10}$
Medium	Half-full	$3d^5$
Small	Partially full	$3d^2$ or $3d^8$

Check your learning 2.5

Check your learning 2.5: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Identify** how many orbitals are present in each of the electron sublevels. (1 mark)

2 **Identify** the maximum number of electrons that can occupy each of the electron sublevels. (1 mark)

3 **Explain** why electrons in the 7s sublevel would have less electrostatic attraction to the nucleus than electrons in the 2s sublevel. (2 marks)

4 **Recall** the relative energies of the s, p and d orbitals. (1 mark)

Analytical processes

5 **Deduce** the condensed electron configuration for chromium. (1 mark)

6 **Determine** the identity of the element from its full electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$. (1 mark)

7 **Determine** the identity of the element from its condensed electron configuration: $Ar4s^1$. (1 mark)

8 **Determine** the maximum number of electrons that can be present in the
   - **a** 4p orbital (1 mark)
   - **b** 5d orbital (1 mark)
   - **c** the third energy level. (1 mark)

9 **Determine** the maximum number of electrons that can be present in the highest occupied sublevel for oxygen. (1 mark)

10 **Apply** the Aufbau principle, Hund's rule and the Pauli exclusion principle to write the full electron configurations of boron, chlorine and lithium. (3 marks)

Knowledge utilisation

11 The complete electron configuration for bromine is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$. The accepted convention for writing electron configurations rearranges this to $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$ to give meaning to the readers. **Evaluate** the significance of this minor rearrangement. (2 marks)

12 **Investigate** the significance of using noble gases as reference points in condensed electron configurations. (1 mark)

Lesson 2.6

Trends in the periodic table

Key ideas

→ Atomic radii and metallic character decrease across a period and increase down a group.

→ Valency increases across a period and stays consistent down a group.

→ Ionic radii are smaller than atomic radii when the ion is positive, and larger than atomic radii when the ion is negative.

→ Ionisation energy and electronegativity generally increase across a period and decrease down a group.

What trends can be observed on the periodic table?

Learning intentions and success criteria

The periodic table is cleverly organised such that the position of an element allows us to predict its physical and chemical properties. In this lesson, we will explore the relationship between where elements are located and properties such as:

- atomic radius
- valency
- ionic radius
- first ionisation energy
- electronegativity
- metallic and non-metallic character
- acidity and basicity of oxides.

What is atomic radius?

The **atomic radius** refers to the size of an atom. The outer edge of an atom is too difficult to measure because of the difficulty of locating the electrons precisely, so the distance between adjacent nuclei is measured (in **picometres**) and that quantity is halved.

atomic radius
used to describe the size of an atom, which is measured by halving the distance between the nuclei of adjacent atoms

picometre
10^{-12} or 0.000000000001 metres

electron shielding effect
the shielding of outer electrons by inner electrons, causing a smaller force of attraction to the nucleus to be felt by the outer electrons

Atomic radii and the periodic table

Atomic radii increase down a group because of the increase in the total number of electrons, combined with the **electron shielding effect**. With each energy level that is added to an atom, there are many more electrons that need to occupy space, so the atom increases in size. For example, within the alkali metal group, lithium is smaller than sodium, which is smaller than potassium and then rubidium (Table 1).

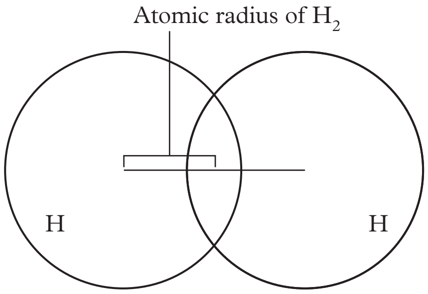


FIGURE 1 A representation of how the atomic radius is determined

TABLE 1 The relationship between atomic number and radius of some alkali metals

Alkali metal	Atomic number	Atomic radius (pm)
Lithium	3	145
Sodium	11	180
Potassium	19	243
Rubidium	37	235
Caesium	55	260

Study tip

Think of adding another energy level as a person putting on another jumper or sweater. The extra clothing is going to take up some room even if it is thin. The more layers of clothing added, the bigger a person looks.

Atomic radii decrease from left to right across a period because of electrostatic attractions. With each increase in atomic number, there is an increase in the number of positively charged protons (in the nucleus) and the attractive force that the negatively charged electrons experience. The electrostatic attraction between the additional subatomic particles pulls the electrons closer to the nucleus, and the atomic radius gets slightly smaller. The smallest atomic radius in any period is that of the noble gas, and the largest is that of the alkali metal.

valence electron
an electron in the outermost energy level of an atom

valency
the combining power of an element as determined by the number of electrons in its outer energy level (valence electrons)

What is valency?

Valence electrons are the electrons occupying the outer most energy level of an atom. For example, silicon has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^2$. The third energy level contains silicon's four valence electrons – two in the 3s orbital and two in the 3p orbital. The term **valency** therefore describes the number of valence electrons that an atom has. Valency determines many of the elements' properties, including their reactivity.

Skill drill

Presenting atomic radii data

Science inquiry skill(s): Processing and analysing data (Lesson 1.7) and Communicating scientifically (Lesson 1.9)

The atomic radii and numbers of the first seven elements are given in Table 2, along with values for Na, P and K.

TABLE 2 Atomic numbers and radii for the selected elements

Element	Atomic number	Atomic radius (pm)
Hydrogen	1	32
Helium	2	37
Lithium	3	130
Beryllium	4	99
Boron	5	84
Carbon	6	75
Nitrogen	7	71
Sodium	11	160
Phosphorus	15	109
Potassium	19	200

Practise your skills

1 **Construct** a graph for the data presented for the first seven elements in Table 2. (2 marks)

2 **Describe** the relationship between atomic number and atomic radius. (2 marks)

3 **Deduce** one explanation as to why the atomic radius of sodium is lower than that of potassium. (1 mark)

4 **Comment** on why the atomic radius of lithium is larger than the atomic radius of phosphorus despite having a smaller number of electrons. (2 marks)

Valency and the periodic table

Valency increases across a period, with the noble gas in any period having complete sublevels. This means noble gases are more stable than the other elements in the period, and largely unreactive. With an increase of one in the atomic number, there is an increase of one electron, which begins to fill the s sublevel in the next energy level.

Valency remains constant down a group. Every alkali metal has one valence electron in its s sublevel, and every halogen has seven valence electrons (two in its s sublevel and five in its p sublevel).

What is ionic radius?

Atoms tend to lose or gain electrons to achieve complete s and p sublevels, particularly the first row of elements. When this happens, atoms gain an **ionic charge**. When an atom has lost one or more electrons, it forms a positively charged **cation** (net positive ionic charge). When it has gained one or more electrons, it forms a negatively charged **anion** (net negative ionic charge).

This process is achieved by the easiest route possible. For example, aluminium has a valency of 3 ([Ne] $3s^1 3p^2$). To have completely filled s and p sublevels, aluminium can either gain five electrons or lose three electrons. A loss of three electrons requires less energy, and therefore is preferred. This gives the aluminium a 3+ ionic charge.

Table 3 shows the general trend for ionic charge across the periodic table.

ionic charge
the difference between the positive charges from the protons and the negative charges from the electrons, represented as a superscript on the right-hand side of an elemental symbol

cation
a positively charged ion

anion
a negatively charged ion

Study tip

For charges greater than 1 (positive or negative), the size of the charge is written before the sign of the charge, e.g. Ca^{2+}, O^{2-} and Fe^{3+}.

TABLE 3 General trend for ionic charge

Group	Ionic charge
1	1+
2	2+
13	3+
14	Valency = 4 (non-metals such as carbon and silicon generally share electrons rather than forming an ion)
15	3–
16	2–
17	1–
18	0

The ionic charge of an element determines its **ionic radius**. Positive ions are smaller (have a smaller ionic radii) than their neutral atom counterparts because they have lost space-occupying electrons (Figure 2). In other words, ionic radii of cations are smaller than the atomic radii of the corresponding uncharged atoms.

ionic radius
the radius of an ion

Negative ions are larger (have a larger ionic radii) than their neutral atom counterparts because they have gained space-occupying electrons. In other words, the ionic radii of monatomic anions are larger than the atomic radii of the corresponding uncharged atoms.

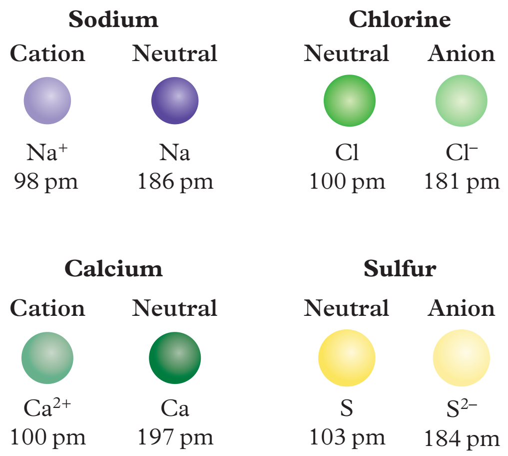


FIGURE 2 Cations are smaller than the neutral atoms. Anions are larger than the neutral atoms. Ionic radii are shown in picometres (pm).

Ionic radii and the periodic table

Ionic radii generally decrease across a period for the cations. This is because as atomic number increases, the number of protons increase, therefore increasing the nuclear charge. There is a stronger attractive force pulling valence electrons towards the nucleus, so the ionic radii decreases.

For elements on the right of the periodic table, anions form. The ionic radii increase significantly for the first anion in each period. This is followed by a decrease. Compare Be^{2+} with N^{3-}: Be^{2+} has just two electrons held close to the nucleus in the first energy level, whereas N^{3-} has 10 electrons, filling both the first and second energy level. The outer eight electrons will not be as strongly attracted to the nucleus and are much further from it (Figure 3).

Down a group, ionic radii generally increases as the electrons occupy more energy levels, with the outer levels being further from the nucleus. This is very similar to the explanation for atomic radii.

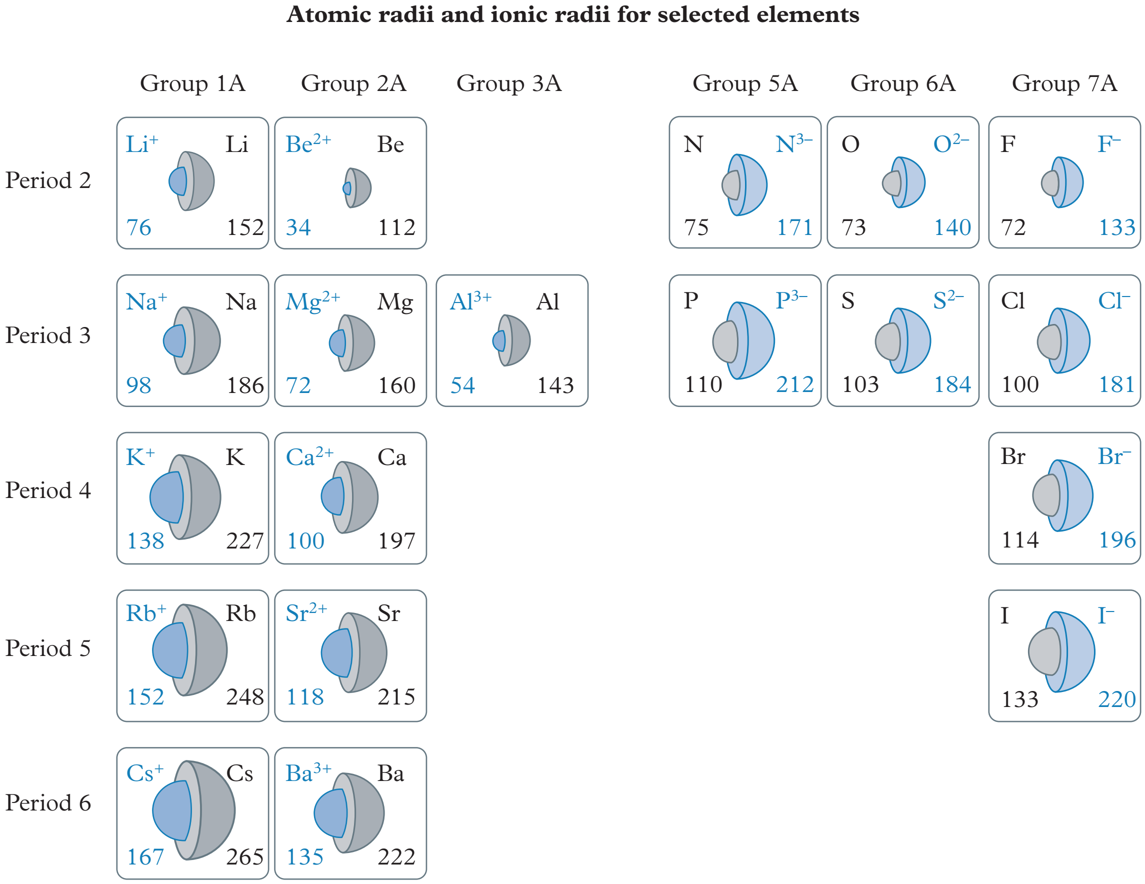


FIGURE 3 Atomic radii (grey) and ionic radii (blue) in picometres (pm); ionic radii decrease across a period (row) for cations. This is followed by a large increase in radius for the first anion in the period, followed by a decrease.

What is ionisation energy?

Ionisation energy (IE) is the energy required to remove an electron from a gaseous atom, forming a positively charged cation. For example:

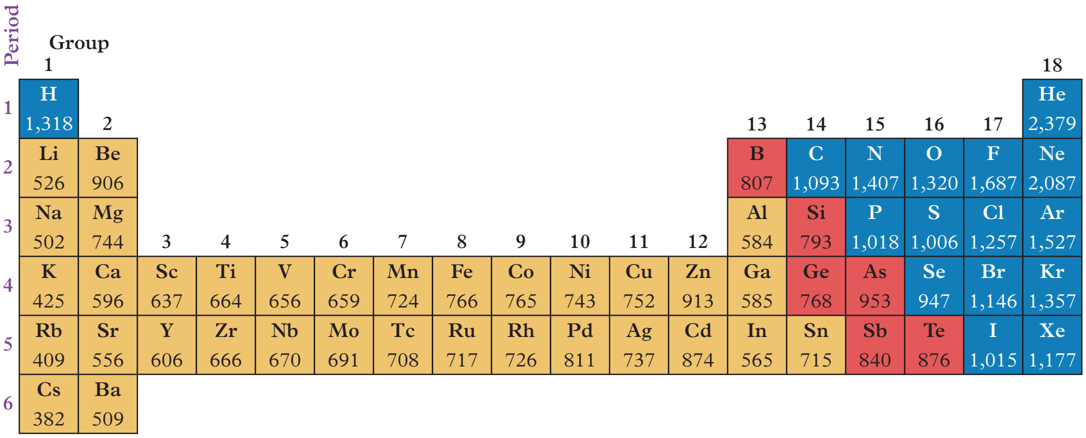
$$K(g) \rightarrow K^{+}(g) + e^{-}$$

ionisation energy
the energy (in kJ mol^{-1}) required to remove one mole of electrons from one mole of atoms in the gaseous phase

Ionisation energy and the periodic table

IE increases across a period because the atom has significantly more electrostatic attraction between its nucleus and valence electrons. For example, sodium has an IE of 495 kJ mol^{-1} and chlorine has an IE of 1257 kJ mol^{-1}. Chlorine has significant electrostatic attractions to its seven valence electrons, and tends to become an ion by gaining an electron to have a complete p sublevel, rather than losing an electron to get further away from stability.

IE decreases going down a group because of the electron shielding effect (Figure 4). With each energy level that is added to an atom, there is additional distance between the negative valence electrons and the positive nucleus. The additional distance causes a weaker electrostatic attraction between the nucleus and the valence electrons.



Group

Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1,318																	He 2,379
2	Li 526	Be 906											B 807	C 1,093	N 1,407	O 1,320	F 1,687	Ne 2,087
3	Na 502	Mg 744											Al 584	Si 793	P 1,018	S 1,006	Cl 1,257	Ar 1,527
4	K 425	Ca 596	Sc 637	Ti 664	V 656	Cr 659	Mn 724	Fe 766	Co 765	Ni 743	Cu 752	Zn 913	Ga 585	Ge 768	As 953	Se 947	Br 1,146	Kr 1,357
5	Rb 409	Sr 556	Y 606	Zr 666	Nb 670	Mo 691	Tc 708	Ru 717	Rh 726	Pd 811	Ag 737	Cd 874	In 565	Sn 715	Sb 840	Te 876	I 1,015	Xe 1,177
6	Cs 382	Ba 509																

FIGURE 4 First ionisation energies (kJ mol^{-1}) of some elements

Worked example 2.6A

Predicting ionisation energies using the periodic table

Worked example 2.6A: Watch a video that shows how to solve this problem.

Predict whether magnesium would have a greater IE than phosphorus. **Justify** your response. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Predict" means to give an expected result. "Justify" means to give reasons or evidence to support an answer. In this question, we need to consider the trend in ionisation energy on the periodic table and compare the two elements. The question is worth 2 marks, so we must complete our analysis and provide a prediction and give our reasoning.
Step 2: Locate the elements on the periodic table.	Both are located in period 3. Magnesium is in group 2 and phosphorus is in group 15.

Think	Do
Step 3: Recall the general trend across periods (between groups).	Across a period, IE increases as there is a greater electrostatic attraction between the nucleus and valence electrons. This is because as atomic number increases, the nuclear charge and number of valence electrons both increase and therefore, the attractive forces increase. (1 mark)
Step 4: Apply the trend to the elements in the question.	Phosphorus will have a greater IE than magnesium, as it has a larger number of valence electrons and a greater nuclear charge. (1 mark)

Your turn

Predict whether boron or indium would have the higher IE. **Justify** your response. (2 marks)

Successive ionisation energies

An atom's first IE is the energy required to remove one electron from its valence energy level. An atom's second IE is the energy required to remove a second electron from its valence energy level.

The valency of an atom has a significant impact on the differences between an atom's successive ionisation energies. Alkali metals have one valence electron, so they have relatively small ionisation energies to remove the valence electron, which results in an ionic charge of +1 and achieves complete valence sublevels (with the electron configuration of the previous noble gas). This first electron has the highest energy and is the furthest away from the nucleus, so it is the easiest to remove. Consider the first ionisation equation for aluminium, where I represents the energy required to remove the electron, in kJ mol^{-1}.

$$Al(g) \rightarrow Al^{+}(g) + e^{-} \qquad I_1 = 580\,kJ\,mol^{-1}$$

A second IE is significantly larger because an electron is being removed from a sublevel that is full and stable, and there are greater forces of electrostatic attraction because the electrons are closer to the nucleus. The second ionisation equation for aluminium is shown; this second electron requires more than three times the energy needed to remove.

$$Al^{+}(g) \rightarrow Al^{2+}(g) + e^{-} \qquad I_2 = 1{,}815\,kJ\,mol^{-1}$$

With each successive IE, the value increases again, as the removal of more electrons means that there is an increasing imbalance between the number of protons and the number of electrons remaining, and each electron is experiencing a greater positive charge. These electrons are closer to the nucleus and a greater electrostatic attraction must be overcome to remove them.

$$Al^{2+}(g) \rightarrow Al^{3+}\ (g) + e^{-} \qquad I_2 = 2{,}740\,kJ\,mol^{-1}$$
$$Al^{3+}(g) \rightarrow Al^{4+}\ (g) + e^{-} \qquad I_2 = 11{,}600\,kJ\,mol^{-1}$$

Alkaline earth metals have relatively small values for the first two ionisation energies, as they develop the same electron configuration of the previous noble gas. However, the third IE is greatly increased in value. Group 13 has three small ionisation energies and then a large fourth IE. This trend continues along the p block of the periodic table, aligning the number of small ionisation energies an atom has with the number of valence electrons it has (Table 4).

TABLE 4 Successive ionisation energies (kJ mol^{-1}) of some elements in the third period

Element	First IE	Second IE	Third IE	Fourth IE	Fifth IE	Sixth IE	Seventh IE
Na	502	4,560					
Mg	744	1,445	7,730				
Al	584	1,815	2,740	11,600			
Si	793	1,575	3,220	4,350	16,100		
P	1,018	1,890	2,905	4,950	6,270	21,200	
S	1,006	2,260	3,375	4,565	6,950	8,490	27,000
Cl	1,257	2,295	3,850	5,160	6,560	9,360	11,000
Ar	1,527	2,665	3,945	5,770	7,230	8,780	12,000

Skill drill

Predicting ionisation energies

Science inquiry skills(s): Processing and evaluating data (Lesson 1.7)

Atoms with full s or d orbitals (group 2 and group 12) and atoms with half-filled outer p orbitals (group 15) have slightly higher first ionisation energies than the elements on either side of them. The first ionisation energies of the first 86 elements are shown in Figure 5, but data for period 5 is deliberately omitted.

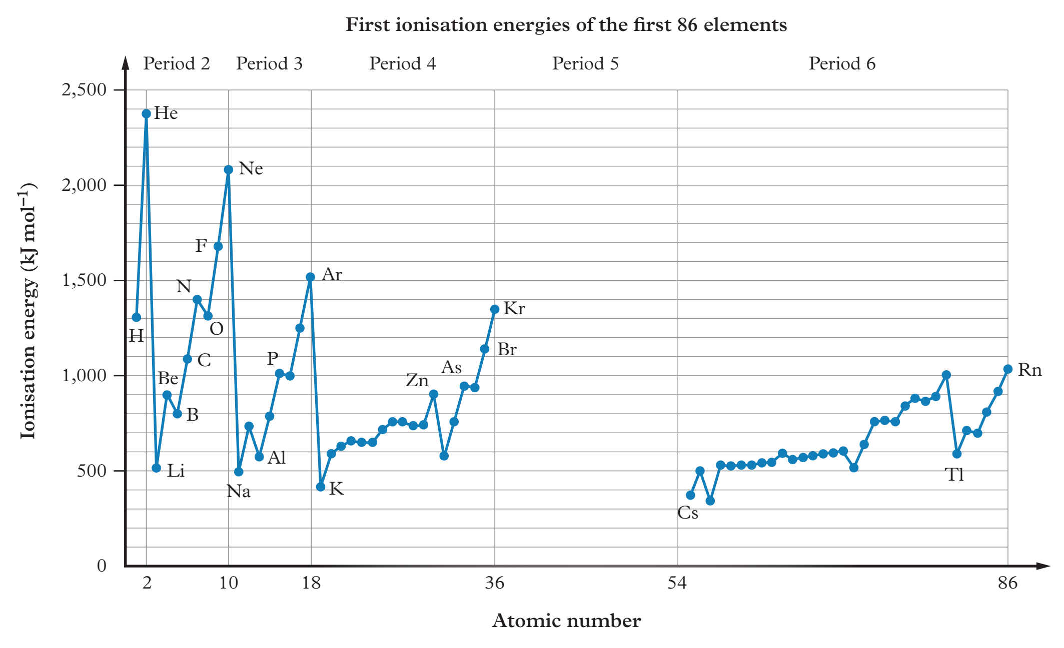


FIGURE 5 The first ionisation energies of the first 86 elements, excluding period 5

Practise your skills

1 **Predict** the approximate values for the first ionisation energies of period 5 elements: xenon, antimony and indium. (3 marks)

2 **Justify** your answers to question **1**, with reference to number of electrons and electrostatic attraction. (3 marks)

What is electronegativity?

electronegativity
the attraction of a positively charged nucleus to negatively charged electrons of a neighbouring atom

We have talked a lot about the electrostatic interactions between the nucleus and valence electrons within a single atom. There are also electrostatic forces between atoms in compounds. The ability of a positive nucleus to electrostatically attract the valence electrons of an adjacent bonded atom is called **electronegativity**. If electronegativity is large, an atom can draw a neighbouring electron closer to itself to share that electron, or even gain the electron for itself (Figure 6). A small electronegativity hinders the atom's ability to share or gain neighbouring electrons.

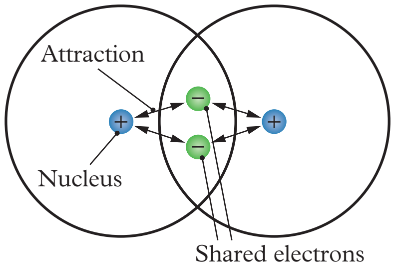
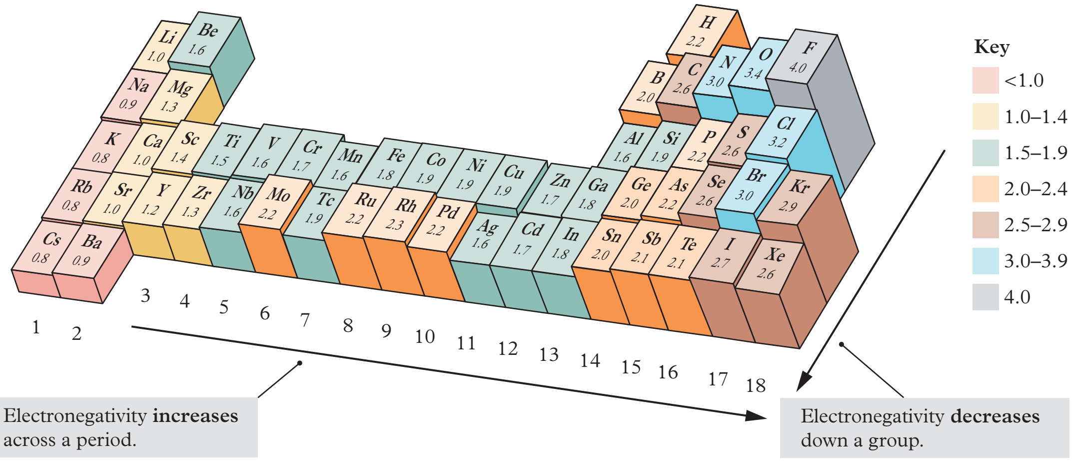


FIGURE 6 Electronegativity allows a positive nucleus to attract negative electrons in a bonded atom, allowing them to share electrons and become stable.

Electronegativity and the periodic table

Electronegativity increases across a period because of the electrostatic attractions caused by the increasing number of an atom's protons and valence electrons.

Electronegativity decreases down a group because of the electron shielding effect (Figure 7). As the distance increases between an atom's nucleus and its own valence electrons, the electrostatic attraction to the valence electrons and the electrons of the adjacent atom decreases.



Electronegativity **increases** across a period.

Electronegativity **decreases** down a group.

FIGURE 7 Electronegativity increases across a period and decreases down a group. Fluorine is the most electronegative element. Caesium and francium are the least electronegative. He, Ne and Ar do not have electronegativity values as these are calculated from the properties of compounds of elements. He, Ne and Ar do not form compounds.

What is metallic and non-metallic character?

metallic character
how readily an element loses valence electrons

non-metallic character
how readily an element gains valence electrons

Metallic character refers to the extent to which an element exhibits metallic properties, such as high melting and boiling points, conductivity, malleability and ductility. It also describes how readily an element loses valence electrons. As such, **non-metallic character** describes how readily an element gains valence electrons.

Metallic and non-metallic character and the periodic table

Metallic character increases from the top right to the bottom left of the periodic table and depends on the fact that most metals have large atomic radii and small electronegativities and ionisation energies. As a result, metals (Li to Cs) tend to lose electrons to other atoms when reacting, forming positive ions. The most reactive metal is francium; it is so reactive that many of its metallic properties have never been observed. Caesium, just above francium and the second highest in metallic character, is so reactive that it can spontaneously ignite in air.

The opposite trend is observed for non-metallic character – it increases from the bottom left to the top right of the periodic table. The halogens (F to I) have the greatest non-metallic character and are most likely to gain valence electrons.

Metallic and non-metallic character and oxide formation

Oxides are compounds that contain oxygen and another element, which can be a metal or non-metal. The element's metallic or non-metallic character determines what type of oxide the compound will be. As metallic character decreases across a period, oxides change from basic through amphoteric to acidic. You will learn more about acids and bases in Module 14.

FIGURE 8 Magnesium oxide is a basic oxide.

Acidic oxides

Acidic oxides form between oxygen and non-metals. Examples are carbon dioxide (CO_2), diphosphorus pentoxide (P_2O_5) and nitrogen dioxide (NO_2).

When an acidic oxide reacts with water, it forms an acid:

$$\underset{\text{acidic oxide}}{CO_2} + \underset{\text{water}}{H_2O} \rightarrow \underset{\text{acid}}{H_2CO_3}$$

When an acidic oxide reacts with a base it forms a salt and water:

$$\underset{\text{acidic oxide}}{CO_2} + \underset{\text{base}}{2NaOH} \rightarrow \underset{\text{salt}}{Na_2CO_3} + \underset{\text{water}}{H_2O}$$

Basic oxides

Basic oxides form between oxygen and metals. Examples are sodium oxide (Na_2O), magnesium oxide (MgO) and copper (II) oxide (CuO).

When a basic oxide reacts with water, it forms a base:

$$\underset{\text{basic oxide}}{Na_2O} + \underset{\text{water}}{H_2O} \rightarrow \underset{\text{base}}{2NaOH}$$

When a basic oxide reacts with an acid, it forms salt and water:

$$\underset{\text{basic oxide}}{Na_2O} + \underset{\text{acid}}{2HCl} \rightarrow \underset{\text{salt}}{2NaCl} + \underset{\text{water}}{H_2O}$$

Amphoteric oxides

Amphoteric oxides have both acidic and basic properties. They are metals that can have different ionic charges. Examples include copper, zinc, tin, lead, aluminium and beryllium.

When an amphoteric oxide, such as zinc oxide, reacts with an acid, it forms a salt:

$$\underset{\text{amphoteric oxide}}{ZnO} + \underset{\text{acid}}{2HCl} \rightarrow \underset{\text{salt}}{ZnCl_2} + \underset{\text{water}}{H_2O}$$

When an amphoteric oxide reacts with a base, it forms a salt:

$$\underset{\text{amphoteric oxide}}{ZnO} + \underset{\text{base}}{2NaOH} + \underset{\text{water}}{H_2O} \rightarrow \underset{\text{salt}}{Na_2ZnO_2}$$

FIGURE 9 Zinc oxide is used as an additive in many different products, such as sunscreen.

What are the periodic table trends?

Figure 10 summarises the trends in the periodic table that we have learnt.

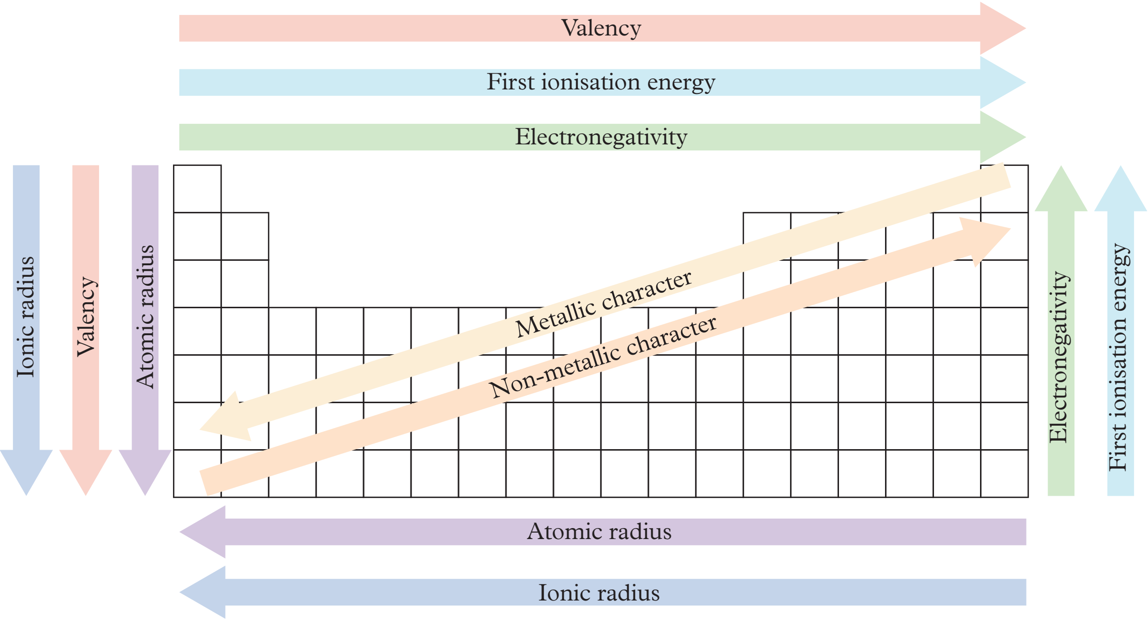


FIGURE 10 Summary of the periodic table trends

Check your learning 2.6

Check your learning 2.6: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** how an atomic radius differs from an ionic radius. (2 marks)

2 **Describe** the electron shielding effect. (2 marks)

3 **Recall** how oxides change moving across period 3. (1 mark)

Analytical processes

4 **Compare** metallic and non-metallic character, using two examples. Reference their positions on the periodic table in your answer. (4 marks)

5 **Determine** which elements have a greater ionisation energy.
   - **a** Al or Cl (1 mark)
   - **b** N or Sb (1 mark)
   - **c** Rb or As (1 mark)

6 **Determine** which elements have a greater atomic radius.
   - **a** K or Br (1 mark)
   - **b** Be or Sr (1 mark)
   - **c** Li or Te (1 mark)

7 **Determine** which elements have a greater electronegativity.
   - **a** Al or Cl (1 mark)
   - **b** O or F (1 mark)
   - **c** Na or N (1 mark)

8 **Analyse** the relationship between the location of elements on the periodic table and their electronegativities, using the data in Figure 7. (3 marks)

Knowledge utilisation

9 **Discuss** the relationship between valency and first ionsation energy. (4 marks)

Practical

Lesson 2.7

Investigating periodic table trends using a database

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

FIGURE 1 Copper is a metallic element. It is a great conductor of heat and electricity.

Lesson 2.8

Review: Atomic structure, isotopes and the periodic table

Summary

2.1
- An atom can be modelled as a nucleus surrounded by negatively charged electrons in distinct energy levels (and sublevels) held together by electrostatic forces of attraction between the nucleus and electrons.
- Inside the nucleus are the positively charged protons and neutral neutrons.
- Nuclear symbol notation ${}^{A}_{Z}M$ summarises the number of subatomic particles in atoms, ions and isotopes.

2.2
- Practical: Simulating Geiger–Marsden's gold foil experiment.

2.3
- Isotopes are atoms of an element that have different numbers of neutrons.
- Isotopes have similar chemical properties by different physical properties.
- Relative atomic mass is a ratio of the weighted average of the atomic masses of all the isotopes of all the naturally occurring isotopes of an element (taking into account their natural abundances on Earth) to one-twelfth the atomic mass of the ${}^{12}C$ isotope.

2.4
- Elements on the periodic table are organised into vertical groups and horizontal columns and into other sections such as metals and non-metals.
- The location of an element is linked to its properties.

2.5
- Electrons are locations within energy levels, which are organised further into sublevels and orbitals (s, p, d and f).
- Electron configurations can be represented in full or condensed formats. They are used to represent the location of the electrons in an atom or ion.
- The Aufbau principle, Pauli exclusion principle and Hund's rule are used to write full and condensed electron configurations.

2.6
- Atomic radii and metallic character decrease across the period and increase down the group.
- Valency increases across the period and stays consistent down the group.
- Ionic radii are smaller than atomic radii when the ion is positive, and larger than atomic radii when the ion is negative.
- Ionisation energy, electronegativity and non-metallic character generally increase across a period and decrease down a group.

2.7
- Practical: Investigating periodic trends using a database.

Key formulas

Protons	$p^+ = Z$
Neutrons for atoms and ions	$n^0 = A - Z$
Neutrons for isotopes only	n^0 = atomic mass – Z
Electrons for atoms and isotopes	$e^- = Z$
Electrons for ions only	$e^- = Z$ – ionic charge

Review questions 2.8A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Identify the number of protons, electrons and neutrons respectively, based on the nuclear symbol $^{9}_{4}\text{Be}$.

A 4, 5, 5
B 4, 4, 4
C 4, 4, 5
D 4, 5, 4

2 What is the number of protons, electrons and neutrons respectively, based on the nuclear symbol $^{24}_{12}\text{Mg}^{2+}$?

A 12, 12, 12
B 12, 10, 12
C 12, 14, 12
D 12, 12, 24

3 Cobalt has only one stable isotope which determines the relative atomic mass for this element. The correct nuclear symbol for cobalt is

A $^{58}_{27}\text{Co}$.
B $^{27}_{59}\text{Co}$.
C $^{12}_{6}\text{C}$.
D $^{59}_{27}\text{Co}$.

4 Identify the correct element based on the full electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$.

A Niobium
B Vanadium
C Phosphorus
D Titanium

5 Which isotope of carbon contains seven neutrons?

A ^{11}C
B ^{12}C
C ^{13}C
D ^{14}C

6 Which characteristics are similar between isotopes of the same element?

A Number of neutrons
B Density
C Electron configurations
D Boiling and melting points

7 Relative atomic mass is calculated by

A dividing the number of isotopes of an element by 12, relative to carbon.
B counting the number of isotopes that exist for an element.
C averaging the mass numbers of isotopes that exist for an element.
D taking the weighted average of the atomic masses of the isotopes for an element and expressing it as a ratio to one-twelfth the atomic mass of carbon-12.

8 The RAM of Cl is

A 17.
B 18.
C 35.45.
D 37.

9 The condensed electron configuration of Cr is

A [Ar] $4s^2 3d^4$
B [Ar] $4s^1 3d^5$
C [Ne] $3s^2 3p^6 3d^4 4s^2$
D [Ne] $3s^2 3p^6 3d^5 4s^1$

10 Which of the terms should be used to complete the following sentence:
As you go across a period in the periodic table, electronegativity __________, and the first ionisation energy generally ______________.

A increases; increases
B increases; decreases
C decreases; decreases
D decreases; increases

11 Which of the following correctly describes the properties of potassium, K, compared to calcium, Ca?

A Higher valency, lower atomic radius
B Higher valency, higher atomic radius
C Lower valency, lower atomic radius
D Lower valency, higher atomic radius

12 Based on its position on the periodic table, oxides of sulfur, S, are expected to be

A basic.
B amphoteric.
C acidic.
D neutral.

13 Hund's rule states that

- **A** every orbital can accept a maximum of two electrons.
- **B** orbitals with lower energy are completely filled before orbitals with higher energy.
- **C** when filling a sublevel, one electron occupies each orbital before any orbitals are occupied by two electrons.
- **D** the total number of electrons in an atom is equal to the atomic number and this is the same as the number of protons.

14 Which lists properties that all increase across a row of the periodic table?

- **A** Basic nature of the oxide, atomic radius, ionic radius
- **B** Electronegativity, first ionisation energy, acidic nature of the oxide
- **C** Atomic radius, acidic nature of the oxide, electronegativity
- **D** Ionic radius, basic nature of the oxide, first ionisation energy

15 As electrons are removed from an atom, the successive ionisation energies

- **A** decrease, because it requires less energy to remove subsequent electrons from an ion.
- **B** increase, because it requires less energy to remove subsequent electrons from an ion.
- **C** decrease, because it requires more energy to remove subsequent electrons from an ion.
- **D** increase, because it requires more energy to remove subsequent electrons from an ion.

Review questions 2.8B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

16 **Describe** the structure of an atom. (1 mark)

17 **Recall** what atomic number and mass number is. (2 marks)

18 **Identify** how isotopes of the same element differ. (1 mark)

19 **Explain** why chromium and copper do not follow the standard electron configuration conventions. (2 marks)

20 **Recall** the relative energy levels of the electron subshell orbitals. (1 mark)

21 **Define** the terms

- **a** "electronegativity" (1 mark)
- **b** "ionisation energy". (1 mark)

22 **Calculate** the RAM for the following:

- **a** copper, which is 69.17% ^{63}Cu with an atomic mass of 62.93 and 30.83% ^{65}Cu with an atomic mass of 64.92 (2 marks)
- **b** silicon, which is 92.23% ^{28}Si with an atomic mass of 27.98, 4.68% ^{29}Si with an atomic mass of 28.98 and 3.09% ^{30}Si with an atomic mass of 29.97. (2 marks)
- **c** nitrogen, which is 99.63% ^{14}N with an atomic mass of 14.00 and 0.37% ^{15}N with an atomic mass of 15.00 (2 marks)
- **d** potassium, which is 93.6% ^{39}K with an atomic mass of 38.96, 0.01% ^{40}K with an atomic mass of 9.96 and 6.73% ^{41}K with an atomic mass of 40.96. (2 marks)

23 90% of the atoms of an unknown element have an isotope with a mass number of 35, 8.0% of the atoms have a mass number of 37, and the remainder of the atoms have a mass number of 38. **Calculate** the RAM of the unknown element. (2 marks)

Analytical processes

24 **Deduce** the nuclear symbol notations for the following atoms and ions. The number of neutrons is specified in brackets.

- **a** Sodium cation (12) (1 mark)
- **b** Fluorine atom (10) (1 mark)
- **c** Cobalt atom (32) (1 mark)
- **d** Titanium atom (26) (1 mark)
- **e** Chloride anion (20) (1 mark)

25 **Deduce** the full electron configurations for the following atoms.

a P (1 mark) **b** Ca (1 mark) **c** Cu (1 mark)

26 **Derive** the condensed electron configurations from the following full electron configurations.

- **a** $1s^2 2s^2 2p^6 3s^2 3p^3$ (1 mark)
- **b** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ (1 mark)
- **c** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ (1 mark)
- **d** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ (1 mark)

27 **Determine** the relative abundances of the isotopes of silver using a RAM of 107.868. ^{107}Ag has an atomic mass of 106.905 and ^{109}Ag has an atomic mass of 108.905. (3 marks)

28 There are two adjacent elements located in period 4 of the periodic table. The first element has a completed third energy level (including all of the sublevels) and the second element has at least one p electron in an unfilled shell. **Determine** the identity of the elements. (2 marks)

29 **Identify errors** in the following electron configurations. **Explain** why they are errors.

- **a** $1s^1 2s^2 2p^4$ (2 marks)
- **b** $1s^2 2s^2 2p^6 3s^3 3p^6 4s^2$ (2 marks)
- **c** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^2$ (2 marks)

30 **Determine** the correct condensed electron configurations for the electron configurations in question **29**. (3 marks)

Knowledge utilisation

31 The full electron configuration given for nitrogen is $1s^2 2s^2 2p^3$. This represents nitrogen as a neutral atom. If nitrogen was ionised to N^{3-}, **determine** the full electron configuration that would best represent this. (1 mark)

32 **Discuss** whether there is a relationship between the atomic radius and the ionic radius of an ion formed from that atom. (2 marks)

Data drill

Atomic number, atomic radii and melting points

The order in which elements are displayed on the periodic table is based on their electron configuration and their properties.

Some students decide to investigate the relationship between atomic number, atomic radii and melting points of some alkali metals: lithium, sodium, potassium and caesium. They collect some secondary data and present them in Table 1.

TABLE 1 Atomic numbers, atomic radii and melting points of four alkali metals

Element	Atomic number	Melting point (°C)	Atomic radius (pm)
Lithium	3	180	134
Sodium	11	98	154
Potassium	19	63	196
Caesium	55	29	225

Apply understanding

1 **Identify** the element with the lowest melting point. (1 mark)

2 **Calculate** the range of the atomic radii collated. (1 mark)

Analyse data

3 **Sketch** graphs of the melting point and atomic radius data from Table 1. (4 marks)

4 **Identify** one trend in the data. (1 mark)

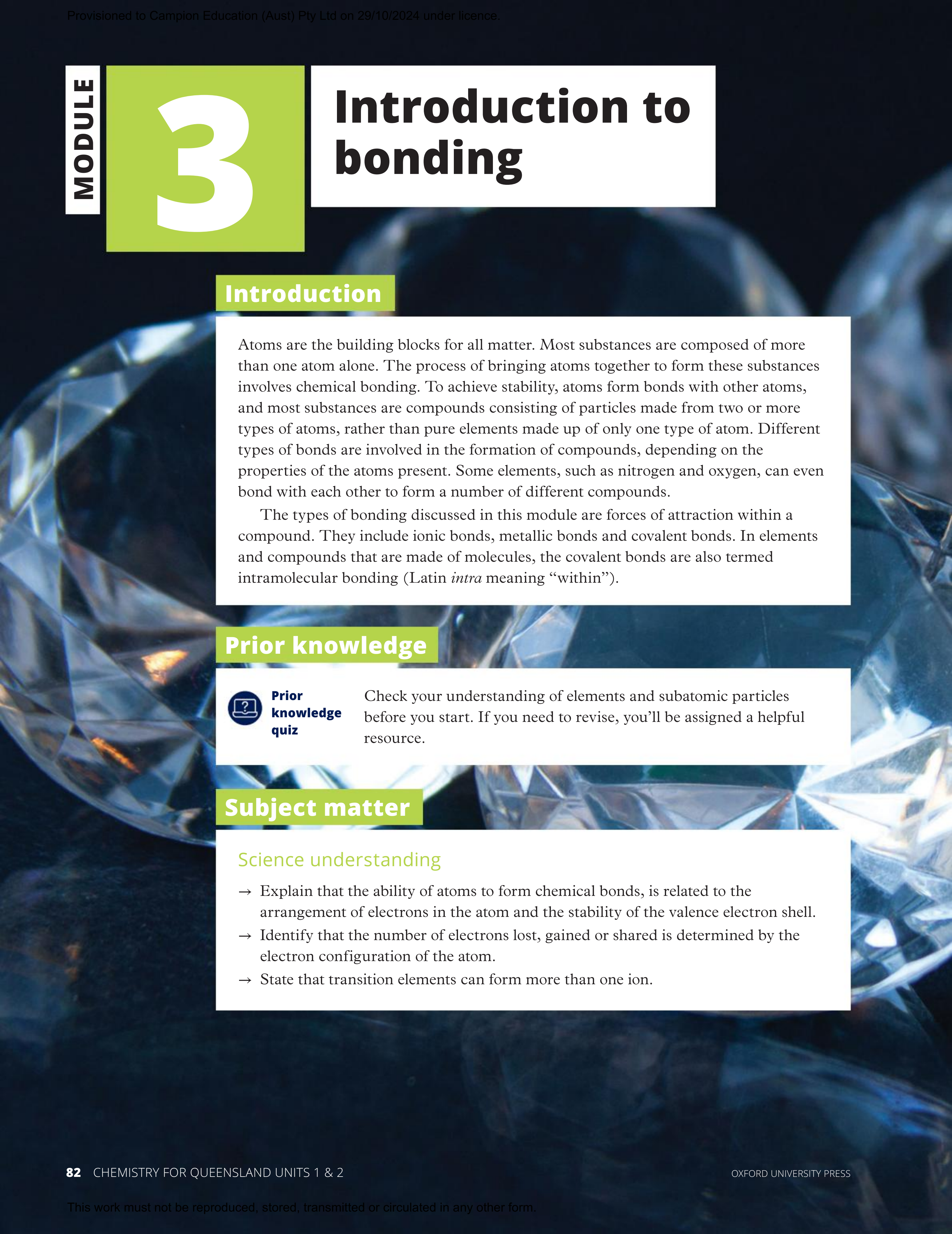
Interpret evidence

5 **Infer** a reason for the trend identified in question **4**. (1 mark)

6 **Predict** the melting point of the element rubidium ($Z = 37$). **Justify** your response. (2 marks)

Module 2 checklist: Atomic structure, isotopes and the periodic table

Quizlet: Revise key terms online to test your understanding



MODULE 3 Introduction to bonding

Introduction

Atoms are the building blocks for all matter. Most substances are composed of more than one atom alone. The process of bringing atoms together to form these substances involves chemical bonding. To achieve stability, atoms form bonds with other atoms, and most substances are compounds consisting of particles made from two or more types of atoms, rather than pure elements made up of only one type of atom. Different types of bonds are involved in the formation of compounds, depending on the properties of the atoms present. Some elements, such as nitrogen and oxygen, can even bond with each other to form a number of different compounds.

The types of bonding discussed in this module are forces of attraction within a compound. They include ionic bonds, metallic bonds and covalent bonds. In elements and compounds that are made of molecules, the covalent bonds are also termed intramolecular bonding (Latin *intra* meaning “within”).

Prior knowledge

Prior knowledge quiz

Check your understanding of elements and subatomic particles before you start. If you need to revise, you’ll be assigned a helpful resource.

Subject matter

Science understanding

→ Explain that the ability of atoms to form chemical bonds, is related to the arrangement of electrons in the atom and the stability of the valence electron shell.

→ Identify that the number of electrons lost, gained or shared is determined by the electron configuration of the atom.

→ State that transition elements can form more than one ion.

→ Explain that ions are atoms or groups of atoms that are electrically charged due to an imbalance in the number of electrons and protons.
→ Explain that chemical bonds are caused by electrostatic attractions that arise because of the sharing or transfer of electrons.
→ Identify that the valency is a measure of the number of bonds that an atom can form.
→ Determine the formula and IUPAC name of ionic and molecular compounds.
→ Determine Lewis (electron dot) structures 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.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Lesson 3.1

Types of bonding

Key ideas

- → Electrostatic attractions cause atoms to gain, lose or share electrons with each other to form compounds.
- → When a metal atom reacts with a non-metal atom, a transfer of electrons can take place to form ions and then an ionic compound.
- → Covalent bonding is the sharing of electrons between atoms.
- → Polyatomic ions are covalently bonded molecules that are electrically charged.

Learning intentions and success criteria

Why are valence electrons important?

The electron configuration of atoms can be separated into two basic components: the inner electrons and the valence (outer) electrons. Many of an element's properties result from the number of valence electrons in the element. For example, the reactivity of an element is determined by its electrostatic attraction to adjacent atoms. Its valence electrons may be gained from, lost to or shared with these adjacent atoms to come to a more stable arrangement of atoms.

octet rule
atoms gain and lose electrons to reach eight electrons in their valence shell; exceptions to this rule are hydrogen and helium

The octet rule

The **octet rule** refers to the tendency of atoms to lose, gain or share electrons in order to achieve eight electrons in the valence energy level, which comprises the highest occupied s and p sublevels. In doing so, atoms acquire the same electron configuration as the nearest noble gas. They therefore become more stable than when their octets were incomplete. Atoms do this by forming chemical bonds, which are essentially forces that hold atoms together in compounds and molecules. Specifically these bonds are categorised as either being ionic or covalent bonds.

Study tip

Exceptions to the octet rule are hydrogen and helium, which obey a "duet" rule with two electrons. Boron and beryllium also often form compounds with fewer than 8 electrons in their valence shell. Third period elements and beyond will either meet or exceed the octet rule.

How does ionic bonding occur?

ionic bonding
the process by which atoms transfer valence electrons to each other; the electrostatic attraction between cations and anions to form an ionic compound

Ionic bonding is the process by which atoms transfer valence electrons to each other. It is the result of electrostatic attractions between oppositely charged ions within a compound. Ionic bonding occurs between positively charged cations and negatively charged anions.

How might ions form? One way is through the transfer of electrons. If an atom gives away one electron, it will develop a positive ionic charge because of the imbalance of positive protons and the negative electrons. For example, sodium has 11 protons and 11 electrons. When sodium donates one electron, it develops a +1 ionic charge because it now only has 10 electrons, but 11 protons. The extra proton provides the excess positive charge. Similarly, the loss of two electrons by group 2 elements, such as magnesium, results in a +2 ionic charge, and the loss of three electrons by group 3 elements results in a +3 ionic charge. Positive ions are called cations.

Unlike metals, non-metallic elements tend to gain electrons from other atoms. They form negatively charged ions called anions. A gain of one electron by an atom causes a –1 ionic charge due to the extra electron. For example, chlorine has 17 protons and 17 electrons.

When it gains an electron, it has 18 electrons, but still 17 protons. The extra electron gives it a –1 ionic charge.

Whether an atom loses or gains electrons and how many it loses or gains can be predicted by its position on the periodic table. Table 1 shows how elements from the second period of the periodic table acquire the electron configuration of either helium or neon, according to the octet rule. Each element tends to either gain or lose electrons to acquire eight electrons, depending on the number of electrons it has in its valence shell and whichever process requires the least amount energy.

TABLE 1 Determining ionic charges in the second row of the periodic table

Element	Number of valence electrons	Lose or gain electrons?	Ionic charge
Li	1 ($2s^1$)	✓ Lose 1 to be like He ✗ Gain 7 to be like Ne	1+
Be	2 ($2s^2$)	✓ Lose 2 to be like He ✗ Gain 6 to be like Ne	2+
B	3 ($2s^2$ and $2p^1$)	✓ Lose 3 to be like He ✗ Gain 5 to be like Ne	3+
N	5 ($2s^2$ and $2p^3$)	✓ Lose 5 to be like He ✗ Gain 3 to be like Ne	3–
O	6 ($2s^2$ and $2p^4$)	✓ Lose 6 to be like He ✗ Gain 2 to be like Ne	2–
F	7 ($2s^2$ and $2p^5$)	✓ Lose 7 to be like He ✗ Gain 1 to be like Ne	1–

For example, sodium can give away an electron to achieve the electron configuration of neon, but chlorine can accept one electron to achieve the electron configuration of argon. When sodium and chlorine are combined, they form an **ionic compound** called sodium chloride, which is held together through the formation of sodium cations that are attracted to the chloride anions. Both ions in sodium chloride have full valence shells – that is, they have completed their octets (Figure 1).

ionic compound
a chemical compound that is held together by ionic bonds

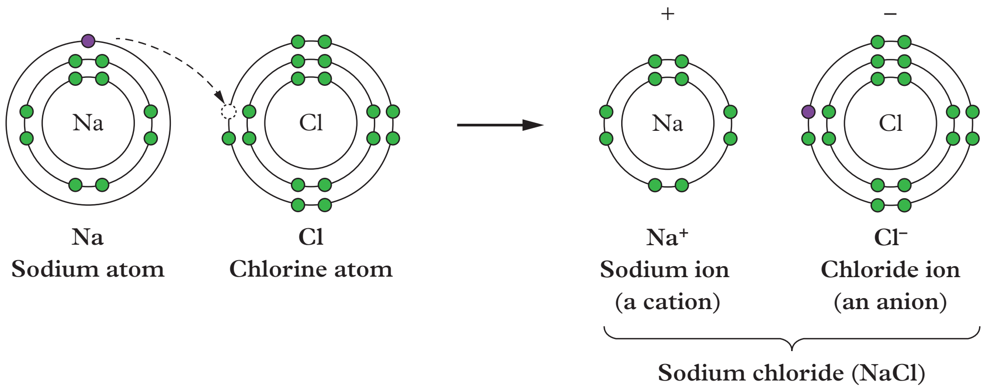


FIGURE 1 A sodium atom loses an electron to a chlorine atom to form the more stable sodium chloride.

We can also predict whether ionic bonding occurs, quantitatively. An ionic electrostatic attraction is likely to occur when elements have a difference in electronegativity of greater than 1.8. Consider sodium chloride. Sodium has an electronegativity of 0.9 and chlorine has an electronegativity of 3.2. The difference is 2.3, which is greater than 1.8. Therefore, we predict that sodium and chlorine combine to form a compound held together by ionic electrostatic attractions.

Study tip

Use the groups in the periodic table to remember common ionic charges. Group 1 is 1+, group 2 is 2+, group 13 is 3+, group 15 is 3–, group 16 is 2– and group 17 is 1–.

Ions of transition elements

Transition elements can form more than one ion. This is because they have partially filled d orbitals, allowing them to lose different numbers of electrons. The result is cations with different ionic charges, which are difficult to predict using the octet rule. An example is iron, Fe. Neutral Fe has the electron configuration $1s^22s^22p^63s^23p^63d^64s^2$. It can lose electrons from the 4s orbital alone or the 4s and partially filled 3d orbital to produce Fe^{2+} or Fe^{3+}, respectively. Other common examples of transition elements that produce more than one ion are copper (which forms Cu^+ and Cu^{2+}) and chromium (which forms Cr^{2+} and Cr^{3+}).

Polyatomic ions

An ion is an atom or group of atoms that has an electric charge. When an ion has only one atom, it is called a monatomic ion. When it contains two or more atoms (of the same or different elements) chemically combined, it is called a **polyatomic ion**. For example, OH^- is a polyatomic ion called the hydroxide ion. Some common polyatomic ions are listed in Table 2. The naming of ions follows rules set by the International Union of Pure and Applied Chemistry (IUPAC).

polyatomic ion
an ion with two or more atoms

TABLE 2 Common polyatomic ions

Polyatomic ion	Name of ion	Polyatomic ion	Name of ion
NH_4^+	Ammonium	$Cr^2O_7^{2-}$	Dichromate
H_3O^+	Hydronium	HCO_3^-	Hydrogencarbonate
CH_3COO^-	Ethanoate (acetate)	OH^-	Hydroxide
CO_3^{2-}	Carbonate	NO_3^-	Nitrate
ClO_3^-	Chlorate	NO_2^-	Nitrite
ClO_2^-	Chlorite	PO_4^{3-}	Phosphate
ClO^-	Hypochlorite	SO_4^{2-}	Sulfate
CrO_4^{2-}	Chromate	SO_3^{2-}	Sulfite

How are ionic formulas written?

A monatomic ion is represented by the element symbol with the ionic charge as a superscript to the right of the element symbol. For example, chloride ions are written as Cl^-. If the ion is polyatomic, then the number of constituent atoms of each element is represented as a subscript to the right of each element symbol (but before the ionic charge superscript). For example, carbonate ions are written as CO_3^{2-}.

When representing an ionic compound as an ionic formula, the cation is written first, followed by the anion. For example, for the ionic compound sodium chloride:

Sodium cation: Na^+
Chloride anion: Cl^-
Sodium chloride: NaCl

The ionic compound does not have a superscript ionic charge because it is overall neutral.

The ratio of cations to anions in an ionic compound is not always one-to-one. Often, electrostatic attraction occurs between one cation and multiple anions, one anion and multiple anions, or even multiple anions and multiple cations. For example, let's consider the compound magnesium chloride. A magnesium ion has a 2+ ionic charge. To balance the 2+ charge, two chloride ions, each with a 1– charge, are required. The ionic formula would therefore be $MgCl_2$ with a subscript 2 to indicate that two chloride ions are present for each Mg^{2+} ion.

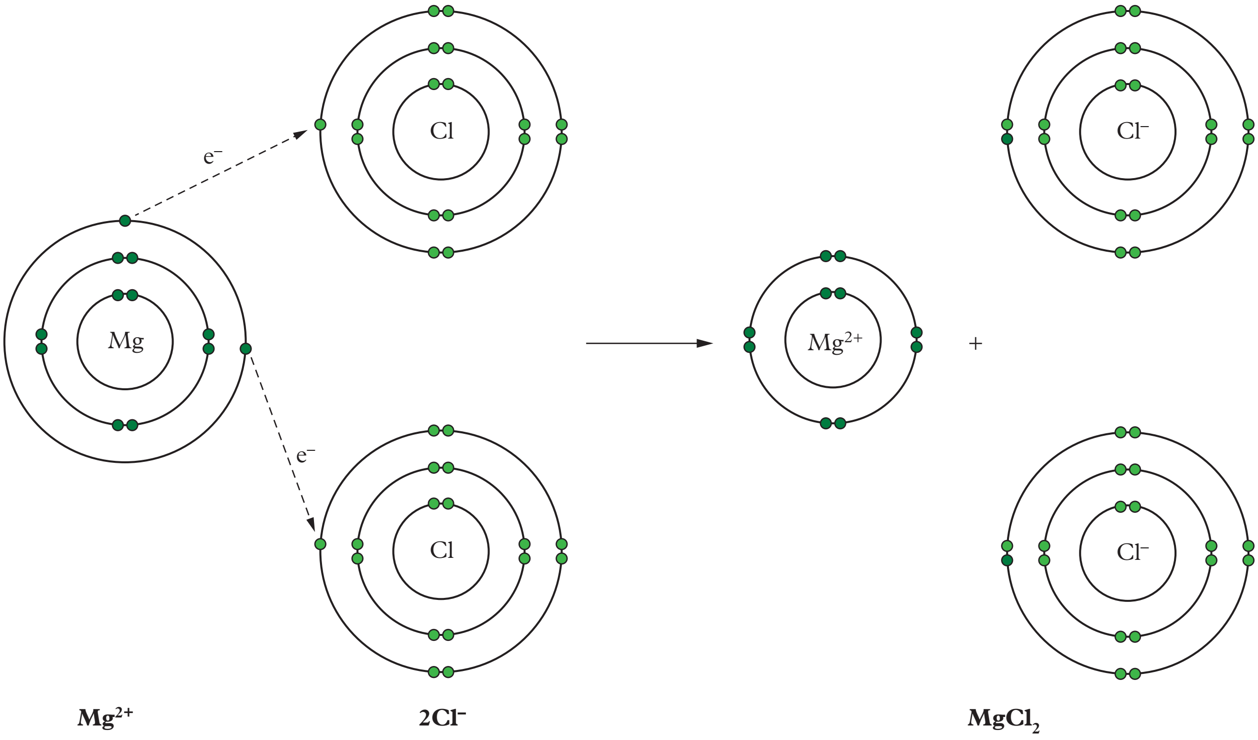


FIGURE 2 Magnesium chloride consists of one magnesium ion and two chloride ions.

The ionic formula can also be obtained using the swap-and-drop method (Figure 3). The ionic charge from the cation is "swapped" over to the anion and then "dropped" to be a subscript. The number loses its positive or negative symbol since it now represents the number of ions of that element in the ionic compound. The ionic charge does not need to be shown. The same process happens for the ionic charge from the anion. It is "swapped" to the cation and then "dropped" to be a subscript for the cation without a positive or negative charge.

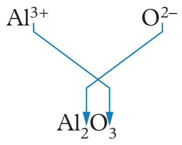


FIGURE 3 The swap-and-drop method for finding the ionic formula

Study tip

When a cation and anion have the same magnitude of charge (but opposite signs), the swap-and-drop method is not required and gives the wrong answer, if used! The ionic compound must be written as the simplest whole number ratio. For example, Ca^{2+} and SO_4^{2-} produce the ionic compound $CaSO_4$.

Worked example 3.1A

Determining ionic formulas

▶ **Worked example 3.1A:** Watch a video that shows how to solve this problem.

Determine the ionic formulas for the compounds formed when

a aluminium and oxygen bond together (1 mark)

b sodium and sulfate ions bond together (1 mark)

c phosphate ions and calcium bond together (1 mark)

d sulfate and calcium bond together. (1 mark)

▶

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to look at the atoms provided and the ions they form, then express the ionic compounds formed using their ionic formulas. The questions are worth 1 mark each, so we must analyse each set of atoms and then correctly express our answers.
Step 2: Determine the ions involved using the periodic table. For the polyatomic atoms, refer to Table 2. Write them with the cation first and then the anion.	**a** Al^{3+} O^{2-} **b** Na^{+} $SO_4^{\ 2-}$ **c** Ca^{2+} $PO_4^{\ 3-}$ **d** Ca^{2+} $SO_4^{\ 2-}$
Step 3: Apply the swap-and-drop method. Take the superscripts for the cation and anion, swap them, and drop them as subscripts. For polyatomic ions, place a bracket around the entire ion and swap and drop the superscript charge outside of the bracket. If the subscripts can be simplified to 1 : 1, you can exclude them.	**a** Al_2O_3 (1 mark) **b** Na_2SO_4 (1 mark) **c** $Ca_3(PO_4)_2$ (1 mark) **d** $CaSO_4$ (1 mark)

Your turn

Determine the ionic formula for ammonium phosphate. (1 mark)

Skill drill

Ionic bonding and melting points

Science inquiry skill(s): Processing and analysing data (Lesson 1.7)

Ionically bonded compounds form crystalline structures of positive ions, surrounded by negative ions. All the ions are attracted to the oppositely charged ions surrounding them. Therefore, these substances have relatively high melting points (Figure 4).

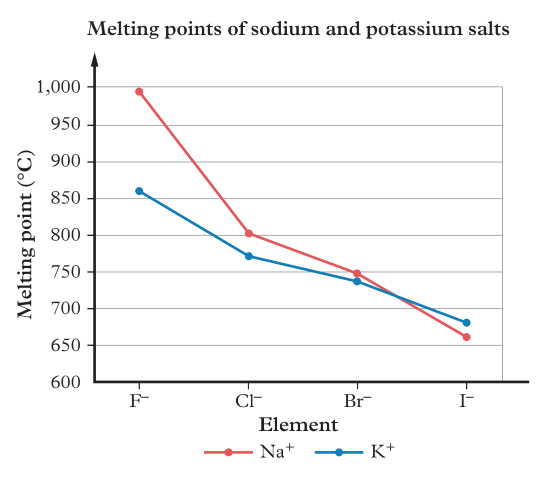


FIGURE 4 Melting points of ionic compounds of Na^+ and anions of group 17 elements, and of K^+ and anions of group 17 elements

Practise your skills

1 **Determine** which of the ionic bonds will be the strongest. **Justify** your answer. (2 marks)

2 **Use** the graph to **identify** the trend between melting point and ion size/ionic bonding. (1 mark)

How are ionic compounds named?

In the same way that ionic formulas are written, ionic compounds are named cation first and anion second. Metals maintain their elemental name as the cation, with d-block transition metals represented using Roman numerals in brackets depending on their charge in brackets, e.g. iron(II) cation for Fe^{2+}. Polyatomic ions also maintain their name. When the anion is monatomic, the last part of the name of the non-metal is removed and replaced with "-ide" (Table 3).

TABLE 3 Naming anions

Element name	Ionic formula	Base name	Anion name
Fluorine	F^-	Fluor	Fluoride
Chlorine	Cl^-	Chlor	Chloride
Oxygen	O^{2-}	Ox	Oxide
Sulfur	S^{2-}	Sulf	Sulfide
Nitrogen	N^{3-}	Nitr	Nitride
Phosphorus	P^{3-}	Phosph	Phosphide

Worked example 3.1B

Determining the name of ionic compounds

Worked example 3.1B: Watch a video that shows how to solve this problem.

Determine the name of the following ionic compounds:

a CaO (1 mark)

b Al_2S_3 (1 mark)

c Li_2CO_3. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to look at the ionic compounds and name them. The questions are worth 1 mark each, so we must analyse each ionic compound and then correctly name them.
Step 2: List the cation first, then the anion.	**a** Calcium oxygen **b** Aluminium sulfur **c** Lithium carbonate
Step 3: If the anion is monatomic, replace the last part of the name of the non-metals and replace it with "-ide".	**a** Calcium oxide (1 mark) **b** Aluminium sulfide (1 mark) **c** Lithium carbonate (1 mark)

Your turn

Determine the name of the ionic compound $(NH_4)_2S$. (1 mark)

Real-world chemistry

Ions in our body

The human body obtains a substantial amount of ions for crucial functions, such as the contraction of the heart, from food. An ion deficiency due to a lack of particular elements in our food has a significant effect on our bodies. Table 4 lists some of the more important ions we need every day. Trace amounts of other ions are also required by the body to function at capacity.

TABLE 4 Important ions in the human diet

Ion name	Ionic formula	Dietary sources	Associated functions	Effects of deficiency
Sodium	Na^+	Salt, additives, meat, fish, dairy, eggs, olives, pickled and smoked foods	Regulation and control of body fluid levels	Diarrhoea, anxiety, cardiac arrest
Potassium	K^+	Whole grains, meat, legumes, fruit and vegetables	Water balance in cells, muscle and nerve function, protein synthesis	Water imbalance, irregular heartbeat, cardiac arrest, tissue breakdown
Calcium	Ca^{2+}	Milk, leafy green vegetables, whole grains, egg yolk, legumes and nuts	Formation of bones and teeth, blood clotting, muscle function, transmission of nerve signals to cells	Rickets, osteoporosis, stunted formation of teeth
Magnesium	Mg^{2+}	Dairy, flour and cereal, dry beans, nuts, peas and leafy green vegetables	Nerve impulses, muscle contraction, enzyme activation	Tremors in muscles, heart spasms, convulsions
Chloride	Cl^-	Salt, seaweed, rye, tomatoes, lettuce, celery and olives	Balance of body fluid levels	Heavy sweating, chronic diarrhoea, vomiting
Phosphate derivatives	$H_2PO_4^-$ and HPO_4^{2-}	Dairy, cereal products and meat	Bone formation and energy production	Low energy, loss of appetite, numbing, burning or tingling extremities, joint pain, swelling and stiffness

Apply your understanding

1 **Evaluate** whether it would be possible to identify the type of ion deficiency from symptoms alone (i.e. without quantitatively measuring ion levels). (2 marks)

2 **Research** the specific roles that trace elements such as iron and zinc have in supporting bodily functions and **identify** the effects of their deficiencies. (4 marks)

How does covalent bonding occur?

covalent bonding
a sharing of electrons between atoms

molecule
a chemical species with no overall charge that is held together by covalent bonds

Covalent bonding occurs when atoms share electrons to form bonds. It is the type of bonding that we typically encounter within the chlorine molecule (Cl_2), ammonia compound (NH_3) and compounds such as water (H_2O) and phosphoric acid (H_3PO_4). Covalent bonds typically form between two or more non-metals and produce a covalent **molecule**, which can be an element or a compound. The number of bonds formed depends on valency of each of the atoms involved.

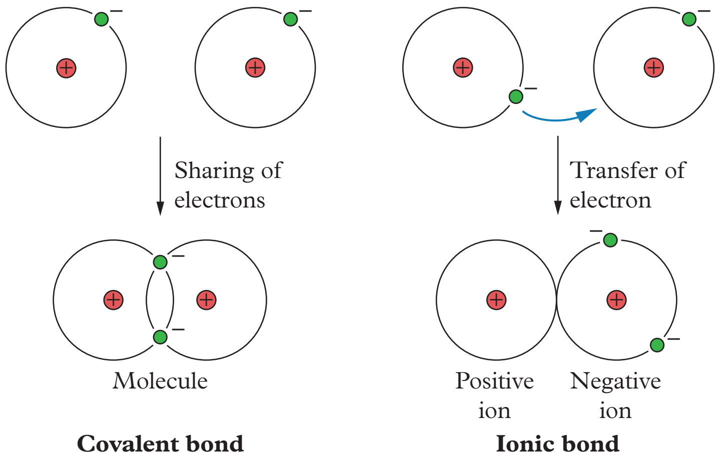


FIGURE 5 Covalent bonding and ionic bonding

For example, oxygen has six valence electrons in its outer shell – it has a valency of 6. It needs two more to complete its octet. We have already learnt that it can source two more electrons through ionic bonding, but it can also do this through covalent bonding by sharing two electrons with a neighbouring atom or atoms. In Figure 6, we can see in the presence of hydrogen, each hydrogen atom shares one electron to complete the octet for oxygen. By doing this, oxygen has achieved stability, in that it has acquired its required octet, and hydrogen being an exception to the octet rule, has also achieved stability.

Hydrogen is an exception to the octet rule and only needs two electrons in its valence shell to become stable. The electrostatic attraction between the shared pair of electrons between the each hydrogen atom and the positively charged nuclei of the one shared oxygen atom holds all three atoms together through two covalent bonds. Just like ionic compounds, we use subscripts to indicate how many atoms of each type are present in a covalent molecule. In this case, the resulting covalent molecule is H_2O, water.

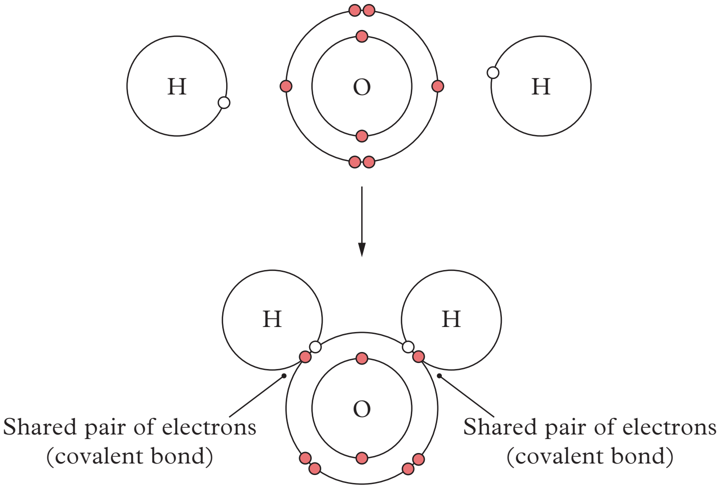


FIGURE 6 Oxygen shares electrons with two hydrogen atoms in order to complete its octet.

Types of covalent bonds

The electrons involved in covalent bonding are considered to belong to both of the atoms sharing the electrons. When investigating whether a molecule has followed the octet rule for each of its atoms, these shared electrons are counted twice.

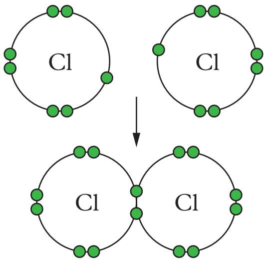


FIGURE 7 Two chlorine atoms share one pair of electrons to form chloride, Cl_2.

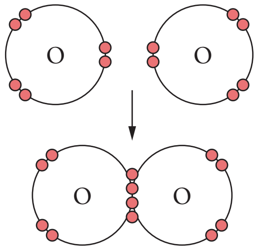


FIGURE 8 Two oxygen atoms share two pairs of electrons to form oxygen, O_2.

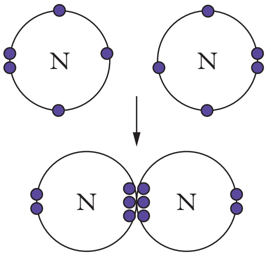


FIGURE 9 Two nitrogen atoms share three pairs of electrons to form nitrogen, N_2.

When two electrons (or one pair of electrons) are shared between two atoms, a single covalent bond has formed. This is what we observed in Figure 6. Another example is chlorine, which has a valency of seven. Each chlorine atom only needs one more electron to complete its outer shell, and it can do this by sharing one electron with another chlorine atom. This forms the element chlorine (Cl_2), made up of diatomic molecules, as shown in Figure 7. The electrons are shared to form a single covalent bond.

Sometimes, more than one pair of electrons must be shared to complete the outer shell and achieve stability. Take oxygen, for example. In Figure 6, we saw that it shared electrons with two hydrogen atoms to form a water molecule with two single bonds and complete its octet. It can also do this by sharing two electrons with another atom with the same valency, such as another oxygen atom. In this case, there are two pairs of electrons shared (four electrons total) and a double covalent bond is formed (Figure 8). This covalent molecule is the diatomic molecule oxygen, O_2.

Atoms can also share three pairs of electrons (six electrons total) to form a triple covalent bond. A common example is gaseous nitrogen, N_2 (Figure 9).

To determine the pattern of electron sharing, you need to consider the valency.

Nitrogen makes up approximately 78% of the air in the atmosphere but it does not readily take part in chemical reactions. This is because triple bonds are very strong. Thus, molecules containing multiple bonds such as double or triple bonds have stronger forces of electrostatic attraction between atoms.

Polyatomic ions are electrically charged groups of atoms that are held together by covalent bonding. For this reason, a compound can contain both ionic and covalent bonding if there is a polyatomic ion present. For example, in calcium carbonate ($CaCO_3$), the three oxygen atoms each have covalent bonds to the central carbon atom, forming the carbonate polyatomic ion. Ionic bonding occurs between negatively charged CO_3^{2-} ions and positively charged Ca^{2+} ions to make a compound that is overall electrically neutral.

Worked example 3.1C

Determining the pattern of valence electron sharing in covalent molecules

Worked example 3.1C: Watch a video that shows how to solve this problem.

Determine how valence electrons are shared between the atoms in the covalent molecule CF_4. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Determine” means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to look at the valency of each atom in the molecule, then find how the electrons are shared. The question is worth 1 mark, so we must analyse the atoms and provide the correct value.

Think	Do
Step 2: Use the periodic table to determine the number of valence electrons in each atom of the molecule.	Carbon is in group 4 so it has a valency of 4. Fluorine is in group 7 so it has a valency of 7.
Step 3: Determine how many electrons are required by each atom to make a full valence shell. From this, determine how many need to be shared.	Carbon needs four more electrons. Fluorine needs one. If carbon shares one electron with each of the four fluorine atoms, all atoms will have complete octets. (1 mark) This results in the formation of four single covalent bonds.

Your turn

Determine how valence electrons are shared between the atoms in the covalent molecule NH_3. (1 mark)

How are covalent molecules named?

Covalent molecules are named in a similar way to ionic compounds. The first element in the molecular formula is named first using the full elemental name, followed by the second element. Prefixes are used (Table 5) to represent the number of each atom present. The exception to this is when there is only one of the first element present. For example, NO is called nitrogen monoxide as opposed to mononitrogen monoxide.

TABLE 5 Naming covalent compounds

Number of atoms	Prefix used in name	Number of atoms	Prefix used in name
1	Mono	6	Hexa
2	Di	7	Hepta
3	Tri	8	Octa
4	Tetra	9	Nona
5	Penta	10	Deca

Table 6 shows how these rules are applied to name various compounds consisting of nitrogen and oxygen atoms.

TABLE 6 Naming covalent compounds

Molecular formula	Name
NO	Nitrogen monoxide
NO_2	Nitrogen dioxide
N_2O	Dinitrogen monoxide
N_2O_3	Dinitrogen trioxide
N_2O_5	Dinitrogen pentoxide

If the second element of the compound starts with a vowel, the final "o" or "a" of the prefix is dropped to avoid confusion. So rather than saying nitrogen monooxide, the compound is nitrogen monoxide. Similarly, dinitrogen pentoxide is correct and dinitrogen pentaoxide is incorrect.

Naming covalent molecules composed of more than two types of atoms can be a little bit more complicated. You will learn about how to name more complex molecules in Units 3 & 4.

Study tip

An ionic compound is often made up of a metal cation and one or more non-metal anions. Covalent compounds often have two or more non-metals. Make sure you check whether a metal is present before naming the compound.

Challenge

Naming covalent compounds

The naming of covalent compounds is slightly different from naming ionic compounds because the ratio of each element within the compound cannot be determined from the ionic charges. For this reason, the ratio forms part of the name. For example, N_2O is dinitrogen monoxide. Note that if the first element is singular, then the quantity is not referred to in the compound's name. For example, CO_2 is carbon dioxide (*not* monocarbon dioxide). **Use** Table 5 to **identify** the names for the covalent compounds CCl_4, N_2Cl_4, CBr_4, SF_6 and P_2O_4. (5 marks)

Check your learning 3.1

Check your learning 3.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the difference between "inner electrons" and "valence electrons". (2 marks)

2 **Recall** which category of elements can form more than one ion. (1 mark)

3 **Explain** how chemical bonding is associated with the electron configuration stability. (2 marks)

4 **Describe** the role of valency in the formation of bonds. (2 marks)

5 **Explain** how ionic bonds are composed of oppositely charged ions but produce a neutrally charged compound, using sodium chloride as an example. (2 marks)

Analytical processes

6 **Compare** ionic and covalent bonding. (2 marks)

7 For the following atoms, **determine** whether electrons must be gained or lost to complete the octet rule and how many electrons must be gained or lost.
   - **a** Chlorine (2 marks)
   - **b** Potassium (2 marks)
   - **c** Nitrogen (2 marks)
   - **d** Bromine (2 marks)
   - **e** Magnesium (2 marks)

8 **Determine** the names for the following covalent compounds
   - **a** $BrCl_3$ (1 mark)
   - **b** NH_3 (1 mark)

9 **Determine** the ionic formulas for
   - **a** lithium carbonate (1 mark)
   - **b** magnesium oxide (1 mark)
   - **c** calcium nitrate (1 mark)
   - **d** potassium chloride. (1 mark)

10 **Use** the periodic table to **deduce** the missing information in Table 7. (20 marks)

11 **Use** the ionic formula $Cr_2O_7^{2-}$ as an example to **infer** what the subscripts and superscripts represent. (2 marks)

TABLE 7 Elements and their ions

Element	Group	Total number of electrons	Number of electrons lost or gained	Ionic charge formed	Anion or cation
Potassium					
		13			
Nitrogen					
		20			

Lesson 3.2

Lewis structures

Key ideas

→ Lewis structures use element symbols and electron dots to represent atoms and compounds.

→ Electron dots are used to show singular bonding electrons and electrons in a lone pair (non-bonding electrons).

→ Lewis structures can also be used to represent ions by taking into account the ionic charge.

Learning intentions and success criteria

What are Lewis structures?

Lewis structures (or **electron dot** structures) are visual diagrams that represent the valence electrons (from the highest s and p sublevels) present in atoms, ions and molecules. Electrons are drawn as dots surrounding the atom's elemental symbol. The elemental symbol represents the inner electrons that do not need to be individually shown because they do not take part in bonding.

Electron dots are added one at a time until the four sides of the element symbol are filled. The second dot for each side is then added, until all the atom's valence electrons are included. The exact location of the dots surrounding the element symbol is irrelevant, except for helium. Helium does not follow the octet rule because it has a full valence 1s sublevel of two electrons, so the two electrons are drawn together on one side.

FIGURE 1 The Lewis structure of a krypton atom

TABLE 1 Lewis structures of helium and period 2 elements of the periodic table

Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
							He
Li	Be	B	C	N	O	F	Ne

How are Lewis structures drawn for covalent molecules?

The electrons in Lewis structures of covalent molecules are categorised as either singular **bonding electrons** or **lone pair** (non-bonding) **electrons**. To create a Lewis structure of a covalent molecule, there are several steps involved.

1 Place the atoms' element symbols together, starting with the atom that has ability to form the most bondings in the centre. Surround the central atom with the other atoms.

2 Draw a single line to connect the central atom with each of the surrounding atoms to represent bonds.

3 Determine the valencies of each atom in the molecule and find the sum.

4 Draw the lone pairs of electrons on the surrounding atoms to complete the octet. The exception is hydrogen, which only requires two electrons to achieve stability.

electron dot represent the number of valence electrons in a Lewis structure

Lewis structure visual diagrams to show the number of electrons in the valence shell

bonding electrons valence electrons that are shared with another atom

lone pair electrons valence electrons that are not bonded to another atom; also known as non-bonding electrons

Study tip

Hydrogen's single electron is located in the 1s shell. It does not follow the octet rule and only requires two electrons to complete its valence shell.

5 Draw any remaining electrons on the central atom. If there are not enough electrons to finish the central atom's octet, then multiple bonds need to be formed: double or triple bonds.

6 Double check your Lewis structure by confirming the total number of valence electrons from Step **3**.

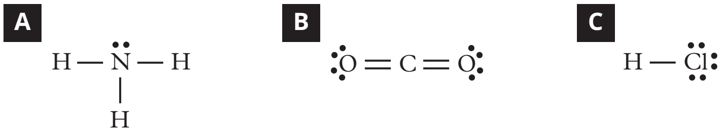


FIGURE 2 The Lewis structures of (**A**) ammonia (NH_3), (**B**) carbon dioxide (CO_2) and (**C**) hydrogen chloride (HCl). The bonding electrons are shown as lines.

Study tip

Electrons can be represented by symbols other than dots if their origin is confusing. For example, the Lewis structure of carbon monoxide could have dots for the electrons from the carbon atom and crosses for the electrons from oxygen.

Study tip

Place the bonding electrons of the surrounding atoms on the side of the element symbol nearest to the central atom. This will make the Lewis structure easier to draw.

Worked example 3.2A shows you how to do this for NF_3.

It is important to understand the limitations of Lewis diagrams. These diagrams are effective in determining the number of bonds formed, as well as whether they are single, double or triple bonds. However, they are not able to reliably determine the shape of molecules, as they are a two-dimensional diagram, and bonds can be arranged in three-dimensional space.

Coordinate covalent bonds

A **coordinate covalent bond** is formed when the two electrons in the covalent bond have been sourced from only one of the atoms involved. In other words, one atom donates a greater number of electrons than another, making the contribution of electrons uneven. For example, carbon monoxide (CO) has a triple bond between carbon and oxygen with six shared electrons – four from oxygen and two from carbon. Oxygen donates one of its lone pairs (which would otherwise be non-bonding) to the triple bond.

coordinate covalent bond
a covalent bond in which two electrons come from only one atom

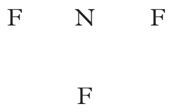
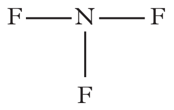
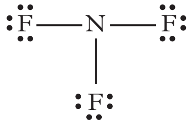
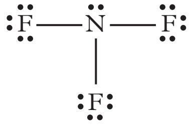
Worked example 3.2A

Determining Lewis diagrams for covalent compounds

Worked example 3.2A: Watch a video that shows how to solve this problem.

Determine the Lewis structures for NF_3. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to look at the valency of each atom in the molecule and use this information to construct a Lewis structure. The question is worth 1 mark, so we must analyse the molecule and represent it correctly.

Think	Do
Step 2: Place the atoms' element symbols together, starting with the atom that has the most bonding electrons in the centre. Surround the central atom with the other atoms.	
Step 3: Draw a single line to connect the central atom with each of the surrounding atoms to represent bonds.	
Step 4: Determine the valencies of each atom in the molecule and find the sum.	Each fluorine has seven valence electrons. Nitrogen has five valence electrons. $(3 \times 7) + 5 = 26$ valence electrons total
Step 5: Draw the lone pairs of electrons on the surrounding atoms to complete the octet.	
Step 6: Draw any remaining electrons on the central atom. If there are not enough electrons to finish the central atom's octet, then multiple bonds need to be formed: double or triple bonds.	 (1 mark)
Step 7: Double check your Lewis structure by confirming the total number of valence electrons from step 4.	The sum of all the electrons is 26, which corresponds to the total number of valence electrons.

Your turn

Determine the Lewis structures for:

a CH_3Cl (1 mark)

b HCN. (1 mark)

How are Lewis structures drawn for ions?

When drawing Lewis structures, the difference between molecules and ions is the third step. The number of valence electrons of an ion is different from the valency of its neutral atom counterpart. We need to take into account the ionic charge, which will require the addition or subtraction of electrons from the total number of valence electrons.

For example, an ionic charge of 3− means that we need to add 3 to the valency of the uncharged atom. In the same way, an ionic charge of 1+ means subtracting 1 electron from the atom's valency.

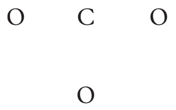
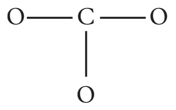
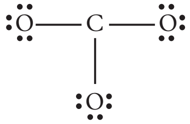
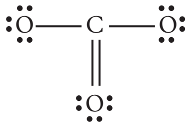
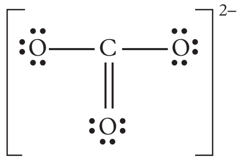
To show that the Lewis structure is for an ion, we use square brackets around the whole Lewis structure and indicate the ionic charge on the top right, like a superscript. See how this is done for the carbonate ion in Worked example 3.2B.

Worked example 3.2B

Determining Lewis diagrams for ions

Worked example 3.2B: Watch a video that shows how to solve this problem.

Determine the Lewis structure for CO_3^{2-}. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation of calculation. We need to look at the valency of each atom in the ion and use this information to construct a Lewis structure. The question is worth 2 marks, so we must analyse the ion and represent it correctly.
Step 2: Place the atoms' element symbols together, starting with the atom that has the most bonding electrons in the centre. Surround the central atom with the other atoms.	
Step 3: Draw a single line to connect the central atom with each of the surrounding atoms to represent bonds.	
Step 4: Determine the number of valence electrons by adding or subtracting electrons from the valency based on the ionic charge.	Each oxygen has six valence electrons. Carbon has four valence electrons. The 2– charge means that there are two additional electrons. $(3 \times 6) + 4 + 2 = 24$ valence electrons total
Step 5: Draw the lone pairs of electrons on the surrounding atoms to complete the octet.	
Step 6: Draw any remaining electrons on the central atom. If there are not enough electrons to finish the central atom's octet, then multiple bonds need to be formed: double or triple bonds. In this case, we have already used up all 24 valence electrons, so to complete the octet for carbon, we need to make a double bond.	
Step 7: Double check your Lewis structure by confirming the total number of valence electrons from Step 4.	The sum of all the electrons is 24, which corresponds to the total number of valence electrons.
Step 8: Draw square brackets around the Lewis structure and indicate the charge using a superscript.	 (1 mark for correct electron placement; 1 mark for charge)

Your turn

Determine the Lewis structure for PO_4^{3-}. (2 marks)

What are the exceptions to the octet rule?

There are three main exceptions to the octet rule.

1. Molecules with an odd number of electrons that cannot be completely paired – these molecules are highly reactive, unstable radicals that react to fix the odd number of electrons. An example is NO (Figure 3).
2. Molecules with a central atom that has less than an octet that cannot form multiple bonds. For example, the fluorine atoms in BF_3 are so electronegative that they do not form a coordinate covalent bond (Figure 4).
3. Molecules with a central atom with more than an octet because the central atom (often from the third period and beyond) can accept additional electrons and it is bonded to highly electronegative surrounding atoms. An example is PCl_5 (Figure 5).

FIGURE 3 Lewis structure for NO

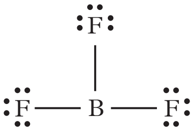


FIGURE 4 Lewis structure for BF_3

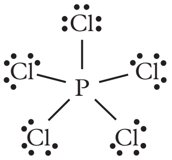


FIGURE 5 Lewis structure for PCl_5

Check your learning 3.2

Check your learning 3.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1. **Describe** the components of a Lewis structure and what they represent. (2 marks)
2. **Explain** why hydrogen is an exception to the octet rule. (1 mark)

Analytical processes

3. **Determine** the Lewis structures of the following covalent compounds:
   - **a** CCl_4 (1 mark)
   - **b** H_2O (1 mark)
   - **c** CO_2 (1 mark)
   - **d** NH_3 (1 mark)
   - **e** HCl (1 mark)
   - **f** C_2H_2. (1 mark)
4. **Determine** the Lewis structures of the following ions:
   - **a** H_3O^+ (2 marks)
   - **b** NH_4^+ (2 marks)
   - **c** OH^- (2 marks)
   - **d** NO_3^-. (2 marks)

Knowledge utilisation

5. Methane is a covalent molecule with the formula CH_4. Due to repulsion between the negatively charged electrons, the angles between the bonds are greater than 90°. **Evaluate** the effectiveness of the two diagrams in Figure 6 for their ability to accurately represent the three-dimensional model of methane. (2 marks)

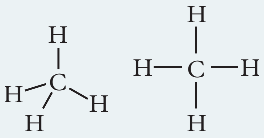


FIGURE 6 Two representations of methane

MODULE 3

Lesson 3.3

Review: Introduction to bonding

Summary

3.1

- Electrostatic attractions cause atoms to gain, lose or share electrons with each other to form compounds.
- When a metal atom reacts with a non-metal atom, a transfer of electrons can take place to form ions and then an ionic compound.
- Covalent bonding is the sharing of electrons between atoms.
- Polyatomic ions are covalently bonded molecular species that are electrically charged.
- The formulas of ionic compounds represent the ratio of atoms, and are named using the original metallic element name and a modified version of the non-metallic element name that ends with "–ide".
- When non-metals are attracted to each other, they share electrons to form covalent bonds.
- Non-metals can form single, double or triple covalent bonds with each other in order to complete their octets.

3.2

- Lewis structures use element symbols and electron dots to represent atoms and compounds.
- Electron dots are used to show singular bonding electrons and electrons in a lone pair (non-bonding electrons).
- Lewis structures can also be used to represent ions by taking into account the ionic charge.

Review questions 3.3A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Ionic bonding is characterised as

A sharing of electrons between a metal and non-metal.
B transfer of electrons between a metal and non-metal.
C transfer of electrons between two non-metals.
D transfer of electrons between two metals.

2 An ion is

A only an atom that is positively charged.
B an atom or group of atoms that are electrically charged.
C a group of atoms.
D an atom or group of atoms that are neutral.

3 How many covalent bonds will an element with six valence electrons, like oxygen, form?

A One
B Two
C Three
D Four

4 Nitrate is characterised as a polyatomic ion that contains a 1– charge and

A one nitrogen and one oxygen.
B one nitrogen and two oxygen atoms.
C one nitrogen and three oxygen atoms.
D one nitrogen and four oxygen atoms.

5 BCl_3 is a compound called

A boron chloride.
B boron chlorine.
C boron trichloride.
D monoboron trichloride.

6 Transition metals

A always form one ion.
B can form more than one ion.
C always form two different ions.
D do not form ions.

7 Valency can be used to determine

A whether an ionic or covalent bond forms.

B the number of bonds an atom can form.

C whether an atom will be ionic or neutral.

D the strength of electrostatic attractions between atoms.

8 The correct ionic formula for calcium sulfate is

A $Ca_2(SO_4)_2$.

B Ca_2SO_4.

C $CaSO_4$.

D $Ca^{2+}SO_4^{2-}$.

Review questions 3.3B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

9 **Explain** why an ion has an ionic charge. (1 mark)

10 **Describe** covalent bonding. (1 mark)

11 **Identify** what electron shell sublevels are represented in a Lewis structure. (1 mark)

12 **Explain** the bonding involved in barium hydroxide, with reference to how electrons are shared or transferred. (2 marks)

13 **Explain** the bonding involved in sodium thiosulfate ($Na_2S_2O_3$), with reference to how electrons are shared or transferred. (2 marks)

Analytical processes

14 **Determine** the Lewis structures of the following ions:

a bromide (2 marks)

b nitride (2 marks)

c carbonate. (2 marks)

15 **Determine** the Lewis structures of the following covalent compounds:

a SCl_2 (1 mark)

b CCl_2F_2 (1 mark)

c $COBr_2$. (1 mark)

16 **Determine** the Lewis structures of the following polyatomic ions:

a BF_4^- (2 marks)

b H_3O^+ (2 marks)

c ClO_2^-. (2 marks)

17 **Determine** the Lewis structures of:

a silicon (2 marks)

b caesium (2 marks)

c selenium. (2 marks)

18 **Determine** the Lewis structures of the following monatomic ions:

a nitride (2 marks)

b sulfide (2 marks)

c iodide. (2 marks)

19 **Determine** the Lewis structures of the following molecules:

a SO_3 (1 mark) **b** PH_3. (1 mark)

20 **Determine** the Lewis structures of the following polyatomic ions:

a NH_4^+ (2 marks) **b** SO_4^{2-}. (2 marks)

21 **Compare** covalent and ionic bonding. (2 marks)

22 **Compare** monatomic and polyatomic ions. (2 marks)

23 **Determine** the ionic formulas for:

a potassium sulfide (1 mark)

b radium phosphide (1 mark)

c aluminium nitride (1 mark)

d strontium chlorite (1 mark)

e boron phosphate (1 mark)

f ammonium sulfite. (1 mark)

24 **Use** the periodic table to **deduce** the missing information in Table 1. (12 marks)

TABLE 1 Elements and their resulting compounds

Group	Element	Total number of electrons	Number of valence electrons
	Chlorine		
		16	
	Silicon		
		9	

25 **Determine** the ionic formulas of:

a boron iodide (1 mark)

b magnesium oxide. (1 mark)

26 **Determine** the names for the following:

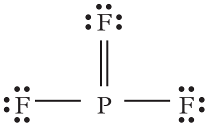
a N_2O (1 mark)

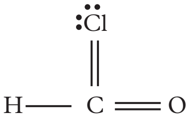
b BCl_3 (1 mark)

c $(NH_4)_2CrO_4$. (1 mark)

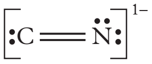
Knowledge utilisation

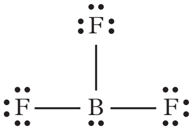
27 **Evaluate** the following Lewis structures to **identify** the errors. Provide a correct representation of each.

a

(2 marks)

b

(2 marks)

28 Evaluate the Lewis structures provided to **identify** the errors. Provide a correct representation of each.

a

(2 marks)

b

(2 marks)

29 A difference in electronegativity of 0–0.44 is considered a non-polar covalent bond because the electrons are equally shared by the atoms – the electrons are placed in the middle of the bond.

A difference in electronegativity of 0.4–1.8 is considered a polar covalent bond because the electrons are not shared equally by the atoms – the electrons are closer to one atom. A difference in electronegativity of >1.8 is considered an ionic bond because the electrons have actually transferred completely to one of the atoms.

a **Use** the electronegativity values in Table 2 to **determine** the location of electrons in Lewis structures of Cl_2, H_2O and SiH_4. Pay particular attention to the placement of electron dots on the bonds. (3 marks)

TABLE 2 Electronegativity of elements

Element	Cl	O	H	Si
Electronegativity	3.2	3.4	2.2	1.9

b **Evaluate** the significance of not assigning a difference in electronegativity of 1.6–1.9 as either ionic or covalent. **Use** HF and NaBr as examples. (2 marks)

c **Investigate** the effect of polar covalent versus ionic bonds on a compound's properties. (2 marks)

Data drill

Properties of chlorine compounds

The melting point, boiling point and electrical conductivity of some chlorine compounds were experimentally determined and recorded in Table 1.

TABLE 1 Properties of various chlorine compounds

Substance	Melting point (°C)	Difference in electronegativity	Electrical conductivity (solid)	Electrical conductivity (liquid form)
Lithium chloride	605	2.2	poor	good
Beryllium chloride	399	1.6	poor	poor
Boron trichloride	question **5**	1.2	question **5**	question **5**
Carbon tetrachloride	–22	0.6	poor	poor
Sodium chloride	801	2.3	poor	good
Potassium chloride	770	2.4	poor	good

Apply understanding

1 **Identify** ionic formula for boron trichloride. (1 mark)

2 **Determine** the range of the melting points in Table 1. (1 mark)

Analyse data

3 **Categorise** the compounds as ionic or covalent. (2 marks)

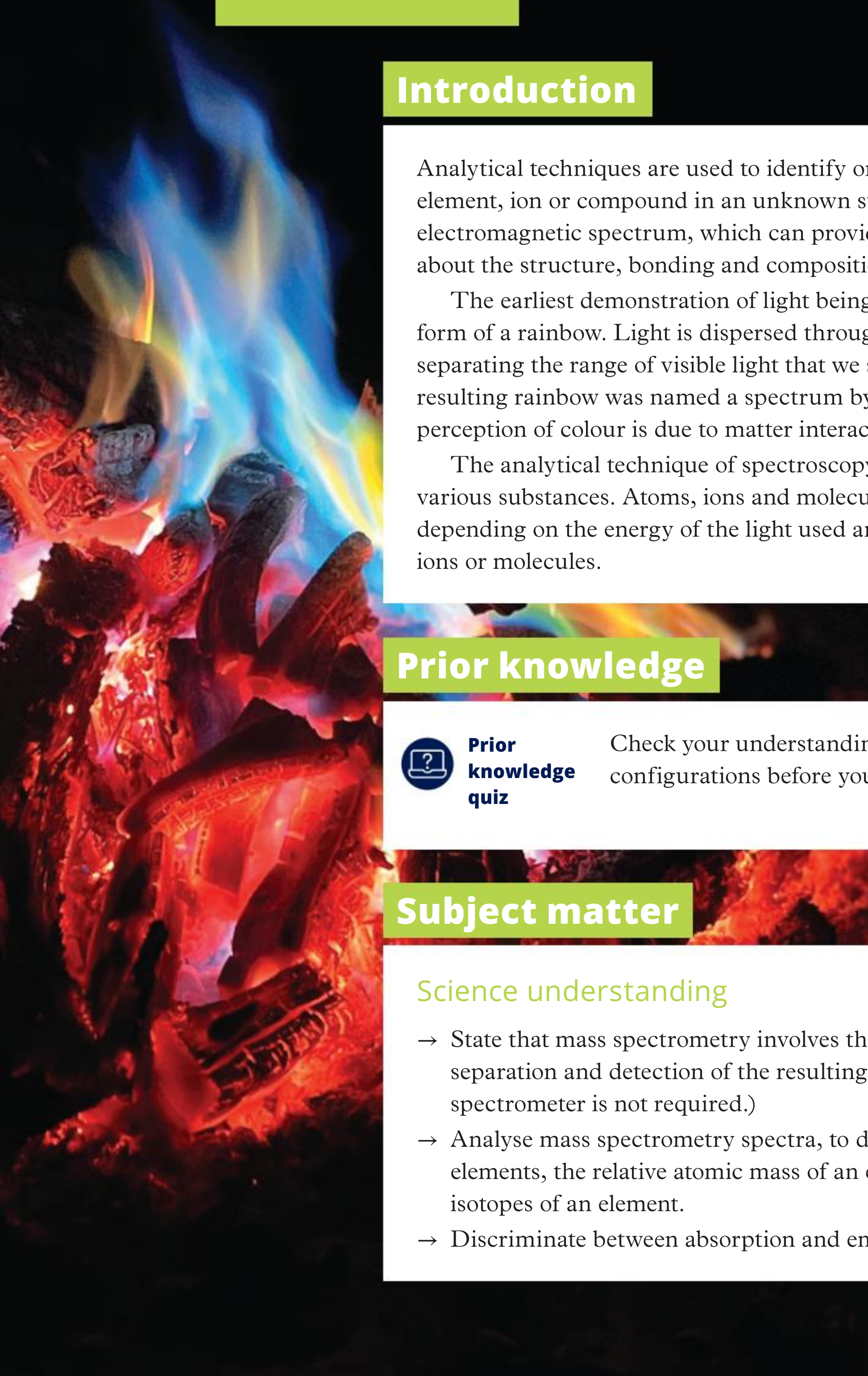
4 **Contrast** the properties of one ionic compound and one covalent compound, using examples. (1 mark)

Interpret evidence

5 **Predict** the melting point and electrical conductivity of boron trichloride. **Justify** your predictions. (3 marks)

Module 3 checklist: Introduction to bonding

Quizlet: Revise key terms online to test your understanding



MODULE 4

Analytical techniques

Introduction

Analytical techniques are used to identify or measure the amount of a molecule, element, ion or compound in an unknown substance. Many techniques use the electromagnetic spectrum, which can provide some of the most valuable information about the structure, bonding and composition of molecules.

The earliest demonstration of light being dispersed through a prism was in the form of a rainbow. Light is dispersed through water droplets, which act as a prism, separating the range of visible light that we see as white light into colours. The resulting rainbow was named a spectrum by Sir Isaac Newton in the 1600s. The perception of colour is due to matter interacting with visible light of different energies.

The analytical technique of spectroscopy uses light to determine information about various substances. Atoms, ions and molecules interact with electromagnetic radiation depending on the energy of the light used and the energy states within those atoms, ions or molecules.

Prior knowledge

Prior knowledge quiz

Check your understanding of ionic charge, isotopes and electron configurations before you start.

Subject matter

Science understanding

→ State that mass spectrometry involves the ionisation of substances and the separation and detection of the resulting ions. (The operation of the mass spectrometer is not required.)

→ Analyse mass spectrometry spectra, to determine the isotopic composition of elements, the relative atomic mass of an element and percentage abundances of the isotopes of an element.

→ Discriminate between absorption and emission line spectra.

→ Explain that flame tests and atomic absorption spectroscopy (AAS) rely on electron transfer between atomic energy levels.

→ Explain that the emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels (Bohr model), which converge at higher energies.

→ Analyse flame tests and atomic absorption spectroscopy (AAS) to identify elements and determine the concentration of metallic ions in solution.

Science as a human endeavour

→ Appreciate that analysis of the distribution of elements in living things, Earth and the universe has informed a wide range of scientific understandings.

Science inquiry

→ Investigate flame tests to identify elements.

→ Investigate mass spectra and isotopes*

→ Investigate atomic absorption spectroscopy (AAS) and the concentration of aqueous metallic ions.*

***Note:** Simulations may be used.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

These lessons are available on Oxford Digital.

Lesson 4.3 Conducting a flame test
Lesson 4.4 Simulating atomic absorption spectroscopy
Lesson 4.6 Simulating mass spectrometry

Lesson 4.1

Principles of spectroscopy

Key ideas

- → Spectroscopy is a branch of physical chemistry that studies the interactions between chemicals and the electromagnetic spectrum.
- → When electrons absorb electromagnetic radiation, an absorption spectrum is produced. When they emit electromagnetic radiation, an emission spectrum is produced.
- → Absorption and emission spectra support Bohr's model for electron configuration.

What is spectroscopy?

Learning intentions and success criteria

Spectroscopy is the branch of chemistry that investigates the interaction between matter and the electromagnetic spectrum. The way in which molecules interact with light (electromagnetic radiation) will depend not only on the type of molecule, atom, element or compound, but also on the type of electromagnetic radiation involved.

The electromagnetic spectrum

The **electromagnetic spectrum** consists of the full range of **frequencies** of light. It includes gamma rays, X-rays, ultraviolet and visible light, through to infrared, microwaves and radio waves. These sections of the electromagnetic spectrum are defined by their different energies, frequencies and **wavelengths**.

Electromagnetic radiation, or light, consists of **photons** (particles lacking mass), which travel at the speed of light in a vacuum, such as space. A photon behaves as both a particle and a wave, depending on the type of the experiment in which it is used. In chemistry, we typically consider light as a particle. These particles of light carry energy and are absorbed or emitted from atoms. The energy of a photon is related to its frequency. The higher the frequency, the higher the energy of the light and the smaller the wavelength.

Figure 1 shows the electromagnetic spectrum. There is an inverse relationship between energy and wavelength: as energy increases, wavelength decreases. Visible light forms the narrowest region of the spectrum.

electromagnetic spectrum
the range of frequencies of electromagnetic radiation and their energies and wavelengths; electromagnetic radiation, or light, is comprised of photons

frequency
the number of waves (referring to wavelength) passing a point each second

wavelength
the distance between two crests or troughs of a wave

photon
a particle and wave of light that carries energy; is absorbed or emitted from atoms

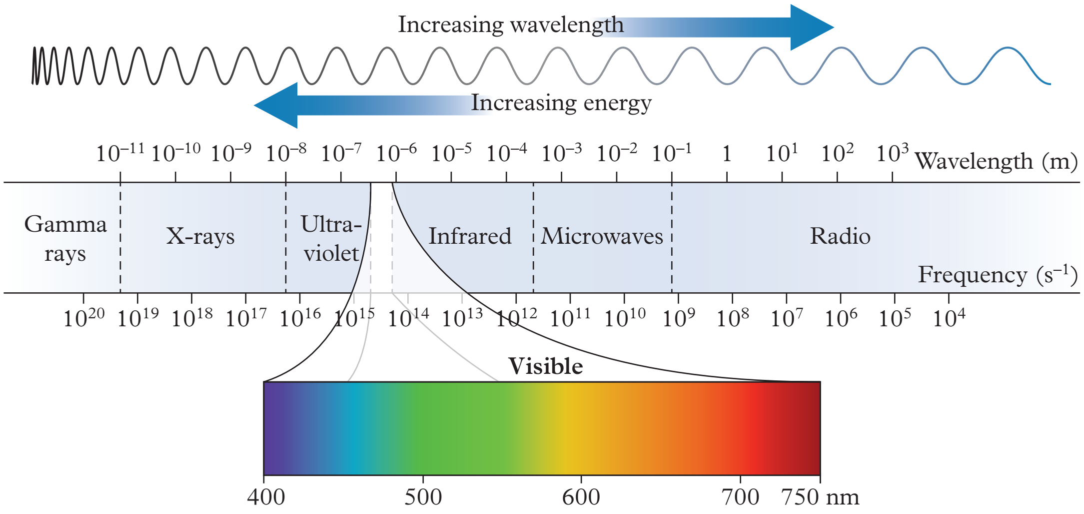


FIGURE 1 The electromagnetic spectrum consists of the full range of frequencies of light.

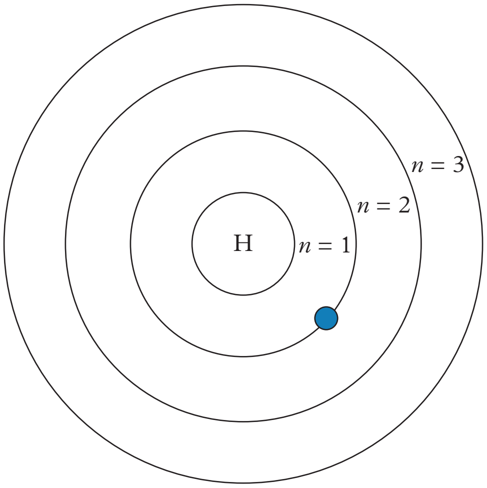


FIGURE 2 Energy level diagram of the ground state of a hydrogen atom, with its electron found in the first electron shell or orbital $n = 1$

The different frequencies of electromagnetic radiation have many everyday applications. For example, radios and televisions use radio waves, microwave ovens use electromagnetic radiation from the microwave region, and X-rays are used to visualise broken bones.

How do electrons move between orbitals of different energy levels?

In an atom, negatively charged electrons are found in orbitals outside the positively charged nucleus. Figure 2 shows an energy level diagram for a hydrogen atom, which contains one positively charged proton in its nucleus, and one negatively charged electron. The atom is in its **ground state**, or most stable state, and the electron is found in the first electron shell ($n = 1$).

When an atom **absorbs** energy, electrons move from an orbital with a lower energy level to an orbital with a higher energy level; for example, from $n = 1$ to $n = 2$ or $n = 3$. This is called the **excited state** because electrons contain more energy than they would if they were in their ground state. The more energy that is absorbed, the further away from the nucleus the electrons will move as they transition to a higher energy level.

Study tip

The plural of spectrum is "spectra".

ground state
an energy state in which electrons are in their lowest possible energy configuration within the orbitals of the atom

absorbs
takes in; in this case, the electron "takes in" energy from a photon

excited state
an energy state of an atom that is higher than the ground state; an atom is unstable in this state and emits energy as a photon of electromagnetic radiation as the atoms return to a lower energy state or electron configuration

For an electron to move from a lower to a higher energy level, it must absorb energy in the form of electromagnetic radiation. The amount of electromagnetic radiation equals the difference between two energy levels. For example, for an electron to transfer from the third ($n = 3$) to the fifth ($n = 5$) energy level, the energy absorbed must be equal to the difference in energy of these levels.

Figure 3 demonstrates the different energy levels of orbitals available for an electron to move to, for the simplest element, hydrogen.

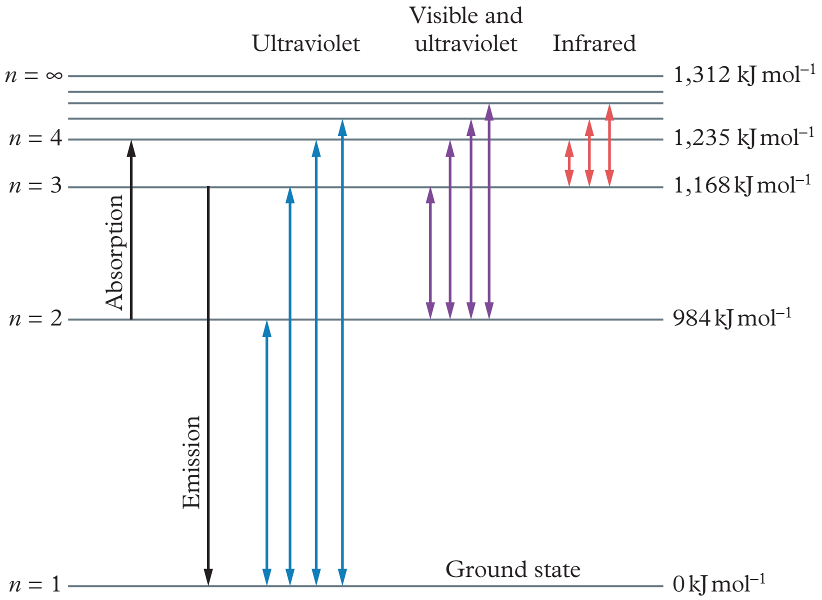


FIGURE 3 Energy levels of electron orbitals for hydrogen

Transitions from higher to lower energy levels involve an emission of energy. Different transitions give off radiation in different parts of the electromagnetic spectrum. For example, transitions from higher levels to the second level ($n = 2$) have energies corresponding to radiation in the visible part of the electromagnetic spectrum.

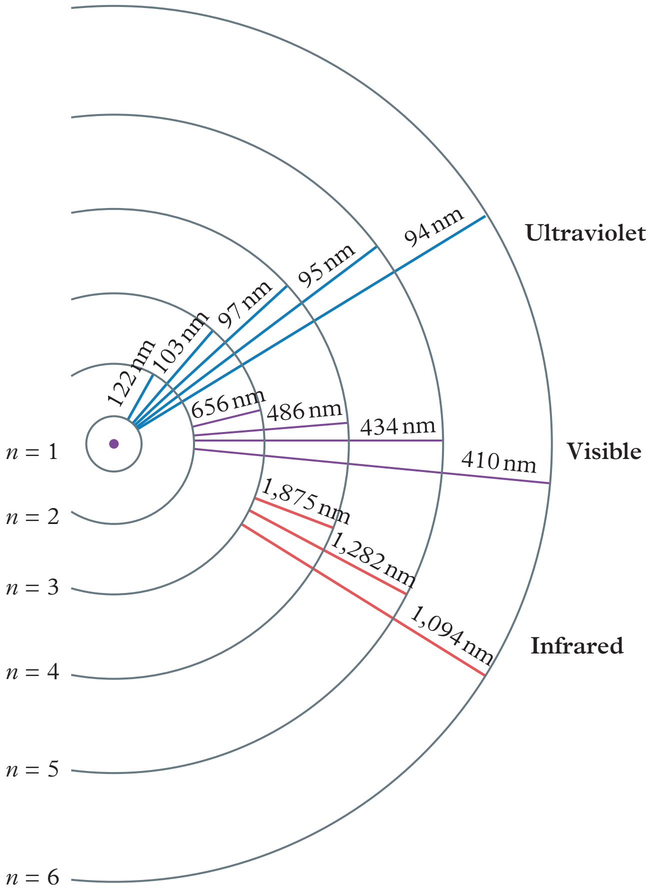


FIGURE 4 The energy emitted by electrons, as a wavelength (nanometres), as they transition from higher energy levels to lower ones. The purple dot represents the nucleus.

For an electron to move from the first ($n = 1$) to the second ($n = 2$) level, it must absorb a single photon containing 984 kJ mol^{-1} of energy with a wavelength of 122 nm. An atom in its excited state is unstable so it will tend to **emit** the excess energy as a photon of electromagnetic radiation.

What are absorption and emission spectra?

As electrons both absorb and emit electromagnetic radiation, they produce an **absorption spectrum** and an **emission spectrum**. The position of the lines on the absorption spectrum is identical to the position of the lines on the emission spectrum, as each line correspond to a transition between the same two energy levels.

A continuous spectrum shows the visible spectrum of light – all of the colours that have wavelengths between 400 and 750 nm. When white light passes through a prism, it produces a continuous spectrum – a rainbow of colours. An emission spectrum consists of a set of bright lines against a black background. An absorption spectrum consists of dark lines against a continuous spectrum.

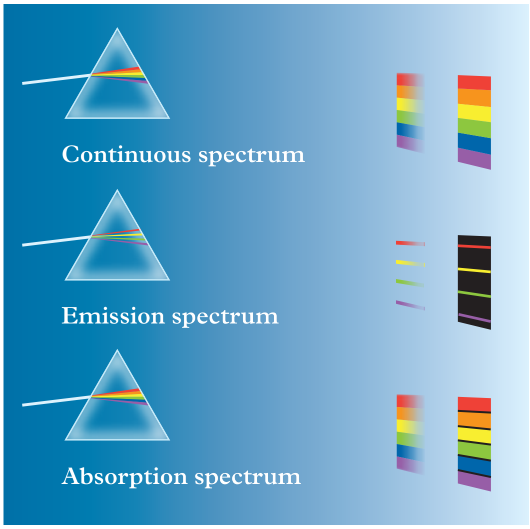


FIGURE 5 Continuous, emission and absorption spectra

Study tip

Review what you have already learnt about energy levels in atoms. There is a limit to the amount of energy that an atom can absorb before an electron is completely removed and the atom becomes an ion. This is called the ionisation energy.

emit
releases; in this case, the atom “releases” energy in the form of a photon

absorption spectrum
the band of lines observed when an atom of an element absorbs light; viewed as a series of black lines on the visible light spectrum

emission spectrum
the band of lines observed when an atom of an element emits light; viewed as a series of coloured lines that represent the wavelengths of electromagnetic radiation emitted by a source

An emission spectrum can be generated by heating a sample to excite its electrons. The light that the sample emits when the electrons move from an excited state to a lower state is focused through a slit and passed through a prism to disperse the light into a spectrum. The photons of light and their corresponding wavelengths that are emitted when the electrons transition to lower orbitals are seen as coloured lines on a black spectrum. Each element produces its own unique spectra. The emission spectrum for hydrogen is shown (Figure 6).

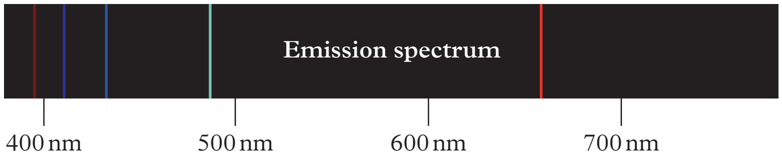


FIGURE 6 The emission spectrum in the visible region, generated by hydrogen when it emits photons of light at specific wavelengths

An absorption spectrum is generated when white light is passed through a gas sample. The sample absorbs photons of light of specific wavelengths as electrons transition to higher energy levels. The light that is not absorbed is focused through a slit and passed through a prism to disperse the light into a spectrum. On a continuous spectrum, the wavelengths that have been absorbed are missing and appear as black lines. These black lines correspond to the energy of the photons required to transition an electron from a lower to a higher energy level. The absorption spectrum of hydrogen is shown in Figure 7.

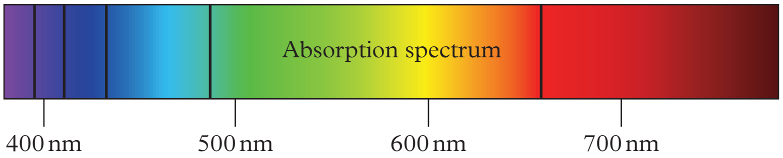


FIGURE 7 The absorption spectrum in the visible region, generated by hydrogen when it absorbs photons of light at specific wavelengths

How do absorption and emission spectra support Bohr's model?

The absorption and emission spectra of the element hydrogen provided the evidence for Niels Bohr to propose what is termed the Bohr model. His model of the atom was the first to recognise that electrons have discrete energies, and can transition between discrete energy levels. Although our understanding of the atom has increased with the quantum mechanical model, the Bohr model was a breakthrough at the time. Up until then, there was no explanation for how spectral lines could be produced.

In 1913, Bohr proposed that:

- electrons orbit an atom's nucleus
- electrons exist in orbits, which are distinct distances from the nucleus and have definite energy levels
- for an electron to move between orbits, it must gain energy or lose energy by absorbing or emitting photons of light with distinct wavelengths.
- the energy levels of the orbits are far apart at the lower level for hydrogen, and are closer together and converge for the higher energy level orbits, as shown by the increasingly close spacing of the absorption lines in the blue–violet region.

Currently, the quantum mechanical model refers to "orbitals", which are regions of three-dimensional space in which electrons are likely to be found in, rather than orbits of distinct distances.

Real-world chemistry

Robert Bunsen, a pioneer of spectroscopy

Robert Bunsen was a German chemist who studied the electromagnetic spectrum, specifically emission spectra. In 1852, he was hired by the University of Heidelberg, who enticed him into employment by building him a new laboratory with gas inlets.

Bunsen had long grappled with the issue of laboratory burners which produced soot, generated very little heat, and burned a brilliant yellow colour. The yellow colour interfered with emission spectra, and without high temperatures, the electrons of metal atoms would not be excited to higher energy levels. The soot also interfered with any data obtained.

Ideally, Bunsen wanted to make a burner to view the emission spectra of metals. He needed a very hot flame (to excite the electrons), which emitted very little to no light – so that it would not interfere with the spectra – and did not produce soot.

Bunsen approached the university's mechanic, Peter Desaga, to build a prototype that would satisfy all of these criteria. The design, developed by 1855, included an adjustable air opening, which produced a cooler yellow flame when it was closed and a hotter blue flame when it was open. The blue flame, as well as being very hot, emitted very little light.

In 1859, Bunsen collaborated with Kirchhoff to invent the first spectroscope, which was able to separate the colours using a prism. This enabled them to identify the specific spectral lines of sodium, lithium and potassium, and for Bunsen to prove that highly pure samples of each element produced unique spectra. Using the burner together with his newly invented spectroscope, Bunsen discovered the elements caesium and rubidium and received several awards for his achievements.

FIGURE 8 Bunsen burners are widely used to conduct flame tests.

The Bunsen burner (Figure 8) is still used in school and professional laboratories and is one of the most effective methods of generating an emission spectrum.

Apply your understanding

1 To create an emission spectrum, Bunsen required a source of heat. **Explain** the effect of heat energy on atoms with a ground state electron configuration, and how adding heat assisted in creating an emission spectrum. (2 marks)

2 **Explain** why different elements would have unique spectral lines. (2 marks)

3 Bunsen discovered that highly purified samples gave unique spectral lines. **Explain** how an impure sample would affect the spectrum Bunsen saw. (2 marks)

Check your learning 4.1

Check your learning 4.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** how the emission spectrum of the hydrogen atom led to:
   - **a** the realisation that electrons must have discrete energies. (2 marks)
   - **b** Bohr's model of the atom. (2 marks)

2 Figure 7 shows the spectral lines in the visible part of the hydrogen spectrum. The line in the red region of the spectrum corresponds to the transition of the electron from the second energy level ($n = 2$) to the third ($n = 3$). The higher energy lines to the left converge closer together. **Explain** what the spectrum shows about the energies of the higher-level orbitals. (2 marks)

Analytical processes

3 **Discriminate** between absorption and emission line spectra by classifying the statements below into the correct columns in Table 1. Two statements have been included for you. (7 marks)
   - appears as dark lines on the electromagnetic spectrum
   - appears as bright lines on the electromagnetic spectrum
   - provides evidence that electrons in an atom can only have discrete energies
   - produced when an electron transitions to a higher energy level
   - produced when an electron transitions to a lower energy level
   - produced by absorption of photons of particular energies
   - can have lines in the ultraviolet, visible or infrared part of the spectrum.

TABLE 1 Features of absorption and emission spectra

Emission spectra	Both emission and absorption spectra	Absorption spectra
• produced by emission of photons of particular energies	• unique for each element	

4 **Analyse** the emission spectrum of hydrogen in Figure 6.
   - **a** **Deduce** the number of energy transitions hydrogen is showing in the visible range spectrum. (1 mark)
   - **b** **Consider** whether it would be possible for hydrogen to have more transitions than those appearing on this spectrum. **Explain** your answer. (2 marks)

Lesson 4.2
Techniques for identifying elements

Learning intentions and success criteria

Key ideas

→ Flame tests are qualitative tests used to identify metals.

→ Atomic absorption spectroscopy is a quantitative technique used to analyse trace levels of metals.

What are flame tests?

Absorption and emission spectra are essential to the qualitative and quantitative analysis of metals. One analytical technique that can be used to identify elements is the **flame test**. It is a qualitative test that relies on electron transfer between atomic energy levels.

flame test
the analysis of metal ions by heating them in a Bunsen burner flame and measuring the light emitted when excited electrons return to their ground state

The flame test is used to demonstrate the characteristic colours that metals emit when placed in a blue Bunsen burner flame (Figure 1). The heat of the flame excites electrons to a higher electron orbital. Photons of light are emitted at distinct wavelengths when electrons return to their ground state.

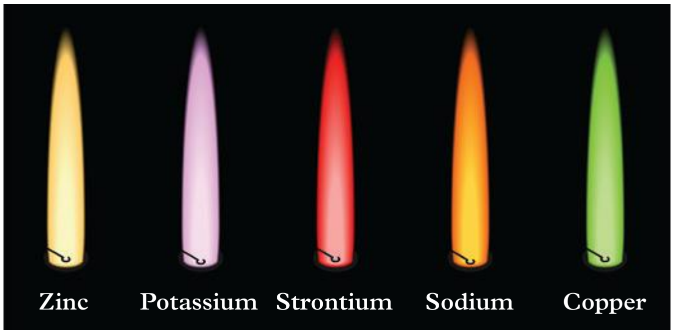


FIGURE 1 The colours produced by some metal ions in a blue Bunsen burner flame

Typically, flame tests are conducted on metal chlorides. Traditionally, a small solid sample was collected on the end of a wire loop and placed into the blue Bunsen burner flame. However, the colours produced can be quite difficult to distinguish from one another. Some metal ions look very similar and therefore the identification of various metals is impossible with the naked eye. For example, lithium, sodium, calcium and strontium appear yellow, red or orange, and this can make it difficult to tell them apart. A newer method uses a flame test bottle with a spray head. A very fine mist of a solution of the metal chloride is sprayed into the flame, which produces the characteristic colours for each metal.

To overcome the issue of similar colours, a spectrograph is used to generate an emission spectrum of each element (Figure 3). Although they may appear to be the same colour to your eyes, the photons of light emitted at various wavelengths differ significantly. Metals can be identified using the spectrum.

FIGURE 2 What metal/s could be on this wire loop?

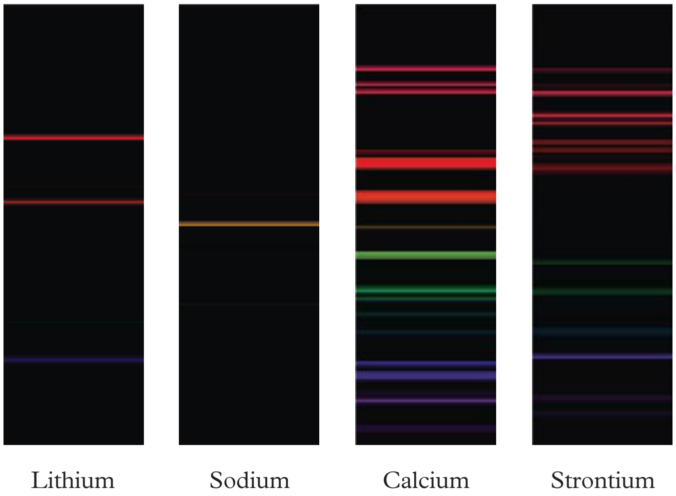


FIGURE 3 Emission spectra of lithium, sodium, calcium and strontium are used to identify the elements.

Skill drill

Assessing risks involved in flame tests

Science inquiry skill(s): Considering safety and ethics (Lesson 1.5)

When working in a chemistry laboratory, a variety of risks will be encountered. As part of your student experiment, you will be required to conduct a risk assessment.

Risks in chemistry can come from hot or cold objects, heavy objects, glassware, electricity, and from the chemicals themselves. The risks of the chemicals used in senior chemistry cannot be understated, and this includes some of the solutions used in flame tests. To assess these risks, become familiar with accessing, reading and using Safety Data Sheets (SDS).

Your school may have a suggested or preferred way for you to record risks in an experimental report.

Practise your skills

1 **Use** the SDS for each solution or solid being used in the flame test and **summarise** the hazards for each chemical. (2 marks)

2 Review the standard operating procedure (SOP) for lighting and using a Bunsen burner, including considering the risk of accidentally extinguishing the flame if using spray bottles for flame tests. **Identify** the hazards associated with using the Bunsen burner for flame tests. (2 marks)

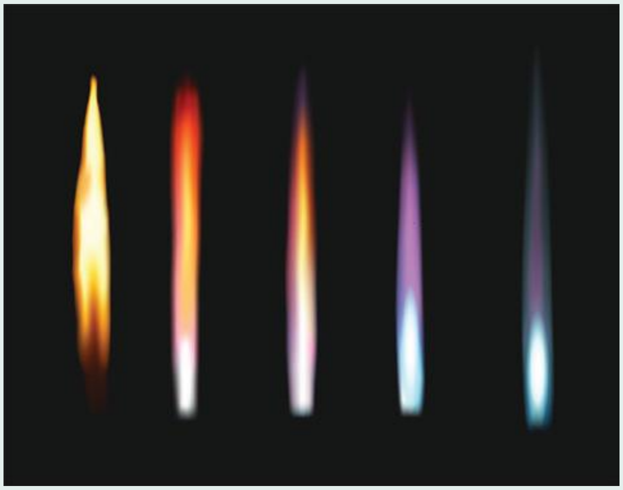


FIGURE 4 There are many risks involved in flame tests.

3 **Develop** a risk assessment in the suggested/preferred format, taking account of all of the hazards identified in questions **1** and **2**, assessing the risk, and describing the action/s required to avoid or overcome each risk. (3 marks)

What is atomic absorption spectroscopy?

atomic absorption spectroscopy (AAS)
the determination of the concentration of metal atoms in the gas phase by measuring the amount of light absorbed from a specific lamp

Atomic absorption spectroscopy (AAS) is an analytical technique in which absorption spectra of metals are used to determine the quantity of the metal in an unknown sample. It is essential in the analysis of trace levels of compounds, as low as micrograms per litre ($\mu g\,L^{-1}$).

An atomic absorption spectrometer works by spraying a solution of the sample into a flame, which converts the sample into atoms (Figure 5). A lamp that produces specific wavelengths for the element being tested creates a light beam. The light is pulsed at the sample in the flame, to excite the electrons in the atom of interest. This causes some of the light beam to be absorbed, reducing its intensity.

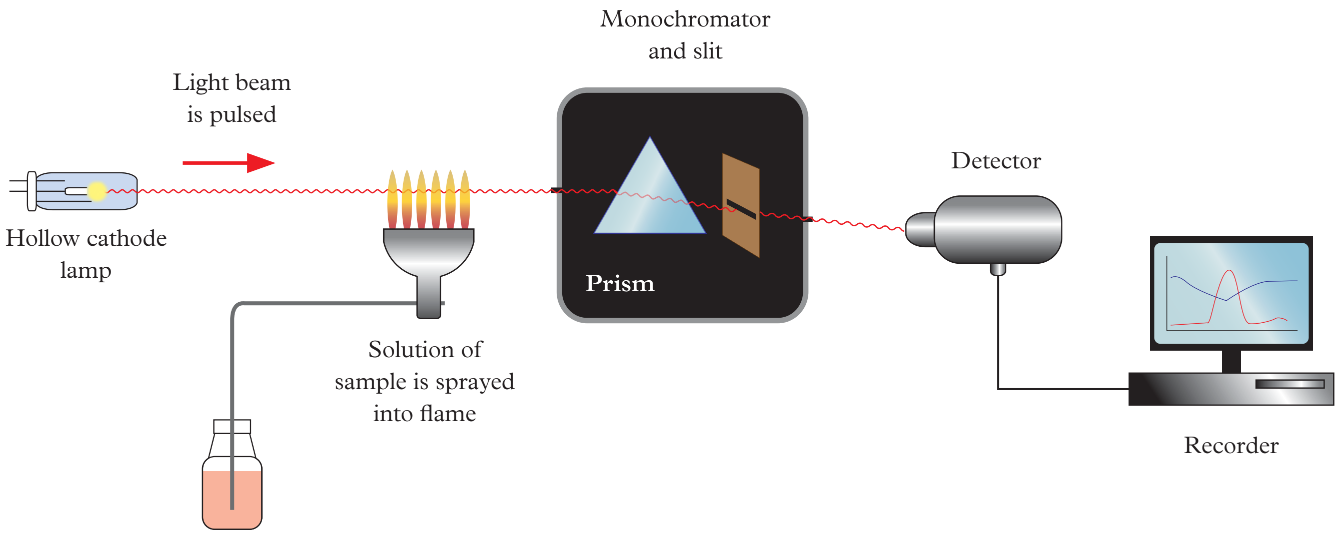


FIGURE 5 Schematic diagram of an atomic absorption spectrometer

The transmitted light is passed through a monochromator, which is a diffraction grating capable of selecting the wavelength that reaches the detector. The detector measures the absorbance caused by the decrease in intensity of the original light beam. The greater the absorbance (decrease in intensity), the higher the concentration of the metal atoms in the sample.

The absorbance corresponds to an electron transitioning to a particular energy level. Contrast this with the absorption spectrum discussed in Lesson 4.1, which displays a number of lines corresponding to transitions which absorb light of different wavelengths and energies. The AAS lamp needs to be specific to the element being tested.

Real-world chemistry

Development of the atomic absorption spectrometer

The invention of AAS is credited to British-Australian physicist Sir Alan Walsh (1916–1998). Until Walsh invented the atomic absorption spectrometer, all atomic spectra analysed were emission spectra. This method had very little sensitivity and could analyse fewer than ten elements.

Walsh's prior research experience was during WWII, in Britain, where he had analysed the metals in enemy bombers which had been shot down. He developed improvements to the equipment used, but it was still difficult to obtain quantitative results. He moved to Australia to work for the Council for Scientific and Industrial Research (CSIR), where he installed and used the first infrared spectrometer in Australia to study infrared absorption by molecules. It was not long before he improved this technique, which was patented and used worldwide.

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FIGURE 6 Sir Alan Walsh – the inventor of atomic absorption spectrometry

Walsh, in a "Eureka!" moment, concluded that scientists had been looking at atomic spectra in the wrong way and that they should be looking at absorption, rather than emission, spectra, as he had been doing for molecules. Amazingly, it took him less than a day to build a working model of an atomic absorption spectrometer. Initial testing of the instrument indicated that it could analyse 65 different elements with increased accuracy and sensitivity.

Walsh could not persuade any companies to manufacture the atomic absorption spectrometer commercially. So, he separated the instrument into multiple components and sent each component to a separate company for manufacture. The result was a kit version that buyers assembled into an operating spectrometer.

The atomic absorption spectrometer has proven essential for metal analysis. It is used in mining to determine the amount of metal in an ore sample, and in medicine and environmental analysis where it has saved lives by determining unsafe levels of metals in blood or the mercury content in fish and waterways.

Apply your understanding

1 Mining companies need to know the economic value of an ore body. **Explain** why determining the concentration of a metal in a newly discovered ore body would help them make decisions about whether it is viable to mine. (2 marks)

2 **Investigate** the meaning of hyperkalemia and hypokalemia, and **explain** why rapid analysis using AAS is important for patients suffering from these conditions. (3 marks)

3 Mercury from the environment accumulates in fish, especially those higher up in the food chain. **Explain** why it is important to be able to analyse mercury content in fish. (2 marks)

4 **Explain** why mining rehabilitation workers restoring the environment after a mine closes might use AAS. (2 marks)

standard
a sample with a known concentration of the chemical that is being tested for

calibration curve
a graph of the absorbance of a set of standards against their concentration

How do you analyse and interpret data from AAS?

AAS is a quantitative technique because it can measure the concentration of metal atoms in a sample. The sample is compared to a set of **standards** of known concentration of the metal being analysed in the sample. A **calibration curve** of the standards is prepared, and the unknown sample's concentration is calculated from this. The calibration curve is a graph that plots the concentration of the standards on the *x*-axis and the absorbance of the light on the *y*-axis.

Worked example 4.2A

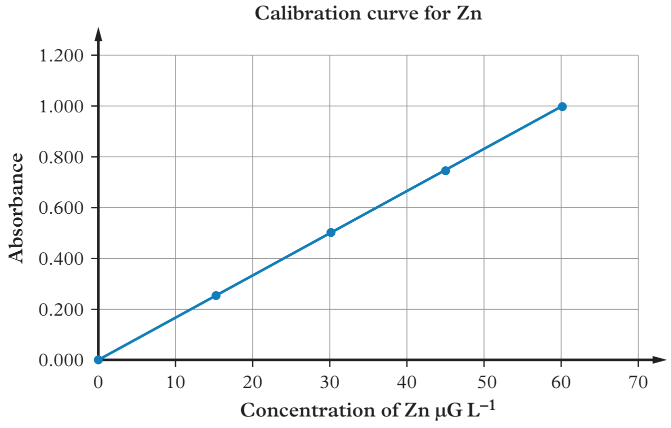
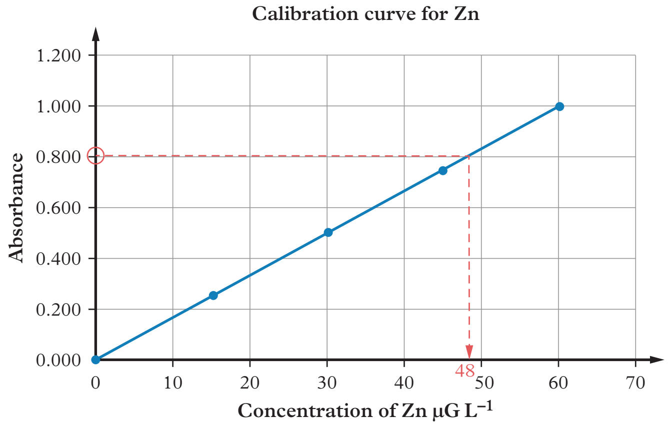
Determining the concentration of metal using AAS

Worked example 3.1A: Watch a video that shows how to solve this problem.

A sample containing zinc is analysed by AAS. A set of standards is run with the sample and the data obtained is shown in Table 1. **Determine** the concentration of zinc atoms in the sample. (2 marks)

TABLE 1 Concentration and absorbance of zinc standards using AAS

Standard/sample	Concentration (μg L^{-1})	Absorbance
Standard 1	0	0.000
Standard 2	15	0.250
Standard 3	30	0.500
Standard 4	45	0.750
Standard 5	60	1.000
Sample	–	0.800

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to use a calibration curve to find the unknown value. The question is worth 2 marks, so we must plot the graph and locate the value.
Step 2: Construct a calibration curve by plotting the concentration of the standards against absorbance. Draw a linear trendline.	 (1 mark for correct curve)
Step 3: Locate the absorbance of the sample on the *y*-axis and draw a horizontal line to the trendline. Draw a vertical line from this point down to the *x*-axis.	
Step 4: Finalise your answer.	The sample has a zinc concentration of 48 μg L^{-1}. (1 mark)

Your turn

A chemist wants to determine the Cu^{2+} content in tap water by AAS. After initial testing, she determines that the sample is too concentrated and dilutes it by a factor of 100. She then reruns the sample with a set of Cu^{2+} standards and obtains the data in Table 2. **Determine** the concentration of Cu^{2+} in the diluted sample and in the original sample of tap water. (3 marks)

TABLE 2 Concentration and absorbance of Cu^{2+} standards using AAS

Standard/sample	Concentration (μg L^{-1})	Absorbance
Standard 1	0	0.000
Standard 2	20	0.210
Standard 3	40	0.420
Standard 4	60	0.630
Standard 5	80	0.840
Diluted sample	–	0.720

Study tip

Double check the counting of the squares on the graph axis to make sure your scale is correct. Always use a ruler to be more accurate!

Challenge

Limitations of AAS

Using AAS has limitations, such as difficulty with solutions of high concentrations. **Use** Excel to plot a calibration curve for the AAS data in Table 3, with an absorbance of 0.0000 for concentration 0 mg L^{-1}. **Determine** the maximum concentration for which the data is linear, and able to provide precise measurements. (2 marks)

TABLE 3 Results from AAS experiment

Concentration (mg L^{-1})	Absorbance	Concentration (mg L^{-1})	Absorbance
1	0.0334	15	0.4956
2	0.0653	20	0.6026
4	0.1315	25	0.7062
7	0.2308	30	0.7781
10	0.3301	40	0.8395

Check your learning 4.2

Check your learning 4.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Use** your knowledge of energy levels of electrons in atoms to:

- a **explain** how a metal can be analysed in a flame test (3 marks)
- b **relate** this to either absorption or emission spectra (4 marks)
- c **describe** what would be expected if the light were passed through a prism or diffraction grating. (1 mark)

2 **Explain** how AAS relates to absorption spectra, with reference to electron energy levels. (2 marks)

Analytical processes

3 An analysis was conducted by AAS to determine the concentration of cadmium in sea turtle blood. **Determine** the concentration of cadmium in the solution “Sample 1” from the data in Table 4. (2 marks)

TABLE 4 Cadmium concentrations of standards

Standard/sample	Concentration ($\mu\text{g L}^{-1}$)	Absorbance
Standard 1	0	0.000
Standard 2	14	0.042
Standard 3	28	0.084
Standard 4	42	0.126
Standard 5	66	0.168
Sample 1	–	0.096

4 Two samples of waste from a manufacturing company are analysed to determine whether the concentration of lead(II) ions, Pb^{2+}, is at a safe environmental level. The recommended maximum concentration of lead(II) ions is 5 $\mu\text{g L}^{-1}$. **Analyse** the results in Table 5 and provide a recommendation to the company based on your findings. (5 marks)

TABLE 5 Concentration and absorbance of lead standards using AAS

Standard/sample	Concentration ($\mu\text{g L}^{-1}$)	Absorbance
Standard 1	0	0.000
Standard 2	2	0.130
Standard 3	4	0.260
Standard 4	6	0.390
Standard 5	8	0.520
Sample 1	–	0.200
Sample 2	–	0.350

Practical

Lesson 4.3

Conducting a flame test

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Video demonstration

Practical

Lesson 4.4

Simulating atomic absorption spectroscopy

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Lesson 4.5

Mass spectrometry

Learning intentions and success criteria

Key ideas

→ Mass spectrometry involves measuring the mass-to-charge ratio of ion fragments.

→ The mass and percentage abundance of isotopes, and relative atomic mass can be determined from mass spectra.

What is mass spectrometry?

Mass spectrometry is one of the most useful analytical techniques available to chemists. It is used to determine the isotopic composition of elements and the bonding structure of molecules.

Every element has isotopes that must be considered when determining its mass. Isotopes have different masses because they have different numbers of neutrons. Therefore, the mass of an element must be calculated as an average of the masses of an element’s isotopes. To determine how much of each isotope exists, mass spectrometry is used.

In mass spectrometry, a sample of an element or compound is injected into the ionisation chamber of a **mass spectrometer**.

1 In the chamber, one of several different methods is used to turn the sample into positively charged ions. This process is called **ionisation**.

2 The ions then pass through an electric field, which accelerates them through a magnetic field.

mass spectrometry an analytical technique in which chemical elements are ionised and the ions are sorted according to their mass-to-charge ratio

mass spectrometer an instrument that ionises elements and molecules and produces spectra showing mass-to-charge ratio

ionisation the formation of ions from atoms by the addition or removal of electrons

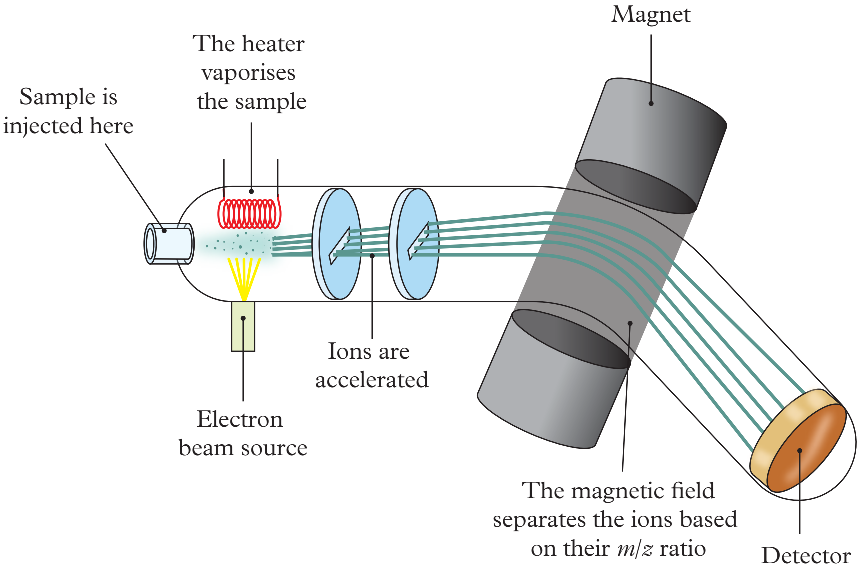


FIGURE 1 The path of particles through a mass spectrometer

3 Within the magnetic field, the ions follow a curved path, with the paths of more positively charged, lighter ions bending more than the paths of less positively charged, heavier ions (Figure 1).
4 The ions are separated according to their mass-to-charge ratio, commonly written as ***m*/*z* ratio**.
5 The detector collects and records this information and generates a mass spectrum, which displays the *m*/*z* ratio of each particle on the *x*-axis and its percentage abundance on the *y*-axis.

***m*/*z* ratio**
mass-to-charge ratio; the molar mass of the positive ion formed divided by its charge

A **mass spectrum** shows:

- how many isotopes an element has, because each has a different *m*/*z* ratio
- the relative mass of each isotope, indicated by the *m*/*z* ratio
- how much of each isotope is in the sample, indicated by the percentage abundance.

Typically, all ions may have a single positive charge, so the lines on the spectrum correspond to each of the isotopes. The mass of each isotope is measured by comparing it to the carbon-12 isotope, which always has a mass of 12. Therefore, isotopic mass is measured relative to ^{12}C and is called **relative isotopic mass (RIM)**.

Study tip
How a mass spectrometer works is outlined in this lesson, but the syllabus does not require you to understand its operation.

How are relative isotopic mass and percentage abundance determined?

In mass spectra, the *m*/*z* ratio represents the RIM of an isotope. The mass spectra for silicon and boron are shown in Figure 2.

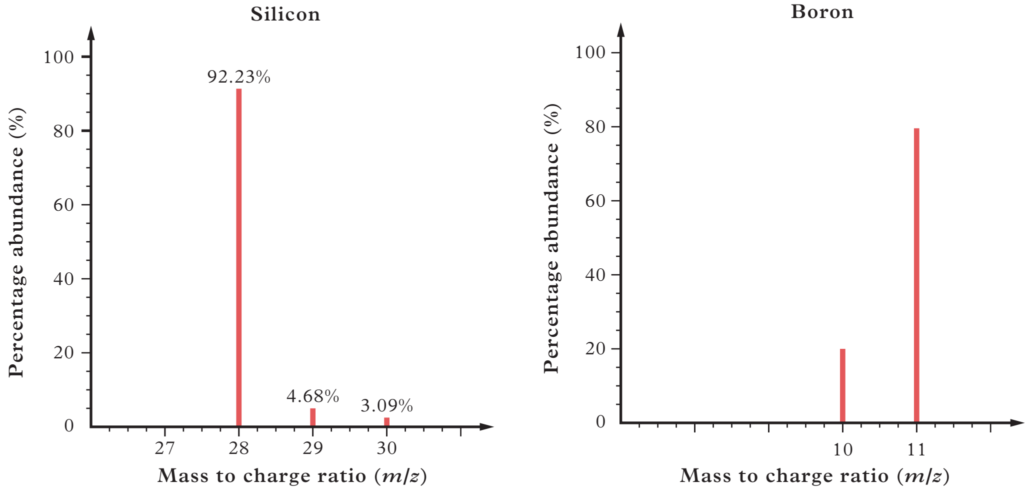


FIGURE 2 The mass spectra of silicon and boron

mass spectrum
the column graph obtained from a mass spectrometer; displays *m*/*z* ratio on the *x*-axis and percentage abundance of each positively charged ion on the *y*-axis

relative isotopic mass (RIM)
the mass of isotopes measured on a mass spectrum, relative to the carbon-12 isotope

- The mass spectrum for silicon reveals that the element has three isotopes. The isotope with a RIM of 28 (^{28}Si) has a relative abundance of 92.23%. The remaining two isotopes, which have RIMs of 29 (^{29}Si) and 30 (^{30}Si), have percentage abundances of 4.68% and 3.09%, respectively.
- In the mass spectrum of boron, the percentage abundances can be found by reading the heights of the peaks off the *y*-axis. Therefore, ^{11}B, with a RIM of 11, has a percentage abundance of 80.1%, whereas ^{10}B, with a RIM of 10, has a percentage abundance of 19.9%.

Note that these RIMs have been rounded to whole numbers. Precisely measured RIM values may be close to, but not exactly, whole numbers.

Study tip

You will study the mass spectra of molecules and their fragments in Unit 4, so make sure you are confident with the basics now!

How is relative atomic mass determined?

When calculating the mass of an element, you must consider all the isotopes of elements, and their different relative masses and abundances. The mass of an element, called its relative atomic mass (RAM), is calculated from the following equation:

$$\text{RAM} = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + (\text{RIM}_3 \times \text{abundance}_3(\%)) + \ldots}{100}$$

Worked example 4.5A

Calculating relative atomic mass of an element from a mass spectrum

Worked example 3.1A: Watch a video that shows how to solve this problem.

A sample of copper is run through a mass spectrometer and the spectrum in Figure 3 is obtained. **Calculate** the relative atomic mass of copper. (2 marks)

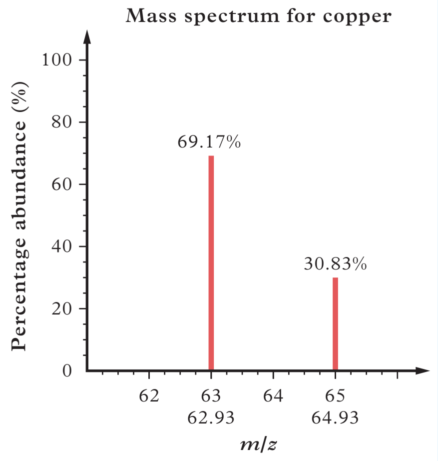


FIGURE 3 Mass spectrum for copper

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to read the information from the graph and determine the RAM. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$\text{RAM} = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%))}{100}$ From the graph: $\text{RIM}_1 = 62.93$, $\text{abundance}_1(\%) = 69.17\%$ $\text{RIM}_2 = 64.93$, $\text{abundance}_2(\%) = 30.83\%$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$\text{RAM} = \frac{(62.93 \times 69.17) + (64.93 \times 30.83)}{100}$ (1 mark) $= 63.5466$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures. RAM does not have a unit.	63.55 (1 mark)

Your turn

A sample of magnesium is run through a mass spectrometer and the following spectrum is obtained. **Calculate** the relative atomic mass of magnesium. (2 marks)

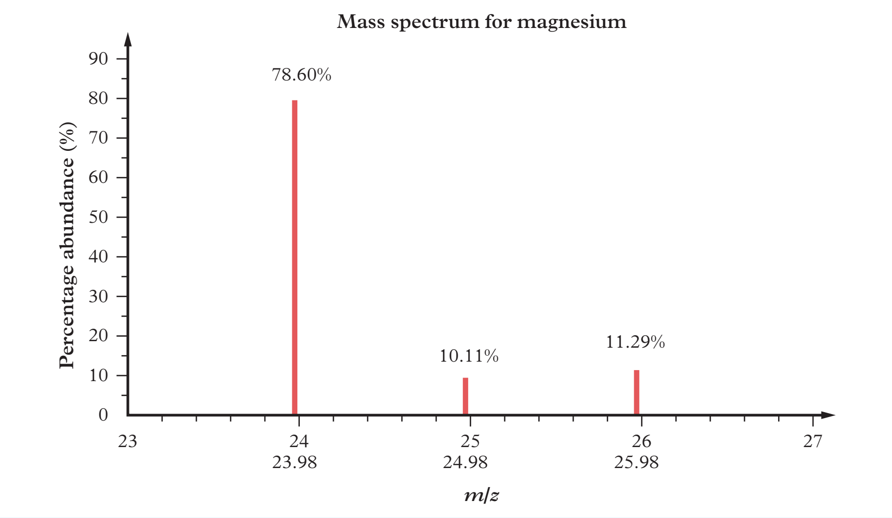


FIGURE 4 Mass spectrum for magnesium

Worked example 4.5B

Calculating relative atomic mass of an element from tabular data

Worked example 3.1A: Watch a video that shows how to solve this problem.

Naturally occurring iron has four isotopes. The data for these isotopes is summarised in Table 1. **Calculate** the relative atomic mass of iron. (2 marks)

TABLE 1 Percentage abundance of iron isotopes

Isotope	RIM	% abundance
^{54}Fe	53.94	4.850
^{56}Fe	54.94	91.75
^{57}Fe	56.94	2.120
^{58}Fe	57.93	0.280

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to read the information from the graph and determine the RAM. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$\text{RAM} = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + \ldots}{100}$ $\text{RIM}_1 = 53.94$, $\text{abundance}_1(\%) = 4.850\%$ $\text{RIM}_2 = 54.94$, $\text{abundance}_2(\%) = 91.75\%$ $\text{RIM}_3 = 56.94$, $\text{abundance}_3(\%) = 2.12\%$ $\text{RIM}_4 = 57.93$, $\text{abundance}_4(\%) = 0.280\%$

Think	Do
Step 3: Substitute the known values into the formula and solve for the unknown value.	$\text{RAM} = \frac{(53.94 \times 4.850)+(54.94 \times 91.75)+(56.94 \times 2.12)+(57.93 \times 0.280)}{100}$ (1 mark) $= 54.3929$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures. RAM does not have a unit.	54.4 (1 mark)

Your turn

Naturally occurring potassium has three isotopes. The data for these isotopes is summarised in Table 2. **Calculate** the relative atomic mass of potassium. (2 marks)

TABLE 2 Percentage abundance of potassium isotopes

Isotope	RIM	% abundance
^{39}K	38.96	93.3
^{40}K	39.96	0.0117
^{41}K	40.96	6.73

How do you determine percentage abundance from relative atomic mass?

If an element has only two isotopes, and the relative atomic mass of the element and the mass of each isotope is known, these values can be used to calculate the percentage abundance of each isotope.

The percentage abundances of the two isotopes must add up to 100% so the abundance of one isotope is denoted as x, and the abundance of the second isotope is $100 - x$.

Worked example 4.5C

Calculating percentage abundances from relative atomic mass

Worked example 3.1A: Watch a video that shows how to solve this problem.

Chlorine has a relative atomic mass of 35.45 on the periodic table. Naturally occurring chlorine has two isotopes: ^{35}Cl and ^{37}Cl. One has a relative atomic mass of 34.968 and the other has a relative atomic mass of 36.966. **Calculate** the percentage abundances of the two isotopes. (3 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to find the percentage abundances. The question is worth 3 marks, so we must gather the correct information, correctly apply the formula and complete the calculations.

Study tip

You may prefer to work in decimals when it comes to percentage abundance. This makes the calculation much simpler. The formula would become $\text{RAM} = (\text{RIM}_1 \times \text{abundance}_1) + \ldots$ but make sure you multiply your answer by 100% afterwards to convert it into a percentage.

Study tip

RAM is measured in u, formerly amu (atomic mass units). However, typically units are not written.

Think	Do
Step 2: Select the correct formula and gather any data required. The percentage abundances of the two isotopes must add up to 100. Clearly identify what is represented by x and $100 - x$.	$\text{RAM} = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%))}{100}$ $\text{RAM} = 35.45$ $^{35}\text{Cl: RIM}_1 = 34.968, \text{abundance}_1(\%) = x$ $^{37}\text{Cl: RIM}_2 = 36.966, \text{abundance}_2(\%) = 100 - x$
Step 3: Substitute the known values into the formula and solve for x, the abundance of ^{35}Cl.	$35.45 = \frac{(34.968x) + (36.966 \times (100 - x))}{100}$ (1 mark) $3545 = 34.968x + 3696.6 - 36.966x$ $1.998x = 151.6$ $x = 75.86$
Step 4: Find the abundance of ^{37}Cl.	$100 - 75.86 = 24.12$
Step 5: Finalise your answer.	^{35}Cl has an abundance of 75.86%. (1 mark) ^{37}Cl has an abundance of 24.12%. (1 mark)

Your turn

Bromine has a relative atomic mass of 79.90 on the periodic table. Naturally occurring bromine has two isotopes: ^{79}Br and ^{81}Br. One has a relative atomic mass of 78.92 and the other has a relative atomic mass of 80.92. **Calculate** the percentage abundances of the two isotopes. (3 marks)

Challenge

Calculating RIM

Silver has two isotopes. One isotope has an RIM of 106.905 and an abundance of 51.839%. If the relative atomic mass of silver is 107.9, **calculate** the RIM and abundance of the second isotope. (2 marks)

FIGURE 5 Silver exists as two isotopes.

Check your learning 4.5

Check your learning 4.5: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Identify** the missing terms from the following passage:

Mass ________________ allows scientists to precisely determine the mass of each _________ of an element. The substance is vapourised. The next step involves the ________________ of the atoms. The ions then follow a curved path through the mass spectrometer. For ions of the same charge, the ions with a ________ mass follow a path with a smaller radius and the heavier ions follow a path with a larger ________. This effectively ___________ the particles of different masses. A detector is able to measure the mass/charge _________ of each ion, as well as the relative _________. Abundances are usually converted to percentages. From this data, the relative ________ mass of each element can be found. (9 marks)

2 Gallium has a relative atomic mass of 69.72 on the periodic table. Naturally occurring gallium has two isotopes. One has a relative isotopic mass of 68.93 and the second has a relative isotopic mass of 70.92. **Calculate** the percentage abundances of the two isotopes. (3 marks)

3 Europium has a relative atomic mass of 151.96 on the periodic table. Naturally occurring europium has two isotopes. One has a mass of 150.92 and the second has a mass of 152.92. **Calculate** the percentage abundances of the two isotopes. (3 marks)

Analytical processes

4 **Analyse** the data obtained from the spectra of silver's two isotopes in Table 3 and **calculate** its relative atomic mass. (2 marks)

TABLE 3 RIM and percentage abundance for isotopes of Ag

Isotope	RIM	% abundance
^{107}Ag	106.9	51.839
^{109}Ag	108.9	48.161

5 **Analyse** the mass spectrum in Figure 6 and **calculate** the relative atomic mass of silicon. (2 marks)

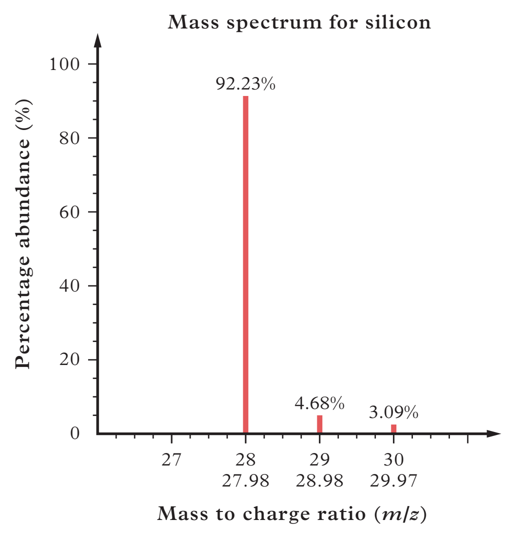


FIGURE 6 Mass spectrum for silicon

6 **Analyse** the mass spectrum in Figure 7 and **calculate** the relative atomic mass of boron. (2 marks)

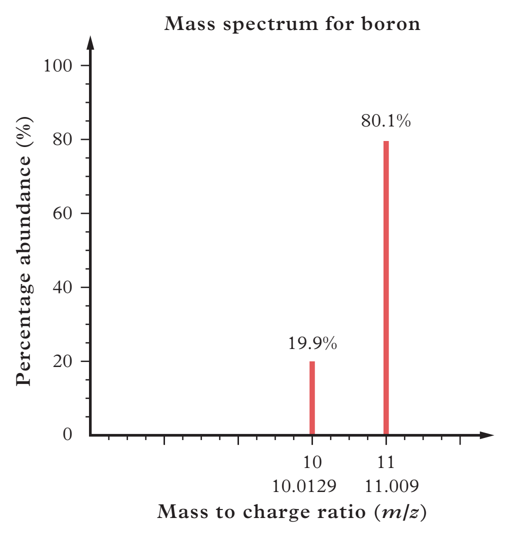


FIGURE 7 Mass spectrum for boron

7 **Analyse** the data in Table 4 and **calculate** the relative atomic mass of antimony from the spectra of its two isotopes. (2 marks)

TABLE 4 RIM and percentage abundance for isotopes of Sb

Isotope	RIM	% abundance
^{121}Sb	120.903818	57.21
^{123}Sb	122.904216	42.79

Practical

Lesson 4.6

Simulating mass spectrometry

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Lesson 4.7

Review: Analytical techniques

Summary

4.1
- Spectroscopy is a branch of physical chemistry that studies the interactions between chemicals and the electromagnetic spectrum.
- When electrons absorb electromagnetic radiation, an absorption spectrum is produced. When they emit electromagnetic radiation, an emission spectrum is produced.
- Absorption and emission spectra support Bohr's model for electron configuration.

4.2
- Flame tests are qualitative tests used to identify metals.
- Atomic absorption spectroscopy is a quantitative technique used to analyse trace levels of metals.

4.3
- Practical: Conducting a flame test

4.4
- Practical: Simulating atomic absorption spectroscopy

4.5
- Mass spectrometry involves measuring the mass-to-charge ratio of ion fragments.
- The mass and percentage abundance of isotopes, and relative atomic mass can be determined from mass spectra.

4.6
- Practical: Simulating mass spectrometry

Key formulas

Relative atomic mass	$\text{RAM} = \frac{(\text{RIM}_1 \times \text{abundance}_1(\%)) + (\text{RIM}_2 \times \text{abundance}_2(\%)) + (\text{RIM}_3 \times \text{abundance}_3(\%))}{100}$

Review questions 4.7A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Which of the following is inconsistent with Bohr's model of electron configuration?

A Electrons in atoms can be at different distances from the nucleus.

B Electrons cannot move into different orbits.

C Electrons orbit the nucleus of an atom.

D For an electron to transition between energy levels, the atom must absorb or emit energy.

2 When an atom absorbs electromagnetic radiation of a specific frequency, electrons

A spin faster around the nucleus of an atom.

B move to an electron energy level further from the nucleus.

C spin in the opposite direction.

D move into an electron energy level closer to the nucleus.

3 An emission spectrum is produced when

A electrons orbiting the nucleus are promoted to a higher energy level.

B heat from a Bunsen burner is used to excite electrons.

C electrons leave an atom's electron orbitals and an ion is formed.

D electrons shed excess energy by releasing it as light.

4 Figure 1 shows some of the lines in the absorption spectrum in the ultraviolet region for hydrogen.

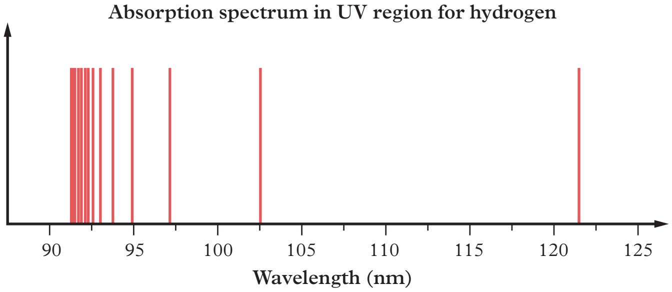


FIGURE 1 Absorption spectrum for hydrogen

At shorter wavelengths (higher energies), the spectral lines are closer together. This closer spacing provides evidence that

A hydrogen atoms have several energy levels available for the electron.

B the electron can transition between energy levels.

C the higher energy levels start to converge and have only small differences in energy.

D the electron emits light as it transitions between levels.

5 What information can a mass spectrometer provide to a chemist?

A The composition of elements in a sample

B The concentration of atoms in a sample

C The properties of molecules within a sample

D The relative mass of atoms within a sample

6 The frequency of electromagnetic radiation can be defined as

A the number of waves that pass a fixed point in one second or other unit of time.

B the energy of electromagnetic radiation, usually expressed in $J mol^{-1}$ or $kJ mol^{-1}$.

C the number of photons that pass a point in a defined period of time.

D the distance between two crests (tops) of a wave.

7 There are two naturally occurring isotopes of rhenium. The relative atomic mass is 186.2. The two isotopes are

A ^{185}Re and ^{187}Re, with relative abundances of 37.4% and 62.6%, respectively.

B ^{184}Re and ^{187}Re, with relative abundances of 94.8% and 5.2%, respectively.

C ^{185}Re and ^{188}Re, with relative abundances of 29.9% and 70.1%, respectively.

D ^{183}Re and ^{187}Re, with relative abundances of 85.6 and 14.4%, respectively.

8 When a particular element's absorption and emission spectra are studied, the lines always have the same wavelengths. This shows that

A the atoms of that element always form ions of the same charge.

B the lines in both types of spectra are due to an electron transitioning to a higher energy level.

C as electrons transition up or down between energy levels, the energy differences are the same.

D the wavelength of the light can be used to calculate the energy change of the electron.

9 Which piece of information cannot be obtained from the mass spectrum of an element with several isotopes?

A The number of isotopes present

B The relative mass of each isotope

C The energy required to ionise each isotope

D The relative amounts of each isotope in the sample

10 Which of the following lists types of electromagnetic radiation?

A Alpha rays, gamma rays, X-rays

B Beta rays, radio waves, infrared radiation

C Infrared radiation, visible light, ultraviolet radiation

D Microwaves, beta rays, X-rays

Review questions 4.7B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

11 **Explain** how absorption and emission spectra are obtained. (4 marks)

12 **Recall** the difference between relative isotopic mass (RIM), relative atomic mass (RAM) and mass number. (3 marks)

13 **Describe** the relationship between the terms in question **12**. (3 marks)

14 AAS is a valuable analytical technique.

- a **Describe** what happens to atoms during AAS and how this is detected. (3 marks)
- b **Explain** whether, for the same element, the wavelengths for absorption and for emission would be the same. (2 marks)

Analytical processes

15 **Determine** the relative atomic mass of the following elements from the isotope data given.

- a Silicon has three isotopes. (2 marks)

TABLE 1 RIM and abundances for the isotopes of silicon

Isotope	RIM	% abundance
^{28}Si	27.977	92.229
^{29}Si	28.976	4.683
^{30}Si	29.974	3.087

- b Europium has two isotopes. (2 marks)

TABLE 2 RIM and abundances for the isotopes of europium

Isotope	RIM	% abundance
^{151}Eu	150.9198	47.81
^{153}Eu	152.9212	52.19

- c Zinc has five isotopes. (2 marks)

TABLE 3 RIM and abundances for the isotopes of zinc

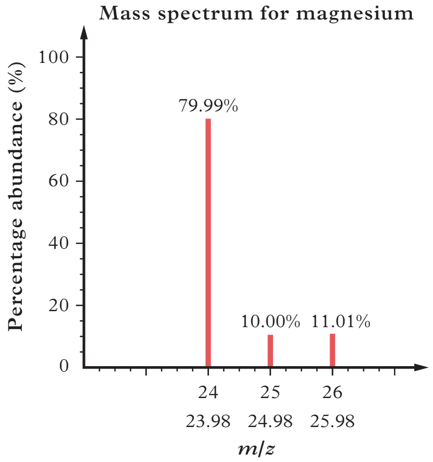
Isotope	RIM	% abundance
^{64}Zn	63.929	48.63
^{66}Zn	65.926	27.90
^{67}Zn	66.927	4.10
^{68}Zn	67.925	18.75
^{70}Zn	69.925	0.62

16 **Determine** the percentage abundance of the isotopes, given the following information.

- a Thallium has a relative atomic mass of 204.38. Its isotopes have relative isotopic masses of 202.97 and 204.97. (3 marks)
- b Indium has a relative atomic mass of 114.8. Its isotopes have relative isotopic masses of 112.904 and 114.904. (3 marks)
- c Rubidium has a relative atomic mass of 85.5. Its isotopes have relative isotopic masses of 84.912 and 86.909. (3 marks)

17 **Determine** the relative atomic mass from the following mass spectra.

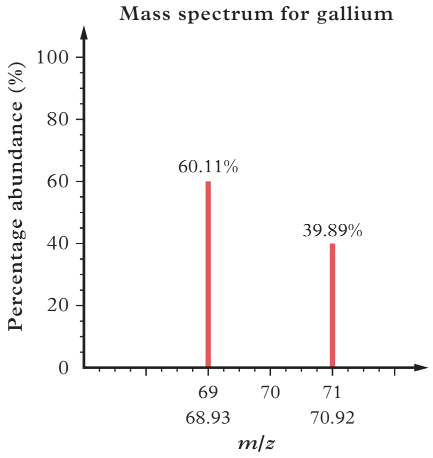
a



(2 marks)

FIGURE 2 Mass spectrum for Mg

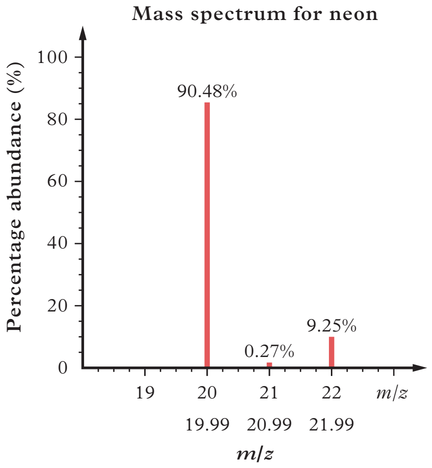
b



(2 marks)

FIGURE 3 Mass spectrum for Ga

c



(2 marks)

FIGURE 4 Mass spectrum for Ne

18 **Determine** the relative atomic mass for the following elements from the isotope data given.

a Isotope data for titanium (2 marks)

TABLE 4 RIM and abundances for the isotopes of titanium

Isotope	RIM	% abundance
^{46}Ti	45.953	8.250
^{47}Ti	46.952	7.440
^{48}Ti	47.948	73.72
^{49}Ti	48.948	5.410
^{50}Ti	49.945	5.180

b Isotope data for chromium (2 marks)

TABLE 5 RIM and abundances for the isotopes of chromium

Isotope	RIM	% abundance
^{50}Cr	49.946	4.3450
^{52}Cr	51.941	83.789
^{53}Cr	52.941	9.5010
^{54}Cr	53.939	2.3650

19 **Determine** the relative atomic mass of the following elements using the mass spectra shown.

a Mass spectrum for rhenium. The relative isotopic masses are 184.950 and 186.956, respectively. (2 marks)

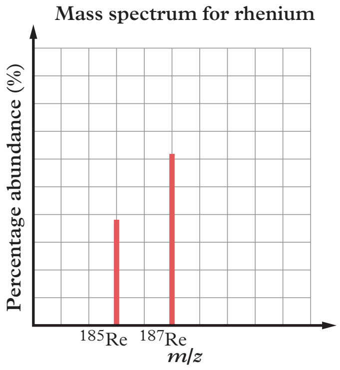


FIGURE 5 Mass spectrum for Re

b Mass spectrum for lead. The relative isotopic masses are 203.973, 204.974, 206.976 and 207.976, respectively. (2 marks)

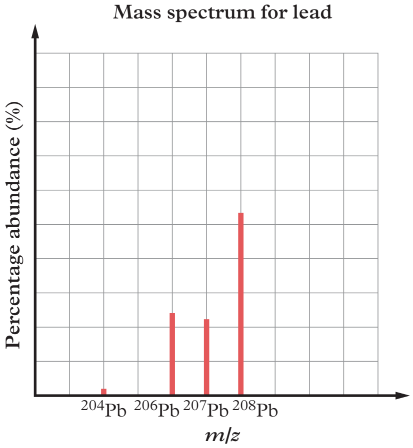


FIGURE 6 Mass spectrum for Pb

20 AAS is commonly used to quantify the amount of substance in a sample, with the help of a calibration curve.

a The magnesium content of blood is analysed against a set of standards and the data in Table 6 is obtained. **Analyse** this data using a calibration curve to **determine** the concentration of magnesium. (2 marks)

TABLE 6 Magnesium concentrations of standards

Standard/ sample	Concentration (mg L^{-1})	Absorbance
Standard 1	0	0.000
Standard 2	4	0.230
Standard 3	8	0.460
Standard 4	12	0.690
Standard 5	16	0.920
Sample	–	0.530

b The calcium concentration in a sample of milk is analysed against a set of standards and the data in Table 7 is obtained. **Analyse** this data using a calibration curve to determine the concentration of calcium. (2 marks)

TABLE 7 Calcium concentrations of standards

Standard/ sample	Concentration (mmol L^{-1})	Absorbance
Standard 1	0	15
Standard 2	10	30
Standard 3	20	45
Standard 4	30	60
Standard 5	40	75
Sample	–	20

- **c** **Compare** the quality of data in Tables 6 and 7, with reference to any potential errors. (3 marks)

21 **Determine** the percentage abundance of the isotopes, given the following information.

- **a** Lithium has a relative atomic mass of 6.9. Its isotopes have relative isotopic masses of 6.015 and 7.016. (3 marks)
- **b** Iridium has a relative atomic mass of 192.2. Its isotopes have relative isotopic masses of 190.96 and 192.96. (3 marks)

Data drill

Atomic absorption spectroscopy

An analysis was conducted by AAS to determine the concentration of magnesium ions contained in a common antacid tablet. The sample was tested along with a set of magnesium ion standards; the calibration curve is shown in Figure 1. The absorbance of the solution made from the tablet was 0.24.

A second solution was made from a magnesium supplement. The label on the supplement states that each tablet contains 110 mg of magnesium. One tablet is crushed and dissolved in 2.0 L of distilled water.

When the absorbance of this solution was tested, it was found to be 0.68. It is known that in general, the calibration curve for AAS is not linear at higher concentrations but tends to give lower readings than expected.

To re-check, a 20.0 mL sample of the second solution was diluted by adding enough water to make 40.0 mL, halving the concentration. The absorbance of the diluted sample was found to be 0.37.

Apply understanding

1 **Identify** the concentration, in mg L^{-1}, of the solution made from the antacid tablet. (1 mark)

2 **Calculate** the concentration of the second solution, made from the supplement. (1 mark)

Analyse evidence

3 **Identify** the trend shown on the calibration curve. (1 mark)

Interpret evidence

4 **Extrapolate** the curve to **determine** the absorbance expected for the concentration calculated in question **2**. (2 marks)

5 **Deduce** one possible reason for the actual absorbance being less than that determined in question **4**. (1 mark)

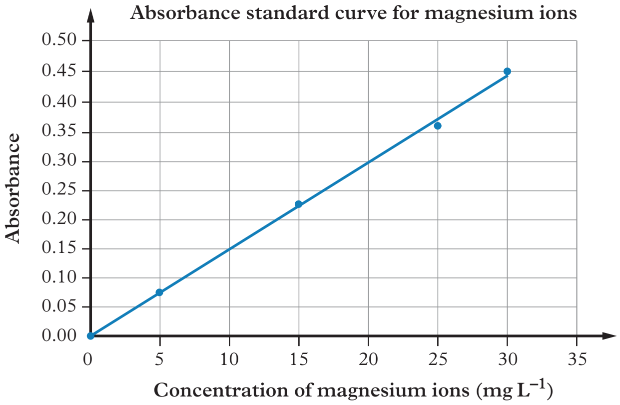


FIGURE 1 Absorbance calibration curve for magnesium ions

Module 4 checklist: Analytical techniques

Quizlet: Revise key terms online to test your understanding

UNIT 1

Topic 1 review

Multiple choice

(1 mark each)

1 Which of the following ions contains the least number of electrons?

A Be^{2+}

B O^{2-}

C He^{+}

D B^{3+}

2 The mass spectrum of naturally occurring magnesium is shown in Figure 1.

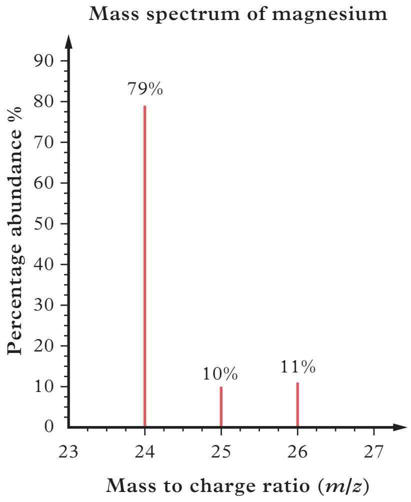


FIGURE 1 Mass spectrum of magnesium

Based on the mass spectrum, the relative atomic mass of magnesium is

A 24.32.

B 24.23.

C 25.00.

D 25.92.

3 Which of the following species has the largest radius?

A F

B K

C Ar

D Na

4 Why does solid iron metal conduct electricity?

A Iron atoms react to form alternating cations and anions.

B The bonding electrons within the iron lattice can move.

C Iron atoms can move to an oppositely charged electrode.

D The iron lattice contains freely moving delocalised electrons.

5 What is the total number of electrons in the d orbitals of a single bromine atom?

A 5

B 10

C 15

D 20

6 Consider the following electron configuration:

$$1s^2\,2s^2\,2p^6\,3s^6\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2.$$

According to the Aufbau principle, the next orbital sublevels that will be filled are

A 4d, then 5s.

B 5p, then 6s.

C 4d, then 4f.

D 4d, then 5p.

7 What is the composition of a nucleus of fluorine-17?

A 9 protons and 9 neutrons

B 8 protons and 8 neutrons

C 8 protons and 9 neutrons

D 9 protons and 8 neutrons

8 Which of the following properties of the period 4 elements increase going from potassium to krypton?

A Nuclear charge and atomic radius

B Atomic radius and nuclear charge

C Nuclear charge and electronegativity

D Atomic radius and electronegativity

9 Which physical property generally decreases down a group but increases from left to right across a period?

A Atomic radius
B Electron density
C Electron affinity
D Ionisation energy

10 Which of the following are covalent compounds?

I NH_4Cl **II** NH_4NO_3 **III** HCl

A III only
B I and II only
C I, II and III
D II and III only

11 What is the formula for the compound formed by aluminium and sulfur?

A Al_3S_3
B Al_2S_2
C Al_3S_2
D Al_2S_3

12 Which of the following interact through hydrogen bonding?

I HBr **II** HCl **III** CH_3OH

A I and II only
B I, II and III
C I and III only
D II and III only

13 What is the correct name for the compound with a molecular formula of N_2F_6?

A Nitrogen fluoride
B Nitrogen trifluoride
C Nitrogen hexafluoride
D Dinitrogen hexafluoride

14 Figure 2 shows the variation of a physical property for the first 20 elements in the periodic table. What is the property?

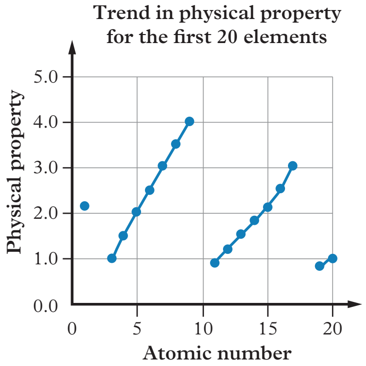


FIGURE 2 Graph of variation of a physical property

A Atomic radius
B Melting point
C Electronegativity
D First ionisation energy

15 A sample of urea contains 2.410 g of N, 0.322 g of H, 0.960 g of C and 1.280 g of O. What is the empirical formula?

A N_2H_2CO
B $N_2H_4C_2O$
C N_2H_4CO
D $N_2H_4CO_2$

Short response

16 Nitrogen and phosphorus belong to the same group of the periodic table.

a **Distinguish** between the terms “group” and “period” in terms of electron arrangement. (2 marks)

b **Identify** the period numbers of nitrogen and phosphorus. (1 mark)

17 The electron configuration of manganese can be expressed as $[Ar]4s^x3d^y$.

a **Explain** what the square brackets around argon [Ar] represent. (1 mark)

b **Deduce** the values of x and y. (1 mark)

c **Consider** the following statement: “After the removal of the first electron from the calcium atom, it is easier to lose the second electron and achieve a noble gas configuration as Ca^{2+}.” **Identify** the incorrect elements within this statement and **justify** your response. (3 marks)

18 A researcher was able to detect the presence of three species (X, Y and Z) with varying abundances. They investigated further to determine the mass and subatomic composition. The results obtained are shown in Table 1.

TABLE 1 Data of three species

Species	Number of protons	Number of neutrons	Number of electrons	Percentage abundance (%)
X	14	14	14	92.23
Y	14	15	14	4.680
Z	14	16	14	3.090

- **a** Species X, Y and Z were determined to be isotopes of each other. **Identify** one piece of evidence from Table 1 that supports this conclusion. (1 mark)
- **b** **Calculate** the relative atomic mass of the element using the data in Table 1. (3 marks)
- **c** **Identify** the element investigated by the researcher. (1 mark)

19 **Compare** the structure and bonding between silicon and oxygen atoms in silicon dioxide and carbon and oxygen atoms in carbon dioxide. (4 marks)

20 Aluminium is the most abundant metal in the Earth's crust. The successive ionisation energies of aluminium are shown in Figure 3.

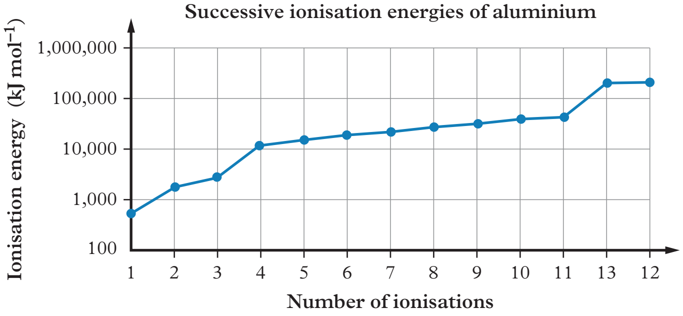


FIGURE 3 Successive ionisation energies of aluminium

- **a** **Construct** an equation for the first ionisation of aluminium. (1 mark)
- **b** **Explain** why ionisation energy increases with each ionisation. (2 marks)
- **c** **Infer** a reason for the large increase between the third and fourth ionisation energies. (3 marks)

21 The emission spectra for five elements and an unknown sample are provided in Figure 4.

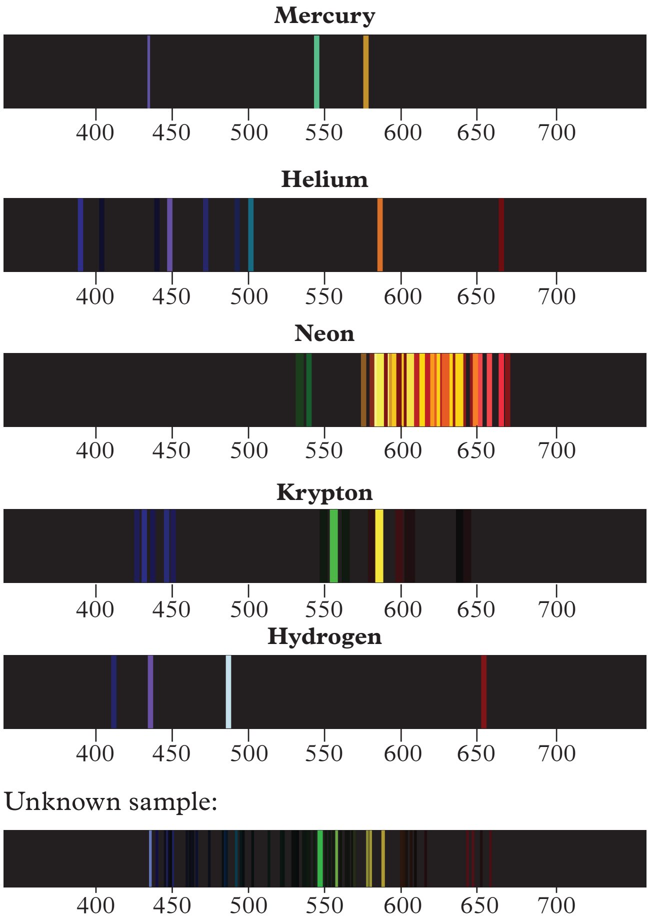


FIGURE 4 The emission spectra for five elements and an unknown sample

- **a** **Identify** all of the elements present in the unknown sample. (2 marks)
- **b** **Explain** how the emission spectra are produced. (2 marks)

22 Ionic radius is the distance between the nucleus of an ion and its valence electron shell.

- **a** **Describe** two reasons why the calcium ion has a smaller radius than a calcium atom. (2 marks)
- **b** **Explain** why the ionic radius of Si^{4+} is less than the ionic radius of P^{3-}. (3 marks)

23 **Determine** the Lewis structures, including non-bonding electrons, for the following species.

- **a** Ethanal (CH_3CHO) (1 mark)
- **b** Phosphate ion (PO_4^{3-}) (2 marks)
- **c** Carbon dioxide (CO_2) (1 mark)

24 **Describe** the acid–base character of the oxides produced from period 2 elements. **Determine** a balanced equation to illustrate this trend for sodium oxide and sulphur trioxide. (4 marks)

25 **Deduce** the names of the following compounds and **identify** the type of bonding involved:

- **a** NH_4SO_4 (2 marks)
- **b** Br_2O (2 marks)
- **c** $K_2Cr_2O_7$. (2 marks)

26 Potassium has a relative atomic mass of 39.10. It exists as three stable isotopes: ^{39}K, ^{40}K, ^{41}K, which have relative isotopic masses of 38.964, 39.964 and 40.962, respectively. The heaviest isotope ^{41}K has a percentage abundance of 6.73%. **Calculate** the percentage abundance of the lightest isotope to two decimal places. (2 marks)

27 A chemist wants to determine the lead content in tap water by atomic absorption spectroscopy. After initial testing, she determines that the sample is too concentrated and dilutes it by a factor of 10. She then reruns the sample with a set of standards and obtains the data in Table 2.

TABLE 2 Lead content in tap water

Lead concentration ($mg\ L^{-1}$)	Absorbance
0.0	0.000
2.0	0.172
4.0	0.348
6.0	0.530
8.0	0.694
10.0	0.892
12.0	0.998
Unknown water sample (diluted 10×)	0.197

- **a** **Construct** a standard curve for the known concentrations of lead. (2 marks)
- **b** **Use** the graph to **determine** the concentration of lead in the sample of tap water. (3 marks)
- **c** If the safe limit for lead in water is $0.01\ mg\ L^{-1}$, **use** the results to **evaluate** how safe the water sample is to drink. (2 marks)

TOTAL MARKS

/70

MODULE 5 Compounds and mixtures

Introduction

All matter is made of atoms. Elements contain the same type of atoms, existing as individual atoms or atoms held together in molecules by chemical bonds. When a compound is formed, atoms of two or more different elements in a fixed ratio are held together by chemical bonds. Elements and compounds are pure substances.

A mixture is made of two or more elements and/or compounds that do not react chemically. Mixtures are not pure substances.

When producing materials, it is essential to know the chemical composition of those materials. Scientists involved in quality control follow strict testing procedures to ensure that the chemical and physical properties of production materials are within quality assurance guidelines.

This module focuses on the physical properties of materials, and how those properties can be used for identification and separation of substances.

Prior knowledge

Prior knowledge quiz

Check your understanding of particle theory and physical properties before you start.

Subject matter

Science understanding

- → State that pure substances may be elements or compounds.
- → Identify that pure substances have distinct measurable properties (e.g. melting and boiling point, reactivity, strength, density) and mixtures have properties dependent on the identity and relative amounts of the substances that make them up.
- → Discriminate between heterogeneous and homogeneous mixtures.
- → Analyse data to determine the physical properties of pure substances and mixtures.

Science as a human endeavour

→ Appreciate that impurities can affect the physical and chemical properties of substances, resulting in inefficient or unwanted chemical reactions.

→ Explore Ellen Swallow Richard's contribution to developing water quality standards and the new discipline of home economics.

Science inquiry

→ Investigate the separation of mixtures based on physical properties.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

oxforddigital **These lessons are available on Oxford Digital.**

Lesson 5.3 Separating mixtures

Lesson 5.1

Pure substances

Key ideas

→ Pure substances are made up of elements or compounds.

→ Pure substances have distinct, measurable properties, including melting point, boiling point, reactivity, strength and density.

What is a pure substance?

The world is made up of substances. The glass on a phone screen, a hot green tea and soil: these items are all made up of **mixtures** of substances.

Pure substances are not mixed with any other type of substance; they are made up of a single type of element or **compound** (Figure 1). In other words, a substance that contains only one type of element or one type of compound is considered pure. A pure substance has no impurities and can exist in different phases (e.g. solid, liquid or gas).

Learning intentions and success criteria

mixture
a substance that is made up of two or more substances that are not chemically bonded

pure substance
a substance that is made up of a single type of element or compound

compound
a substance that is made up of atoms of different elements, joined together by chemical bonds

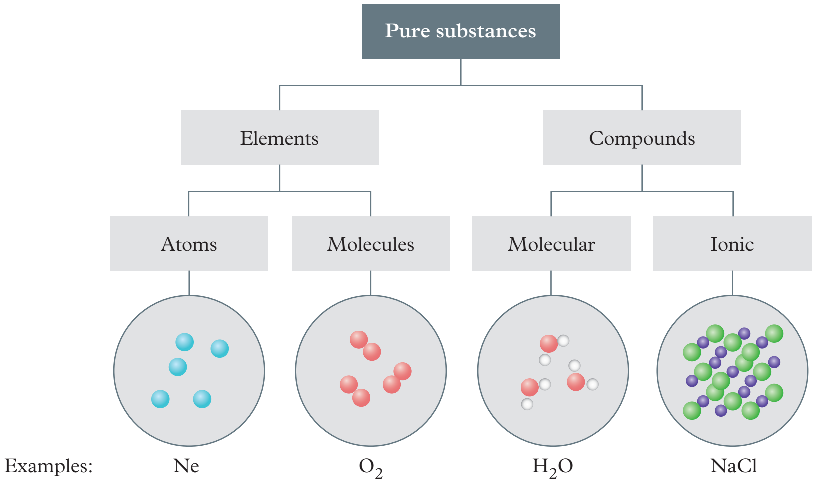


FIGURE 1 Pure substances consist of elements or compounds.

For example, water molecules grouped together are considered a pure substance. However, if sodium chloride (NaCl) is added to the water, it is no longer a pure substance because more than one kind of compound is present.

Table 1 summarises the key features of elements and compounds.

TABLE 1 The key features of elements and compounds

Type of substance	Description	Examples
Element	Made up of one type of atom	Oxygen (O_2), hydrogen (H_2) and carbon (C)
Compound	Made up of atoms of different elements, joined together by chemical bonds	Water (H_2O) and carbon dioxide (CO_2)

What are the properties of pure substances?

Pure substances have distinct, measurable properties. The purity of a substance can be established by collecting or analysing data of its physical and chemical properties, such as melting point, boiling point and density. Pure substances have characteristic reactivities and bond strengths.

TABLE 2 Physical properties of some substances under standard laboratory conditions

Element	Physical state	Melting point (°C)	Boiling point (°C)	Density (g cm^{-3})
Water (H_2O)	liquid	0	100	0.997
Oxygen	gas	−218.79	−182.962	1.429×10^{-3}
Nitrogen	gas	−210	−195.975	1.2504×10^{-3}
Mercury	liquid	−38.8290	356.73	13.534
Iron	solid	1,538	2,862	7.874

Melting point and boiling point

Melting point refers to the temperature at which a solid substance becomes a liquid. A pure solid substance should have a narrow melting point. In other words, if the substance is pure, all of the substance should melt at the same temperature. Boiling point refers to the temperature at which a liquid substance becomes a gas.

Impurities in a substance typically affect its melting and boiling points because the different compounds may have a range of melting and boiling points.

For example, the boiling point of pure water is approximately 100°C at a pressure of 1 bar. More accurately, it is 99.61°C. However, any dissolved impurities in the water can increase the boiling point (Figure 2). Melting points of mixtures are generally lower than that of the separate pure substances, and they also melt over a range of temperatures.

FIGURE 2 The boiling point of water can be affected by impurities in the water.

Reactivity

Reactivity refers to how readily a substance reacts with other substances. During a chemical reaction, reactant bonds are broken and product bonds are formed. When a substance is pure, we can predict the products that can form from certain chemical reactions. The potential energy changes associated with these processes determine whether a reaction will occur and can influence the rate at which the chemical reaction occurs.

reactivity
a measure of how readily a substance reacts with other substances

Bond strength

The chemical bonds that are present in a substance directly influence the properties of that substance. Covalent bonds are extremely strong and require a lot of energy to break. Pure substances with very strong covalent bonds extending throughout the substance in three dimensions are among the hardest substances known and have extremely high melting points. Examples are diamond, which is pure carbon, and quartz, which is silicon dioxide.

But this is not true of all covalent substances. It may seem strange that many covalent substances have quite low melting points. The low melting and boiling points are due to the much weaker intermolecular forces. For example, water melts at 0°C and boils at 100°C. The product of boiling water is water vapour, and the strong covalent bonds within the water molecules have not been broken.

Sodium chloride contains strong ionic bonds. It forms a hard substance with high melting and boiling points.

Density

Density is the mass per unit volume of a substance. The density of a substance is based on the relative atomic masses of the nuclei present and how compact the atoms or ions are. If the mass and volume of a substance are known, the following formula can be used to calculate density:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

The units of density depend on the units of mass and volume used in the calculation. If mass is measured in grams and the volume is measured in cubic centimetres, the units of density are grams per cubic centimetre (g cm^{-3}).

Worked example 5.1A

Calculating density

Worked example 5.1A: Watch a video that shows how to solve this problem.

A student measured the mass of a measuring cylinder to be 15.23 g. A 10.0 cm^3 volume of ethanol was added to the measuring cylinder, and it was reweighed. The mass was now 23.12 g.

a **Calculate** the mass of ethanol. (1 mark)

b **Calculate** the density of ethanol at this temperature. (2 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to apply the density formula. We must gather the correct information, correctly apply the formula and complete the calculations.

Study tip

Including units when substituting values helps to check that the units of the final answer are correct for that quantity.

Study tip

Spaced learning assists in long-term recall. Check your learning 5.1 question 2 as an opportunity to practise prior learning of naming elements and compounds and writing formulas.

Think	Do
Step 2: Select the correct formula and gather any data required.	Density = $\frac{\text{mass}}{\text{volume}}$ Density = ?, mass = ?, volume = 10.0 cm^3
Step 3: For part **a**, find the mass of the substance based on the measurements provided.	mass(ethanol) = 23.12 g – 15.23 g = 7.89 g (1 mark)
Step 4: For part **b**, substitute the known values into the formula and solve for the unknown value.	Density = $\frac{7.89\text{ g}}{10.0\text{ cm}^3}$ (1 mark) = 0.789 g cm^{-3}
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	0.789 g cm^{-3} (1 mark)

Your turn

An empty measuring cylinder has a mass of 35.75 g. 18.5 cm^3 of water was added. The measuring cylinder was reweighed and found to have a mass of 54.20 g.

a **Calculate** the mass of water. (1 mark)

b **Calculate** the density of water at this temperature. (2 marks)

Check your learning 5.1

Check your learning 5.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Recall** the two types of pure substances and give an example of each. (2 marks)

2 **Identify** the missing information for the pure substances listed in Table 3. The first row has been done for you. (5 marks)

TABLE 3 Formulas and classification of pure substances

Name	Formula	Element or compound?
Magnesium nitrate	$Mg(NO_3)_2$	compound
Chlorine		
Sulfur trioxide		
Aluminium oxide		
	Fe	
	Li_2CO_3	

3 **Recognise** five measurable properties of pure substances that can be used to help identify a substance. (5 marks)

Analytical processes

4 A sample of iron has a volume of 8.00 cm^3 and a mass of 57.55 g. The density of pure iron is known to be 7.87 g cm^{-3}.

a **Calculate** the density of the sample. (1 mark)

b **Determine** if the sample of iron is pure. (1 mark)

5 **Analyse** the melting points of the unknown substances in Table 4.

a **Deduce** which of the substances is most likely to be pure water. (1 mark)

b **Compare** the melting point of substance B with the melting points of the other substances listed. (2 marks)

c **Infer** why substance B behaves differently when melted. (1 mark)

TABLE 4 Melting points of some unknown substances

Substance	Melting point (°C)
A	–114
B	8–26
C	0
D	6.5

Lesson 5.2
Mixtures

Key ideas

→ The properties of mixtures depend on the identity and relative amounts of their components.

→ Heterogeneous mixtures are uniform, whereas homogeneous mixtures do not have a consistent composition.

→ Mixtures can be separated based on the physical properties of the components.

What is a mixture?

Learning intentions and success criteria

A mixture is made up of two or more substances that are not chemically bonded. The substances in a mixture retain their individual properties. Therefore, the properties of a mixture depend on the identity and amount of each substance that is a component of the mixture. For example, fried rice is a mixture that can be easily separated into its different components.

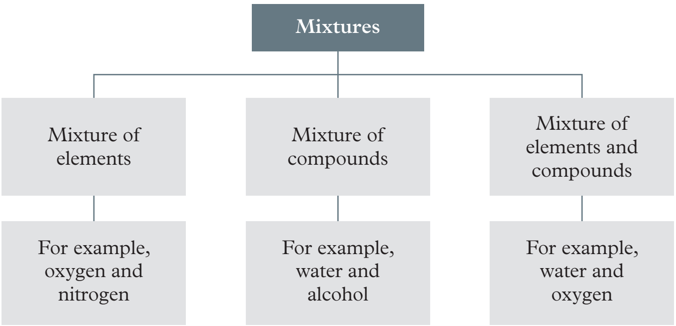


FIGURE 1 Mixtures may consist of a mixture of elements, a mixture of compounds or a mixture of elements and compounds

Mixtures are different from pure substances in several ways:

- Mixtures can be physically separated.
- Mixtures have not undergone a chemical reaction to combine or bond their components.
- The properties of a mixture depend on the substances that form the mixture.
- The composition of a mixture is not always the same throughout the mixture.

Mixtures are also different from compounds, whose substances are chemically bonded together (Table 1). For example, water is made up of hydrogen and oxygen, but it is not considered a mixture because the elements are chemically bonded together; therefore, water is considered a compound.

TABLE 1 Differences between a mixture and a compound

	Structure	Properties	Separation
Mixture	The amount of each substance in a mixture can vary. The different substances are not chemically bound. A chemical formula cannot be written for a mixture.	The substances in a mixture retain their own properties.	A mixture can be separated based on the physical properties of its component substances.
Compound	The amount of each element in a compound does not vary. The different elements are chemically bonded in a fixed ratio and can be represented by a chemical formula.	A compound has properties different from those of its component elements.	A compound cannot be separated based on its physical properties.

Skill drill

Analysing melting points to identify substances

Science inquiry skill(s): Processing and analysing data (Lesson 1.7)

Pure substances have fixed measurable melting points, so the melting point can be useful to identify a substance. However, you should not rely on it as the only evidence, as substances can share the same melting point.

When a pure substance is contaminated with small amounts of another substance, the resulting impure substance will melt over a wider range of temperatures that are lower than the known melting point of the pure substance.

The melting points of four samples are measured and shown in Table 2. The four samples can be any of the seven substances in Table 3, and they may be pure or impure.

TABLE 2 The melting points of the four unknown samples

Sample	Melting point range (°C)
Sample A	125–130
Sample B	172–173
Sample C	70–71
Sample D	115–121

TABLE 3 A list of the melting points of some pure substances; there is a very slight range in the accepted values.

Substance	Melting point range (°C)
Ferrocene	172–174
Naphthalene	80–81
Biphenyl	69–70
Aspirin	136–138
Benzoic acid	122–124
Stearic acid	69–70
2-napthol	121–123

Practise your skills

1 **Deduce** which samples are likely to be pure and which are likely to be mixtures. **Justify** your answers using evidence. (4 marks)

2 **Analyse** Tables 2 and 3 to **determine** the identities of the four samples. (4 marks)

3 For the substances that could not clearly be identified, **discuss** why. (2 marks)

What are homogeneous and heterogeneous mixtures?

Mixtures can be classified as homogeneous or heterogeneous as shown in Figure 2.

Homogeneous mixtures

A **homogeneous mixture** has a uniform composition. It is made up of substances of the same state (gas, liquid or solid), and is consistent throughout the mixture. A sample taken from a homogeneous mixture would appear the same and have the same composition, no matter where the sample was taken from in the mixture. For example, brewed tea, salt solution or wine are all mixtures that contain a number of substances but have a uniform composition. A solution, such as cordial, is also a homogeneous mixture.

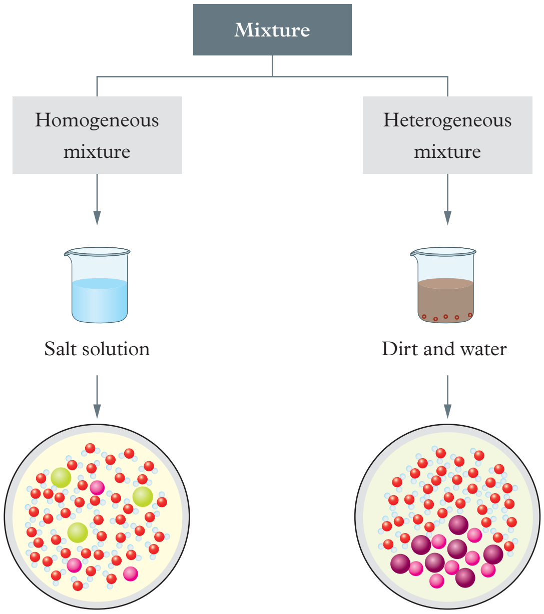


FIGURE 2 A salt solution is an example of a homogeneous mixture, and dirty water is an example of a heterogeneous mixture of soil and water.

Heterogeneous mixtures

A **heterogeneous mixture** has an inconsistent composition. For example, a mixture of oil, water and sand is heterogeneous as the three substances do not mix well, and their different physical properties are clearly identifiable (Figure 3A). If a sample was taken from the top of the mixture, it is likely to display the properties of the oil at the top, rather than properties that represent the whole mixture.

Foods such as salads, muesli or a bowl of mixed lollies are also heterogeneous mixtures (Figure 3B). In these mixtures, it is clear that there are a number of substances present, which can be easily separated. One sample of the mixture may have different properties from another sample of the mixture.

homogeneous mixture
a mixture with a uniform composition

heterogeneous mixture
a mixture with an inconsistent composition

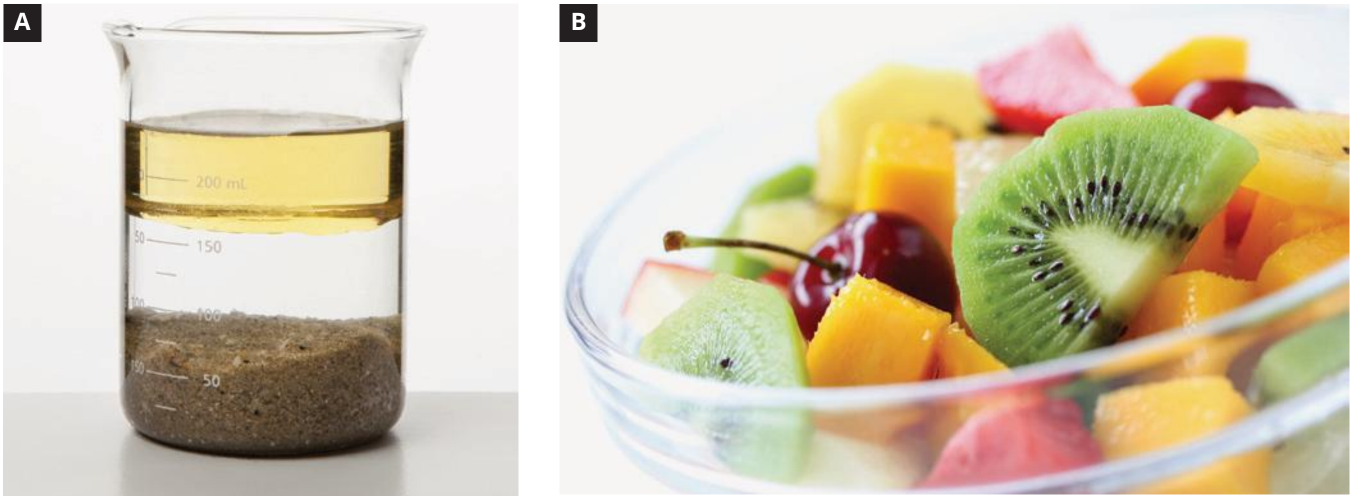


FIGURE 3 Two different heterogeneous mixtures: (A) oil, water and sand, and (B) a fruit salad

colloid
a heterogenous mixture in which the dispersed insoluble substance does not separate over time

Table 4 shows the different types of mixtures that include solutions, suspensions and **colloids**.

TABLE 4 Different types of mixtures

Type of mixture	Description	Example	Homogeneous or heterogeneous?
Solution	A mixture in which a substance is dissolved evenly in another substance	Sugar in water	homogeneous
Suspension	A mixture in which a substance is dispersed (but not dissolved) in another substance; will separate over time	Oil and sand in water	heterogeneous
Colloid	A mixture of an insoluble substance in another substance. The dispersed substance is finely divided, so the components do not separate over time.	Milk	heterogeneous

How are physical properties used to separate mixtures?

centrifugation
the technique for separating mixtures where tubes are placed into a centrifuge that spins in a circular motion very fast to push solids or higher density liquids to the bottom

flotation
the technique for separating different substances in a powdered ore depending on whether they sink in, or float on, a given liquid

The physical properties of substances can be used to separate substances from mixtures. Therefore, it is useful to understand the physical properties of each of the individual substances within the mixture.

Different techniques for separation, including distillation, crystallisation by evaporation, filtration, using a separating funnel, **centrifugation**, decantation, **flotation**, sedimentation and physical separation, have been developed on the basis of these physical properties. The technique chosen for separation will depend on the properties of the substance being separated from the mixture. Some of these techniques are outlined in Table 5.

TABLE 5 Techniques for separating mixtures based on physical properties

Physical property	Techniques for separation	Example
Boiling point	**Distillation**: A liquid can be separated from a liquid mixture by boiling the mixture until the liquid evaporates and is then collected by condensation. The liquid substance with the lower boiling point evaporates first.	Separating ethanol from water in the production of spirits
	Crystallisation by evaporation: Evaporating the solvent can separate a solid substance that is dissolved in a liquid mixture.	Collecting salt from a solution of salt and water by evaporating the water (solvent) to leave the salt crystals behind
Density	**Sedimentation and decantation**: Dense solid substances in a liquid mixture will naturally sink to the bottom of a container over time because of gravity (sedimentation). The liquid can then be decanted or carefully poured off the top to leave the undisturbed solids behind.	Separating soil from a mixture of soil and water, by allowing soil sink to the bottom of the mixture
	Using a separating funnel: If two liquids do not dissolve in each other, but form a heterogeneous mixture, they can be separated using a separating funnel.	The more dense liquid forms a layer at the bottom. The stop-cock (tap) can be opened for long enough for the bottom layer to be released and collected in a beaker.

FIGURE 4
A separating funnel

Physical property	Techniques for separation	Example
Size	**Physical separation**: Some substances in mixtures can be easily removed by physically picking them out of the mixture, either by hand or with a tool such as tweezers.	Picking out a chocolate chip from a mixture of cookie dough ingredients
	Filtration: A solid substance can be separated from a liquid mixture by pouring the mixture through a filter (a porous barrier such as filter paper). The liquid passes through the filter, while the solid is caught in the filter.	Filtering tea leaves using a strainer

Figure 5 shows how the apparatus used in distillation separate the components of a homogeneous mixture. This relies on the substances making up the mixture having different boiling points.

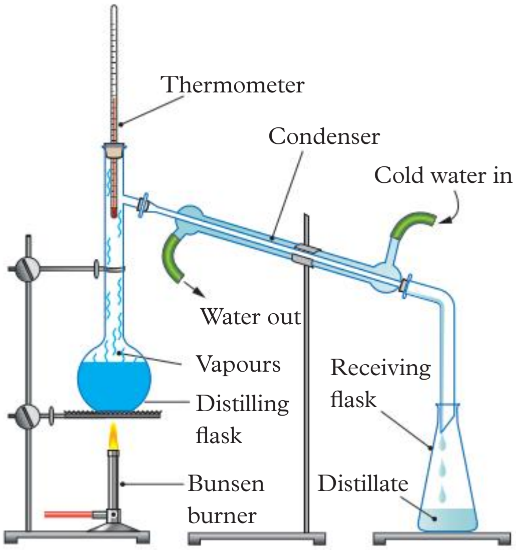


FIGURE 5 Distillation apparatus for boiling and then condensing liquids to separate a homogeneous liquid mixture

Figure 6 shows two techniques for separating heterogeneous mixtures: sedimentation and decantation, and filtration.

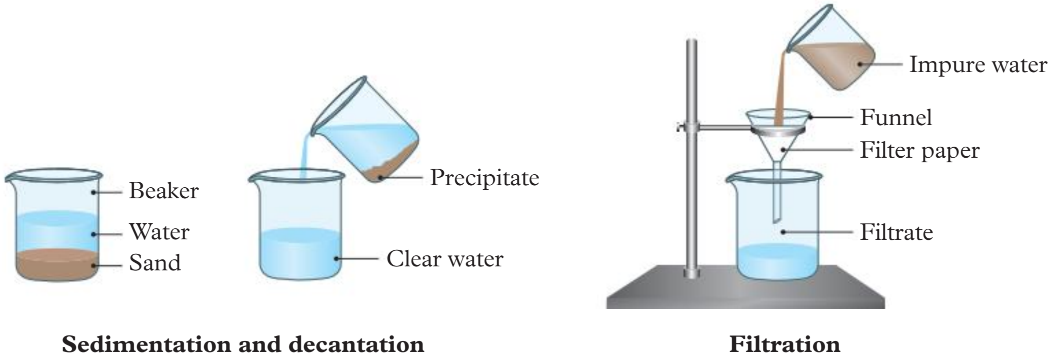


FIGURE 6 Techniques for separating heterogeneous mixtures

Real-world chemistry

Ellen Swallow Richards: Developing water quality standards

Ellen Swallow Richards (1842–1911) from Massachusetts, USA, studied chemistry and became the first American woman to obtain a science degree. Her research was diverse, starting with geology and leading her towards water and air quality. She applied her knowledge to the science of hygiene in the home, monitoring of pollutants and designing treatments of wastewater.

Richards introduced the term "ecology" to the English language. She saw the importance of monitoring and controlling chemical contamination of water and treatment of bacterial contamination for the good health of humans and the environment.

Richards knew larger than normal concentrations of nitrates (NO_3^-) and chlorine from salt in water indicated pollution. But what was "normal"? Testing was needed. Over 40,000 water samples from rivers and lakes all over Massachusetts were tested for chlorine levels and these were recorded on a map, the "Normal Chlorine Map" (Figure 7). Natural levels are 6 to 7 ppm close to the coastline and drop to about 1 ppm further inland.

Her methodologies of water-sampling rely on physical separation of a mixture. Water samples are collected from below the surface into identically sized bottles which are rinsed with the water being sampled. Samples are analysed within 24 hours. Silver nitrate solution is added in excess to produce silver chloride, AgCl, precipitate. The AgCl is separated by filtration, dried and weighed and the concentration of chloride ions in the original sample can be determined.

Richards' experiments in water quality led to the filtration of water supplies for towns and cities. Bacteria had recently been discovered, and Richards researched filtration of water through sand, to reduce bacterial contamination, testing the effect of rate of filtration and size of the sand grains.

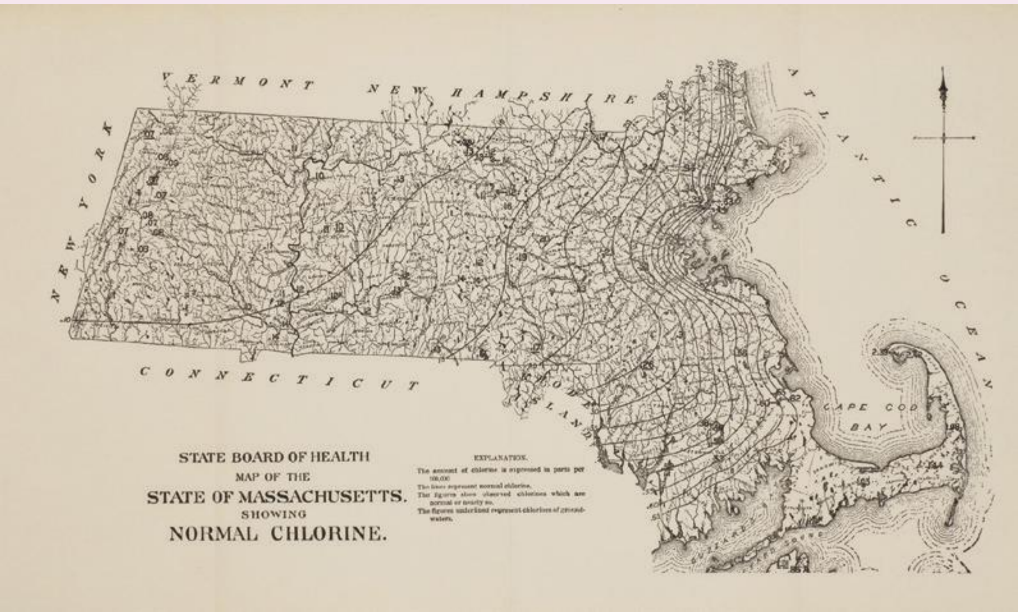


FIGURE 7 Richards' "Normal Chlorine Map" was published in her book *Air, Water, and Food from a Sanitary Standpoint*.

Apply your understanding

1 **Propose** a reason why water samples needed to be collected and tested in the same way. (1 mark)

2 **Consider** the two variables in sand filtration. **Construct** justified hypotheses that Richards may have tested. (3 marks)

Check your learning 5.2

Check your learning 5.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** what a homogeneous mixture is. (1 mark)

2 **Identify** two types of heterogeneous mixtures. (2 marks)

3 **Explain** why a bowl of fruit-and-nut muesli would not be considered a homogeneous mixture. (2 marks)

4 **Explain** why it is important to know the physical properties of substances when trying to separate a mixture. (2 marks)

Analytical processes

5 **Discriminate** between homogeneous and heterogeneous mixtures. Give examples to support your answer. (4 marks)

6 **Classify** each of the mixtures below as homogeneous or heterogeneous.

- **a** Clean tap water (1 mark)
- **b** Full cream milk (1 mark)
- **c** Water from a river (1 mark)
- **d** A glass of soft drink with bubbles in it (1 mark)
- **e** Hydrochloric acid (1 mark)
- **f** Smoke in air (1 mark)

Knowledge utilisation

7 **Justify** why clean air is considered a homogeneous mixture. (2 marks)

8 A student has a mixture of sand and vegetable oil. **Design** a method that would allow the student to easily separate the two substances. (5 marks)

9 **Design** a method to separate a mixture of water, sand and table salt. (5 marks)

Practical

Lesson 5.3

Separating mixtures

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Video demonstration

MODULE 5

Lesson 5.4

Review: Compounds and mixtures

Summary

5.1
- Pure substances are made up of elements or compounds.
- Pure substances have distinct, measurable properties, including melting point, boiling point, reactivity, bond strength and density.

5.2
- Properties of mixtures depend on the identity and relative amounts of their components.
- Heterogeneous mixtures are uniform, whereas homogeneous mixtures do not have a consistent composition.
- Mixtures can be separated based on the physical properties of the components.

5.3
- Practical: Separating mixtures

Key formulas

Density $\text{Density} = \frac{\text{mass}}{\text{volume}}$

Review questions 5.4A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Pure water is best described as
A an element.
B heterogeneous.
C a compound.
D a solution.

2 The melting point of a pure substance
A is typically very high.
B is typically very low.
C generally falls within a narrow range.
D generally falls within a wide range.

3 A 38.6 g sample of pure gold has a volume of 2.00 cm^3. The density of gold is
A 77.2 g cm^{-3}.
B 38.6 g cm^{-3}.
C 0.05 g cm^{-3}.
D 19.3 g cm^{-3}.

4 A beaker of salty water contains crystals that will not dissolve. This mixture is best described as
A pure.
B dilute.
C unsaturated.
D heterogeneous.

5 Which of the following is an example of a homogeneous mixture?
A Concrete
B Gold
C Methanol
D Salt water

6 A sample of clean air is best described as
A an element.
B heterogeneous.
C homogeneous.
D a solution.

7 The density of aluminium is $2.71\ g\ cm^{-3}$. If a sample of aluminium has a mass of 14.8 g, the volume of the sample is

A 0.183 cm^3.

B 40.1 cm^3.

C 5.46 cm^3.

D 12.1 cm^3.

8 During a chemical reaction

A reactant bonds and product bonds are broken.

B reactant bonds and product bonds are formed.

C reactant bonds are broken and product bonds are formed.

D reactant bonds are formed and product bonds are broken.

9 Which of the following separation methods could be used to separate components of a homogeneous mixture?

A Flotation

B Sedimentation

C Using a separating funnel

D Distillation

10 A heterogeneous mixture consists of two components with similar boiling points but very different densities. Which of the following methods would best separate the components?

A Distillation

B Using a separating funnel

C Crystallisation by evaporation

D Filtration

Review questions 5.4B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

11 **Explain** what pure substances are, using examples. (2 marks)

12 **Explain** what dictates the properties of a mixture. (2 marks)

13 **Define** the term “density”. (1 mark)

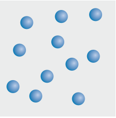
14 **Describe** what happens to a substance at its boiling point. (1 mark)

Analytical processes

15 **Discriminate** between homogeneous and heterogeneous mixtures. (1 mark)

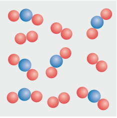
16 **Classify** the following diagrams as representing pure substances, homogeneous mixtures or heterogeneous mixtures.

a



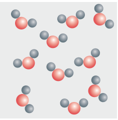
(1 mark)

b



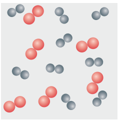
(1 mark)

c



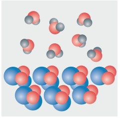
(1 mark)

d



(1 mark)

e



(1 mark)

17 **Classify** the following mixtures as homogeneous or heterogeneous.

a A fruit salad (1 mark)

b A glass of diluted cordial (1 mark)

c A milkshake (1 mark)

d Smoke from a fire (1 mark)

e A chocolate chip muffin (1 mark)

18 **Analyse** the melting and boiling temperatures in Table 1 to **determine** the physical state of each substance at –200°C. (5 marks)

TABLE 1 Physical properties of some substances under standard laboratory conditions

Element	Physical state (at 25°C and 1 bar)	Melting point (°C)	Boiling point (°C)
Water (H_2O)	liquid	0	100
Oxygen	gas	−218.79	−182.962
Nitrogen	gas	−210	−195.975
Mercury	liquid	−38.8290	356.73
Iron	solid	1,538	2,862

Knowledge utilisation

19 The density of water varies with temperature as shown in Figure 1.

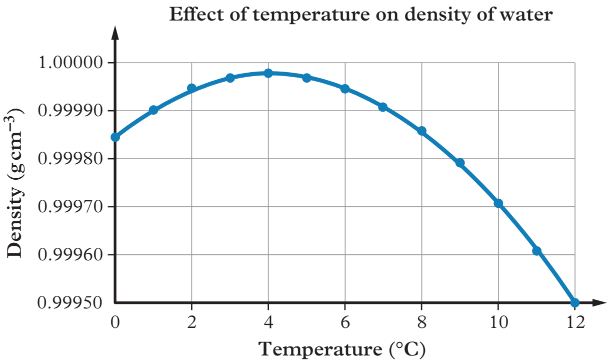


FIGURE 1 Density of water over different temperatures

a **Identify** the temperature at which water has the greatest density. (1 mark)

b Assuming that no water is lost through evaporation, **describe** what would happen to a given volume of water as the temperature increases from 6°C to 12°C. (1 mark)

c 1.0000 L of water has a mass of 999.5 g. **Determine** the temperature of the water. (1 mark)

d A sample of water has a density of 0.99990 $g\,cm^{-3}$. From this information alone, **consider** whether you can determine the temperature of the water. **Justify** your answer. (2 marks)

20 **Design** a method to separate a mixture of sand and sodium chloride (salt) and **justify** your design, based on differing physical properties. (3 marks)

21 Combustion engines have an air filter.

a **Explain** the role of the air filter in a combustion engine by considering the physical properties of the surrounding air. (2 marks)

b **Propose** a reason why this filter must be changed at regular intervals. (1 mark)

Data drill

Densities of common substances

Density is the measure of an object's mass compared with the volume that it takes up. It also helps to explain why some things will float on water, while other things will sink. The masses of different volumes of some common substances were measured, and graphed in Figure 1.

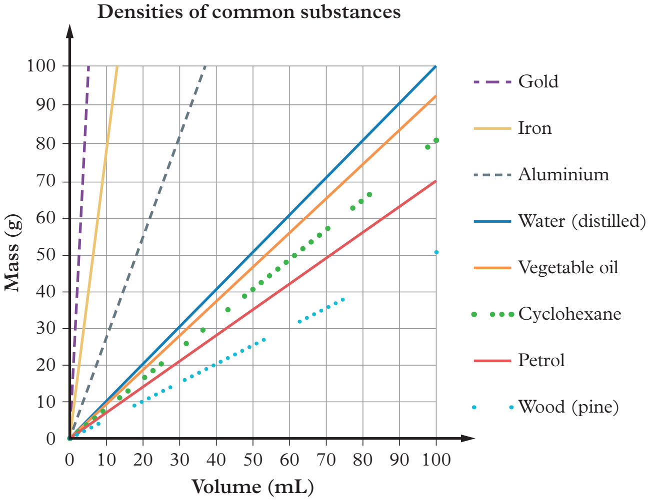


FIGURE 1 Densities of common substances

Apply understanding

1 **Identify** the substance in the graph which has the highest density. (1 mark)

2 **Calculate** the densities of aluminium, wood (pine) and water. (3 marks)

Analyse evidence

3 **Contrast** the densities of the metals – gold, iron and aluminium, with those of the other substances tested. (2 marks)

Interpret evidence

4 **Determine** which of the substances from the graph will float on water and which will not. **Justify** your answer using the calculations from question **2**. (4 marks)

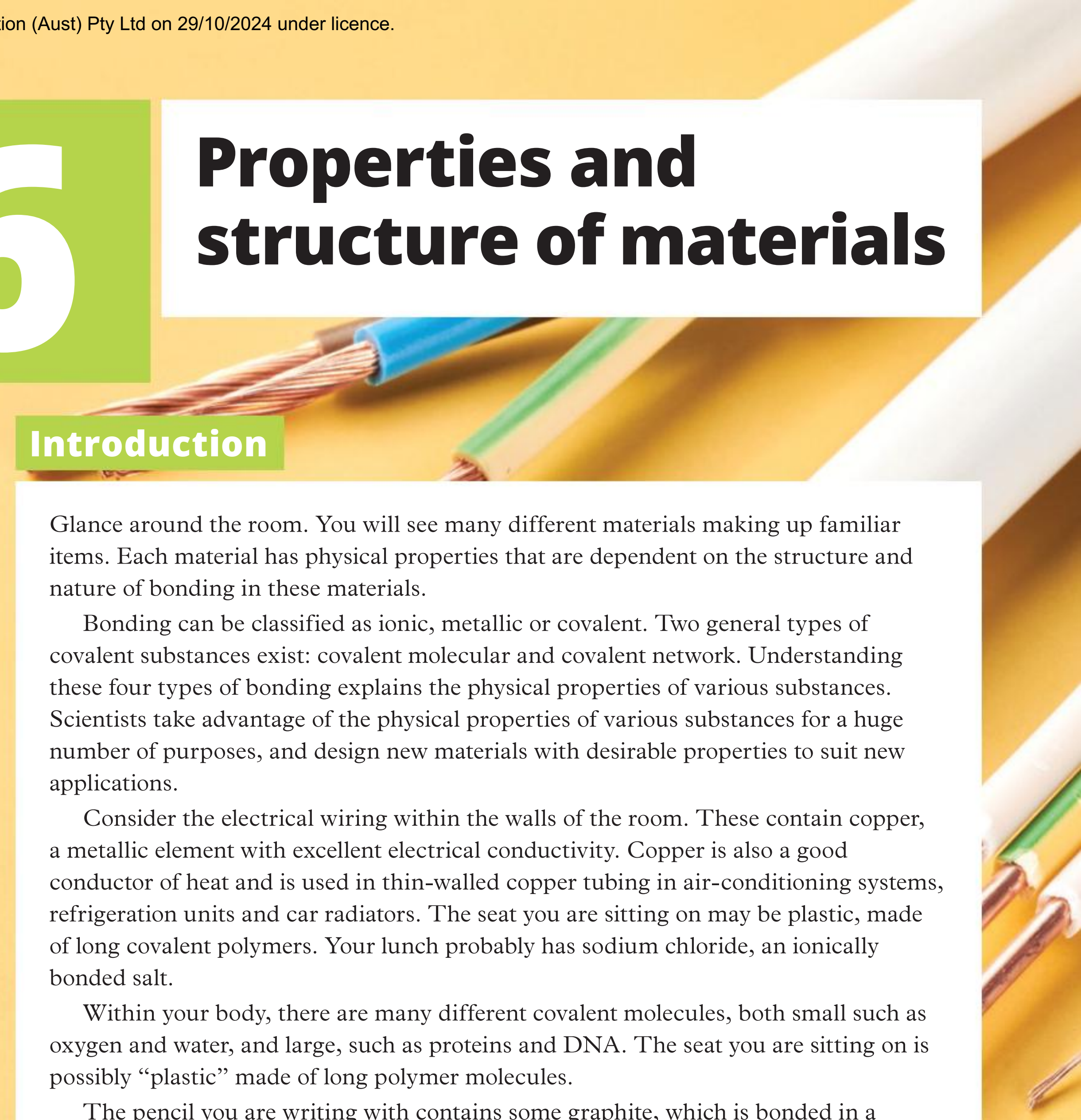
5 Petrol and cyclohexane are mixed together to form a homogeneous mixture. Water is then added to this mixture. Petrol and cyclohexane do not mix well with water. The mixture is then placed in a separating funnel. **Predict** the position and composition of the layers of liquid in the separating funnel by sketching a labelled diagram. (2 marks)

Module 5 checklist: Compounds and mixtures

Quizlet: Revise key terms online to test your understanding

MODULE 6

Properties and structure of materials



Introduction

Glance around the room. You will see many different materials making up familiar items. Each material has physical properties that are dependent on the structure and nature of bonding in these materials.

Bonding can be classified as ionic, metallic or covalent. Two general types of covalent substances exist: covalent molecular and covalent network. Understanding these four types of bonding explains the physical properties of various substances. Scientists take advantage of the physical properties of various substances for a huge number of purposes, and design new materials with desirable properties to suit new applications.

Consider the electrical wiring within the walls of the room. These contain copper, a metallic element with excellent electrical conductivity. Copper is also a good conductor of heat and is used in thin-walled copper tubing in air-conditioning systems, refrigeration units and car radiators. The seat you are sitting on may be plastic, made of long covalent polymers. Your lunch probably has sodium chloride, an ionically bonded salt.

Within your body, there are many different covalent molecules, both small such as oxygen and water, and large, such as proteins and DNA. The seat you are sitting on is possibly “plastic” made of long polymer molecules.

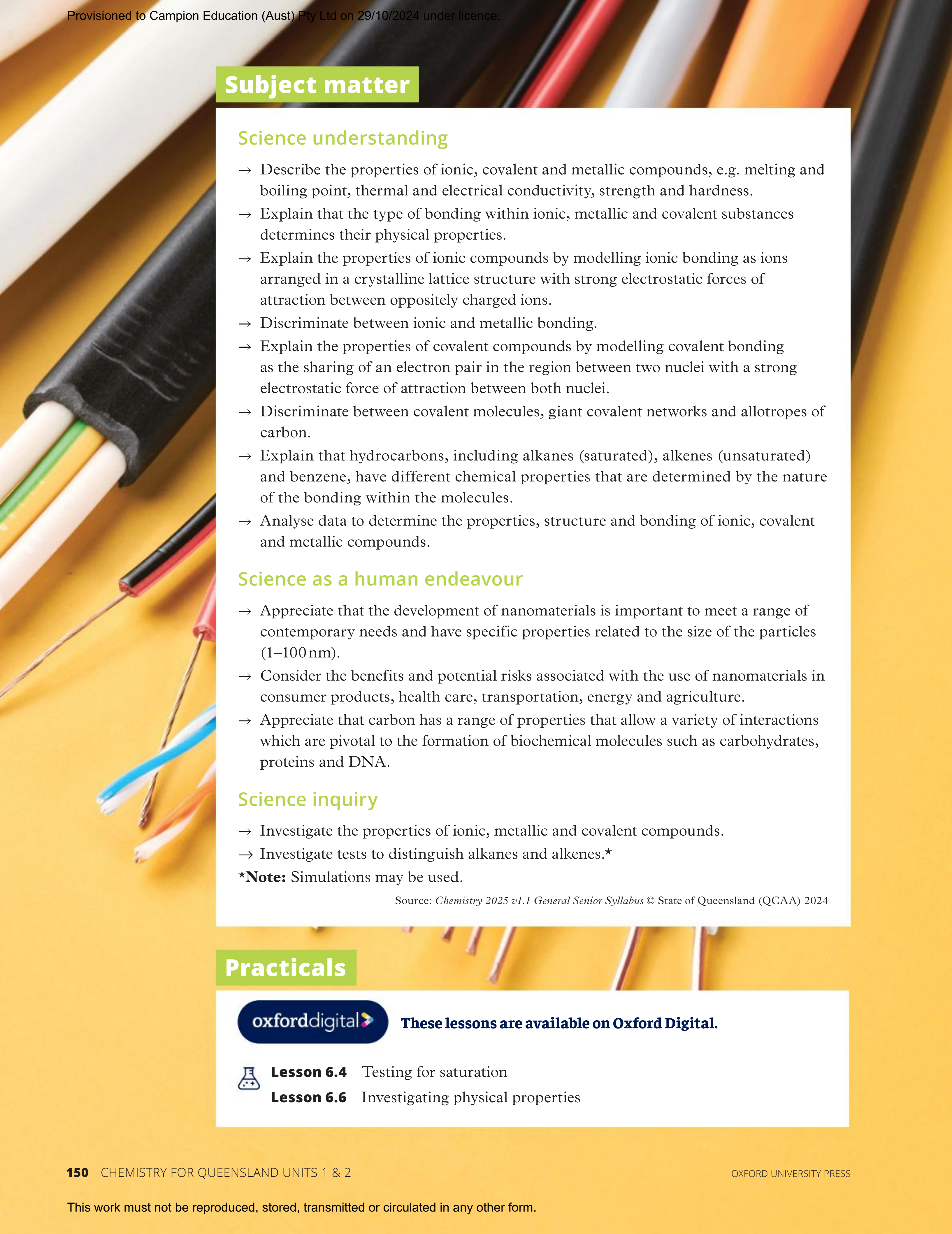
The pencil you are writing with contains some graphite, which is bonded in a giant covalent network. Scientists have developed special glass for the touch screen of electronic devices, which also consists of a covalent network, taking advantage of their knowledge of structure, bonding and properties.

This module explores the type of bonding found in ionic compounds, metallic bonding, and the two types of covalently bonded substances, and introduces the intermolecular forces that occur between molecules. An understanding of chemical bonding and intermolecular forces is needed to know how substances will react and how they can be transformed into new and useful materials.

Prior knowledge

Prior knowledge quiz

Check your understanding of ionic, metallic and covalent bonding before you start.



Subject matter

Science understanding

- → Describe the properties of ionic, covalent and metallic compounds, e.g. melting and boiling point, thermal and electrical conductivity, strength and hardness.
- → Explain that the type of bonding within ionic, metallic and covalent substances determines their physical properties.
- → Explain the properties of ionic compounds by modelling ionic bonding as ions arranged in a crystalline lattice structure with strong electrostatic forces of attraction between oppositely charged ions.
- → Discriminate between ionic and metallic bonding.
- → Explain the properties of covalent compounds by modelling covalent bonding as the sharing of an electron pair in the region between two nuclei with a strong electrostatic force of attraction between both nuclei.
- → Discriminate between covalent molecules, giant covalent networks and allotropes of carbon.
- → Explain 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 data to determine the properties, structure and bonding of ionic, covalent and metallic compounds.

Science as a human endeavour

- → Appreciate that the development of nanomaterials is important to meet a range of contemporary needs and have specific properties related to the size of the particles (1–100 nm).
- → Consider the benefits and potential risks associated with the use of nanomaterials in consumer products, health care, transportation, energy and agriculture.
- → Appreciate that carbon has a range of properties that allow a variety of interactions which are pivotal to the formation of biochemical molecules such as carbohydrates, proteins and DNA.

Science inquiry

- → Investigate the properties of ionic, metallic and covalent compounds.
- → Investigate tests to distinguish alkanes and alkenes.*

*__Note:__ Simulations may be used.

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Practicals

These lessons are available on Oxford Digital.

Lesson 6.4 Testing for saturation

Lesson 6.6 Investigating physical properties

Lesson 6.1

Properties of ionic compounds

Key ideas

→ Ionic compounds are arranged in a crystalline lattice structure.

→ The main characteristics of ionic compounds are related to their ionic structure and bonding.

→ Ionic compounds typically have high melting and boiling points, are brittle and conduct electricity in molten liquid or aqueous states.

What do you remember about ionic compounds?

Ionic substances are always compounds and not elements. Remember from Module 3 that they consist of positively charged and negatively charged ions. In most cases, the positively charged ion is a metal ion, as they lose valence electrons more readily to form cations. A notable exception is the ammonium ion NH_4^+. The negative ion can be a non-metal like chloride (Cl^-), a polyatomic ion made of non-metals like the sulfate ion SO_4^{2-}, or a polyatomic ion that contains a metal and non-metal like the permanganate ion MnO_4^{2-}. Non-metals gain electrons more readily to form anions.

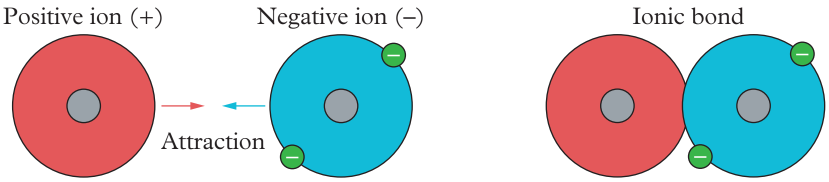


FIGURE 1 Ionic bonds form between positive and negative ions.

What is the structure of ionic compounds?

The ionic model of bonding describes how cations and anions are held together in ionic compounds. An ionic compound is an ordered, repeating three-dimensional arrangement of cations and anions in a **crystalline lattice** structure (Figure 2). In a crystalline lattice, cations are surrounded by anions and vice versa. Ionic bonding is the strong electrostatic attraction between each ion and all the neighbouring ions of the opposite charge.

In sodium chloride (NaCl, Figure 2), each interior sodium ion (Na^+) is surrounded by and ionically bonded to multiple chloride ions (Cl^-), while each chloride ion is surrounded by and ionically bonded to multiple sodium ions. This arrangement maximises the attraction between the oppositely charged ions. The way that ions pack together depends on their charges and ionic radii:

- Ions with smaller ionic radii or size have greater attractions between the ions than an ionic compound with larger ions with the same charges, as long as the ions are arranged in the same way.
- Ions of the same charge repel each other. Ions of opposite charge are attracted to each other.

Learning intentions and success criteria

Study tip

Ionic bonding was introduced in Module 3. It is a good idea to revise this material before advancing.

Study tip

It will be helpful to review and memorise the polyatomic ions and their charges, especially when you come to Units 3 & 4. Make sure you are aware of the metals that can form two ions with different charges, how to name their compounds, and how to determine the charge of the ion from the formula. One way to learn the names and formulas is to put them on either side of flash cards or use a flash card app.

crystalline lattice
the ordered, repeating three-dimensional arrangement of atoms, molecules or ions in crystalline solids

In a few compounds, water molecules are also incorporated into the lattice of ions in a specific ratio. For example, copper(II) sulfate has five water molecules in its lattice for every one unit of copper sulfate. This gives it the formula $CuSO_4{\cdot}5H_2O$.

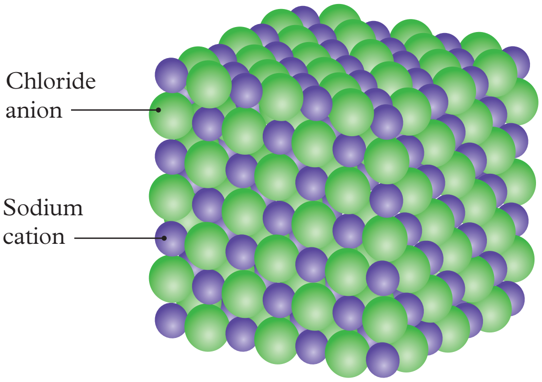


FIGURE 2 Ionic compounds, such as sodium chloride (NaCl), are arranged in a crystalline lattice structure.

FIGURE 3 Sodium chloride (table salt) has a melting point of 801°C and a boiling point of 1,413°C.

Study tip

Visualising the structure of substances and knowing about the strength of the bonds will help you to explain the various properties. This approach works for all types of substances, not just ionic substances. Drawing small diagrams can help you visualise.

How does the structure of ionic compounds explain their physical properties?

The physical properties of ionic compounds can be explained by considering the nature of ionic bonding and the structures they form.

Melting and boiling points

The ions in an ionic compound are held in place in the crystal lattice by the electrostatic attraction between oppositely charged ions. It takes a large amount of heat energy to overcome these strong ionic bonds, giving rise to high melting points.

To convert an ionic compound from a liquid to a gas at its boiling point, it is necessary to overcome all of the electrostatic interactions between oppositely charged ions. Ionic compounds therefore have relatively high melting and boiling points, usually hundreds to thousands of degrees Celsius.

Electrical conductivity

For a substance to conduct electricity, charged particles need to be able to move and carry current between the electrodes (Figure 4). In the solid state, ions are held in a rigid lattice and cannot move freely to carry electrical charge. There are no free electrons to carry charge either. Therefore, ionic compounds do not conduct electricity in the solid state.

However, they can conduct electricity as molten liquids. In the molten state, ions are no longer held in a rigid lattice and are free to move between the electrodes in a circuit. The current in the liquid is due to moving ions, not electrons.

Aqueous solutions of ionic solids can also conduct electricity. When dissolved in water, the ions separate. Similar to the ions in molten liquids, the ions can move towards an electrode with the opposite charge. Solutions and molten liquids that conduct electricity are termed “electrolytes”.

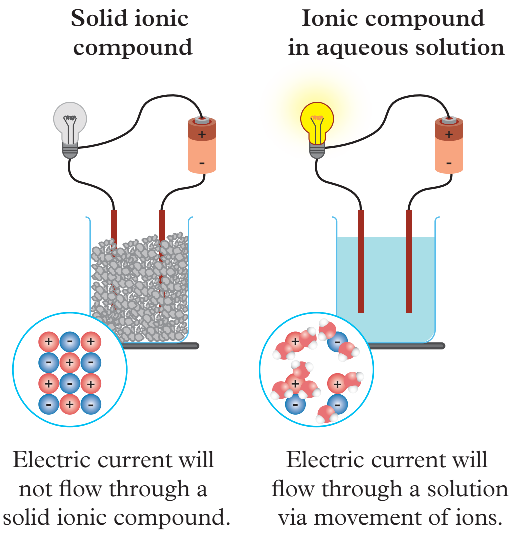


FIGURE 4 Ionic compounds will not conduct electricity as a solid since ions cannot flow freely. However, as a molten liquid or in aqueous solution, the ions are free to flow towards the electrodes. Anions move towards the positively charged electrode (anode), cations move towards the negatively charged electrode (cathode). Electrons carry current through the wire circuit to light the bulb.

Hardness and brittleness

Hardness is the ability of a solid to resist being scratched by other solids. Ionic bonding is strong because of the electrostatic attraction between oppositely charged ions throughout the entire crystal lattice. This makes it difficult to scratch. Therefore, ionic solids are relatively hard.

Despite this, ionic solids are brittle. Under pressure, such as when hit with a hammer, the crystal lattice structure becomes disrupted. Ions of the same charge are brought close together, resulting in repulsion (Figure 5). The ionic crystals shatter in a regular manner along smooth planes.

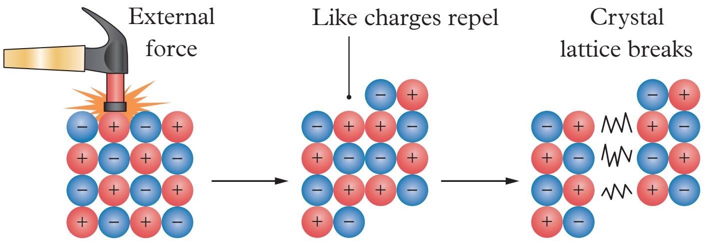


FIGURE 5 Ionic compounds are brittle solids because of their ionic bonds.

Solubility

Many (but not all) ionic solids are soluble in water. To dissolve in water, the strong ionic bonds holding the ions together must be overcome. Solubility depends on the amount of energy required to do this; it is also the energy released by interactions between water molecules and ions. For example, $Mg(NO_3)_2$ is soluble in water, whereas MgO is insoluble.

Water molecules also partially shield the ionic charges from each other, reducing the attraction between oppositely charged ions. This happens because water molecules are slightly negative near the oxygen atom and slightly positive near the hydrogen atoms. The oxygen end of the water molecule is attracted to cations, pulling them away from the crystal lattice and surrounding them in solution. Likewise, the hydrogen end is attracted to anions, and each anion is surround by several water molecules.

Study tip

You will learn more about solubility in Module 12.

Anions and cations already exist in charged form within a solid ionic compound, so dissolving simply causes the ions to dissociate and move away from each other.

Worked example 6.1A

Explaining properties of ionic compounds

▶ **Worked example 6.1A:** Watch a video that shows how to solve this problem.

Explain why solid NaCl does not conduct electricity. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Explain" means to make an idea or situation plain or clear by describing it in more detail or revealing relevant facts. We need to identify the type of substance, recognise its structure, and relate this to the property. The question is worth 4 marks, so we must analyse the substance and explain why it has the specified property.
Step 2: Identify the type of substance.	Sodium chloride is an ionic compound. (1 mark)
Step 3: Describe the structure and bonding present in the substance. A small 2D sketch can help.	NaCl is made of Na^+ and Cl^- ions, held in a crystalline lattice by electrostatic attractions between oppositely charged ions. (1 mark)
Step 4: Link the structure and bonding to the property, in this case, not conducting.	To conduct electricity, charged particles (electrons or ions) must be free to move and carry current. (1 mark) In solid NaCl, the ions are held in fixed positions, unable to move. There are no free electrons to carry current. (1 mark)

Your turn

Explain why a solution of copper(II) sulfate conducts electricity. (4 marks)

Skill drill

Distinguishing between scientific and non-scientific ideas about ionic compounds

Science inquiry skills: Evaluating evidence (Lesson 1.8)

Three students are researching the properties of ionic compounds for an assignment. Each student has suggested what sort of information they should include in their presentation.

→ Student 1 suggests that they should include information on how ionic compounds are brittle because they once saw a large salt crystal break into pieces after it fell to the floor.

→ Student 2 suggests that they should include information on how ionic compounds have high melting points because data from several peer-reviewed journal articles they have read supports this concept.

→ Student 3 suggests that they should include information on how all ionic compounds can conduct electricity because in a classroom experiment, they were able to keep a circuit connected using an ionic compound in aqueous solution.

Practise your skills

1 **Identify** and **explain** whether each student's suggestion is based on opinion, anecdote, or evidence. (6 marks)

2 **Identify** and **explain** whether each suggestion presents a valid scientific idea or non-scientific idea. (6 marks)

Check your learning 6.1

Check your learning 6.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Identify** the formulas for ionic compounds formed from

- a caesium and iodine (1 mark)
- b aluminium and fluorine (1 mark)
- c gallium and oxygen (1 mark)
- d lithium and nitrogen (1 mark)
- e sulfur and sodium (1 mark)
- f bromine and calcium. (1 mark)

2 **Identify** the correct name for each of these ionic compounds

- a Na_3N (1 mark)
- b KNO_3 (1 mark)
- c MgS (1 mark)
- d Li_2CO_3 (1 mark)
- e $Fe(OH)_2$ (1 mark)
- f $Fe(NO_3)_3$. (1 mark)

3 Ionic compounds have a range of properties in common.

- a **Recall** two typical properties of ionic compounds. (2 marks)
- b **Describe** what each property is. (2 marks)

4 **Sketch** a labelled 2D diagram to show how positive and negative ions are arranged in ionic compounds. (2 marks)

5 **Explain** why crystals of ionic compounds are brittle and shatter easily. (3 marks)

6 **Explain** why ionic compounds have relatively high melting and boiling points. (2 marks)

Analytical processes

7 **Infer** two trends based on the melting and boiling point data shown in Table 1. (2 marks)

TABLE 1 The melting and boiling points of selected ionic compounds

Compound	Melting point (°C)	Boiling point (°C)
NaF	993	1,695
NaCl	801	1,413
NaBr	747	1,396
KF	770	1,420
CsF	682	1,251
MgO	2,852	3,600

Lesson 6.2

Properties of metallic substances

Key ideas

- → Metallic bonding is explained using the free electron model.
- → The main characteristics of metallic substances are related to their metallic bonds.
- → Metallic substances typically have high electrical conductivity, high boiling points, are malleable and ductile.

What do you remember about metallic substances?

Learning intentions and success criteria

Most elements (about 80%) in the periodic table are metals. They are to the left and middle of the periodic table; for example, iron (Fe), and aluminium (Al). Metals are often available as pure substances.

There are also many alloys, which are mostly mixtures of two or more metals. Alloys have properties similar to pure elemental metals. Ancient alloys include brass and bronze.

In modern times, many more alloys have been developed with specific properties, such as stainless steel which contains iron but does not rust and bisalloy, which contains iron but is extremely hard and durable, and thus used in heavy equipment in the mining industry. A few alloys incorporate a non-metallic element into the metallic structure, like cast iron which contains 2% to 4% carbon and mild steel which contains traces of carbon and other metals.

The formula for pure metals is simply the elemental symbol of the metal. As alloys are mixtures, it is not possible to write a formula for them. However, their composition may be expressed in terms of the percentages of each component.

FIGURE 1 Metallic substances include pure metals (like iron) and alloys (like the stainless steel used to make the mixing bowl).

What is the structure of metallic substances?

Like ionic compounds, metals form a crystalline lattice. For metallic substances, this is explained using the **free electron model of metallic bonding**. The model considers metals in the solid state as lattices of metal ions surrounded by a sea of valence electrons (Figure 2). Metallic elements can lose valence electrons relatively easily because they have low ionisation energies. These electrons are free to move randomly in the solid and are not associated with any single metal ion but are shared between several neighbouring metal ions. These electrons are called **delocalised electrons**.

The metal ions can vibrate but are otherwise fixed in a repeating three-dimensional lattice, with the metal ions packed closely together. The forces of attraction between positively charged metal ions and negatively charged electrons are the **metallic bonds** keeping the solid together. Metallic bonding is quite strong, similar in strength to ionic bonding.

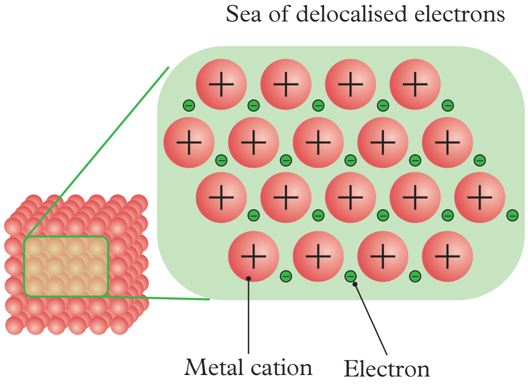


FIGURE 2 Metallic bonding occurs where delocalised electrons surround metal ions in a lattice.

free electron model of metallic bonding
a model of the bonding in metals that presents metals as ions surrounded by one or more valence electrons that are not associated with a particular atomic nucleus

delocalised electron
an electron that is not retained in bonds between two atoms or associated with a single nucleus of an atom but is instead spread over several atoms or is highly mobile within a solid

metallic bond
the attractive force between a metal ion and delocalised electrons that hold metals together

How does the structure of metallic substances explain their physical properties?

The physical properties of metallic substances can be explained using the free electron model of metallic bonding.

Melting and boiling points

Metallic substances have high melting and boiling points. The strength of the metallic bonds depends on the number of delocalised electrons, and the charge and size of the cation. In a molten metal, the ordered arrangement no longer exists, but the attractive forces are not completely overcome until enough heat energy is added to reach the boiling point. Therefore, this is a better indicator of the strength of the metallic bonds than the melting point.

Table 1 shows the melting and boiling points for selected alkali and alkali earth elements. These values are relatively high, meaning that metallic substances are solid at room temperature. Exceptions include mercury, which has a melting point of –38.83°C, and on a particularly hot day, gallium, which as a melting point of 29.8°C.

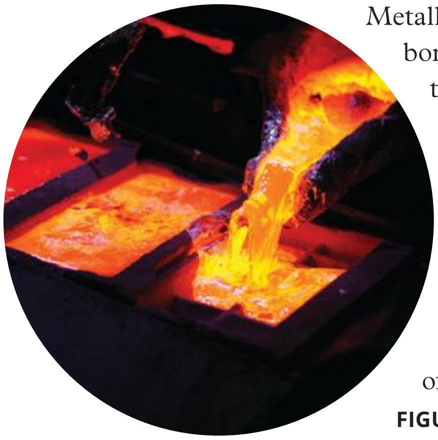


FIGURE 3 Very high temperatures are required to melt metallic substances.

TABLE 1 Melting and boiling points of selected alkali and alkali earth elements

Alkali metal	Melting point (°C)	Boiling point (°C)	Alkali earth metal	Melting point (°C)	Boiling point (°C)
Lithium (Li)	181	1,330	Beryllium (Be)	1,287	2,470
Sodium (Na)	98	883	Magnesium (Mg)	650	1,090
Potassium (K)	63	759	Calcium (Ca)	842	1,484
Rubidium (Rb)	39	688	Strontium (Sr)	777	1,382
Caesium (Cs)	28	671	Barium (Ba)	727	1,870

Electrical conductivity

Metallic substances have high electrical conductivity. If a piece of metal such as a wire is connected to an electrical source such as a battery and then connected into an electrical circuit, electrons from the battery will enter the wire and displace electrons from the end of the wire closest to the source of the current. Electrons that are part of the delocalised sea of electrons will flow through the solid, displacing electrons from the other end of the wire.

Thermal conductivity

Solid metals have high thermal conductivity, meaning that they can transfer heat easily. This is also due to the free electrons. Heat is transferred in metals as electrons move freely from atom to atom, distributing the energy supplied as heat through the solid. Some evidence for this is that the thermal conductivity of metals is proportional to how electrically conductive they are and both depend on delocalised electrons.

When you touch a piece of hot metal, relatively large amounts of heat from various parts of the metal can be easily transferred from some distance into your finger. Hot metal feels much hotter and is more likely to burn you than a less conductive substance of the same temperature.

Hardness

Many metals are relatively hard, although a few are softer. This is due to the strong force of attraction between positively charge ions in the lattice and the negatively charged delocalised or free electrons.

Malleability and ductility

Metallic substances can easily be shaped into sheets and drawn into wires. In other words, they are malleable and ductile. When a metal is beaten with a hammer, layers of ions can slip past each other (Figure 4). After the movement, the environment around each metal atom is unchanged because the sea of electrons can easily adapt to the change due to their delocalised nature.

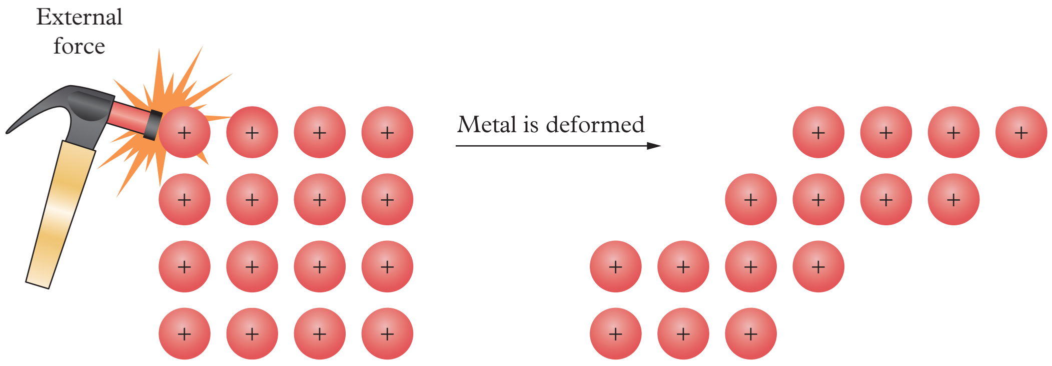


FIGURE 4 When force is applied to metal, layers of metal ions move and the delocalised electrons adapt to the change, resulting in deformation.

FIGURE 5 The metallic lustre of metals is part of its attraction for metals such as silver and gold in jewellery.

Lustre

Metals are typically shiny and lustrous, often silvery. The sea of delocalised electrons reflects light well, causing a fresh metallic surface to appear shiny. If the surface is oxidised or corroded, it will not be shiny.

Solubility

Due to the strong forces of attraction between the lattice of ions, and the delocalised sea of electrons, metals are insoluble in water. The positive metal ions cannot be dissolved and hydrated, unlike the positive and negative ions in ionic substances.

Water does assist in creating an environment where some metals will react chemically and corrode. A few metals will even react violently when placed in water. However, these are chemical reactions rather than the physical process of dissolving.

What is the difference between ionic and metallic bonding?

Ionic and metallic bonding have some similarities and some major differences. Both types of bonds arise from the strong electrostatic attractions between oppositely charged particles within a crystalline lattice. However, ionic bonding occurs between positive ions and negative ions held in fixed positions in the lattice. Metallic bonding occurs between positive ions and free electrons. The positive ions are held in fixed positions in the lattice, but the outer electrons of each metal are delocalised, and free to move anywhere in the lattice.

Check your learning 6.2

Check your learning 6.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Recall** the symbols for these common metals and the ions they typically form.
  - a Zinc (1 mark)
  - b Silver (1 mark)
  - c Copper (two possible ions) (1 mark)
  - d Magnesium (1 mark)
  - e Iron (two possible ions). (1 mark)

2 **Explain** the following properties of metals by considering the bonding.
  - a Metals conduct electricity very well when solid or molten. (3 marks)
  - b Metals generally have relatively high melting and boiling points. (3 marks)
  - c Metals are lustrous. (2 marks)

Analytical processes

3 Metallic substances and ionic substances have some similarities but some key differences. Both consist of ions arranged in a lattice in the solid state. The bonding is due to electrostatic attraction between oppositely charged particles.
  - a **Construct** labelled diagrams that describe the differences and similarities between ionic and metallic bonding. (2 marks)
  - b **Compare** the properties of ionic compounds and metallic substances using a Venn diagram. Properties that are shared by both types of substance appear in the section of overlap. The first property has been done for you. (6 marks)

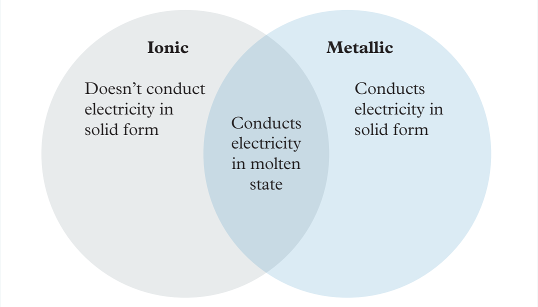


FIGURE 6 Similarities and differences of metallic and ionic substances

4 Some properties of two metals and two ionic compounds are listed in Table 2.

TABLE 2 Properties of magnesium, magnesium oxide, silver and silver chloride

Substance	Melting point (°C)	Boiling point (°C)	Colour	Electrical conductivity (S cm^{-1})
Mg	650	1,090	silver	2.3×10^7
MgO	2,852	3,600	white powder	10^{-14} to 10^{-12}
Ag	962	2,162	silver	6.2×10^7
AgCl (anhydrous form)	455	1,547	white powder	10^{-9} to 10^{-6}

- **a** **Compare** the properties of the two metals with the properties of the two ionic compounds. (2 marks)
- **b** **Consider** your answer to part **a** and the bonding involved to **discriminate** between metallic and ionic substances. (4 marks)

5 **Analyse** the following *incorrect* statements about metallic bonding. **Identify** errors in these statements and rewrite them to be correct.

- **a** Metals can conduct electricity in the solid state because the metal atoms can easily slip past each other to carry electrical charge through the solid. (2 marks)
- **b** Metals are held together by ionic bonds. (2 marks)
- **c** The strength of metallic bonding in different metals is constant because only a single valence electron from each atom becomes delocalised. (2 marks)
- **d** Metals melt once sufficient heat energy has been provided to boil away the sea of electrons. (2 marks)
- **e** The melting point of a metal is a good indication of the strength of the metallic bonding within that metal. (2 marks)

Lesson 6.3

Properties of simple covalent molecules

Key ideas

→ The main physical characteristics of covalent substances are related to the relatively weak intermolecular forces that hold them together.

→ Simple covalent molecules have low melting and boiling points and do not conduct electricity.

→ The properties of hydrocarbons relates to the type of carbon–carbon bonds present.

What do you remember about covalent molecules?

Learning intentions and success criteria

As discussed in Module 3, covalent bonding occurs when electrons are shared between atoms. The shared electrons are electrostatically attracted to the nuclei of both atoms involved in the bond.

Non-metallic elements in the periodic table form covalent bonds with themselves and each other. These elements are positioned to the right-hand side of the periodic table and have greater valency than metals. Non-metals cannot gain electrons from other non-metals, so they share electrons when bonded.

FIGURE 1 Carbon dioxide

Nearly all non-metallic elements and compounds are in the form of covalent molecules. Covalent molecules can be represented by a chemical formula which shows the elements and number of atoms of each element.

Simple molecules that have covalent bonds include elements such as oxygen (O_2), nitrogen (N_2), and compounds such as water (H_2O), methane (CH_4), ammonia (NH_3), and carbon dioxide (CO_2). Lesson 6.5 discusses those few non-metallic elements and compounds that form giant covalent network substances instead.

The noble gases, consisting of single atoms, have physical properties similar to covalent molecular substances, so are often grouped together in discussion. Other larger covalent molecules exist, such as the proteins in all living things, and the many different natural and synthetic polymers you will learn about in Unit 4.

What is the structure of covalent molecules?

intramolecular force
an attractive force between the atoms in a molecule; *intra* = "within"

Covalent molecules consist of discrete molecules with very strong covalent bonding between the atoms in the molecule. These are **intramolecular forces** (i.e. between the atoms in a single molecule). Intramolecular forces are extremely strong and require a lot of energy to break. For example, to break the covalent O–H bonds in water requires heating it to thousands of degrees, or blasting the molecules with high-energy ultraviolet (UV) radiation. The strength of bonds increases from single to double to triple bonds, as the number of shared electrons increases.

intermolecular force
a force between molecules; *inter* = "between"

In a solid or a liquid made up of many covalent molecules, there are no covalent bonds between distinct molecules. Instead, molecules are attracted weakly to each other through **intermolecular forces**. Intermolecular forces are explained fully in Module 9, including dispersion forces, dipole–dipole attractions and hydrogen bonding.

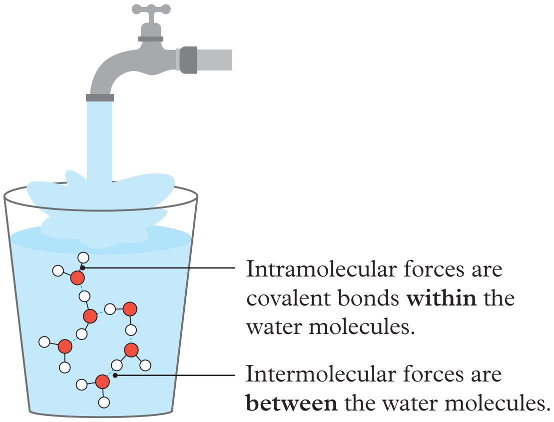


FIGURE 2 Covalent molecules have strong intramolecular forces within the molecule and weak intermolecular forces between molecules.

How does the structure of covalent molecules explain their physical properties?

Just as we have seen with ionic compounds and metallic substances, the physical properties of covalent molecules are strongly tied to their structures. To explain these properties, a distinction needs to be drawn between the forces that keep atoms in a molecule together and those responsible for the attractions between different molecules. In general, the physical properties are due to the much weaker intermolecular forces.

Melting and boiling points

Covalently bonded molecules typically have low melting and boiling points. The smaller molecules tend to be gases or liquids at room temperature. Larger molecules may form solids but with melting points that are typically less than a couple of hundred degrees Celsius.

When a molecular solid is heated, the thermal energy disrupts some of the intermolecular forces between different molecules and the solid can go from a solid to a liquid state, and eventually form a gas. These intermolecular forces take much less energy to overcome than covalent bonds and the molecules remain intact during melting and boiling. For example, intermolecular forces between water molecules can be broken simply by heating to temperatures above 100°C at normal pressures.

Intermolecular interactions are also constantly breaking and re-forming in the liquid state, unlike covalent bonds. Differences in the boiling points of different molecular compounds are due to differences in the strength of the intermolecular interactions between the molecules.

Hardness

When solid, covalent molecules may pack into a regular, repeating lattice and form crystals, such as sugar does. This depends on the shape and rigidity of the molecule. Other covalent molecules may solidify in an irregular fashion, and form soft, waxy or greasy solids. This is because they are held together by relatively weak intermolecular forces. The same weak intermolecular forces make it easier to scratch or break a covalent molecular solid.

Thermal and electrical conductivity

Solid covalent molecules poorly conduct heat or electricity as there are no free electrons within the solid to carry electrical charge or heat energy. All the electrons are either held strongly by their atom (as lone pair electrons) or strongly as covalent bonds between atoms (as bonding electrons).

Aqueous solutions of covalent molecules do not usually conduct electricity, but some do. Most molecular substances remain as discrete molecules when dissolved; the solution has no free ions available to carry current. A few covalent molecules react with water molecules in aqueous solution to form ions; then these ions carry electrical current, for example, sulfuric acid molecules in a lead–acid car battery. In Unit 3, you will learn more about acids and the degree to which aqueous solutions of acids react to form ions.

Solubility

Some covalent molecular substances dissolve in water, and others are insoluble in water, but may dissolve in other solvents. This property depends on the types of intermolecular forces, which are explained in detail in Module 9. If a covalent molecule has the same types of intermolecular forces as a particular solvent, it will dissolve in that solvent. For example, sugar dissolves in water because the intermolecular forces are similar. Cooking oil does not dissolve in water because the intermolecular forces are not similar.

What are saturated hydrocarbons?

Hydrocarbons are an important class of covalent molecules. They are composed entirely of carbon and hydrogen atoms. Because of the small electronegativity difference between carbon and hydrogen, all hydrocarbons are non-polar with weak intermolecular forces acting between molecules.

Saturated hydrocarbons contain the maximum number of hydrogen atoms possible for the number of carbon atoms present. These molecules are also called **alkanes**.

Study tip

Go back and practise drawing Lewis structures for covalent molecules, including those with double and triple bonds.

Study tip

Intermolecular forces are explained in detail in Module 9.

FIGURE 3 Butter is an example of a substance composed primarily of covalent molecules. It is a soft, waxy and greasy solid.

hydrocarbon
a covalent compound that only contains carbon and hydrogen atoms

saturated hydrocarbon
a compound of carbon and hydrogen that contains the maximum possible number of hydrogen atoms for the number of carbon atoms present

alkane
a hydrocarbon that contains the maximum number of hydrogen atoms possible for the given number of carbon atoms; for linear or branched alkanes with n carbon atoms, the number of hydrogen atoms is $2n + 2$; for a cycloalkane with one ring and n carbon atoms, the number of hydrogen atoms is $2n$

Alkanes have single bonds between the atoms, with carbon atoms arranged in a linear fashion to form linear alkanes, or branches coming off the main chain to form branched alkanes. Since each carbon atom can form four other bonds, the number of linear and branched compounds based on a carbon skeleton is practically limitless.

Linear and branched alkanes have the same general formula C_nH_{2n+2}, where n is the number of carbon atoms and $2n + 2$ is the number of hydrogen atoms. If $n = 1, 2, 3, 4 \ldots$, then $2n + 2 = 4, 6, 8, 10 \ldots$, to give the molecules CH_4, C_2H_6, C_3H_8, C_4H_{10}.

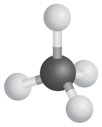
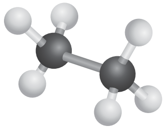
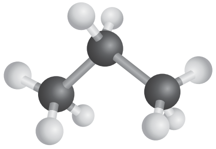
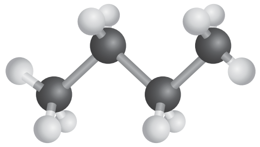
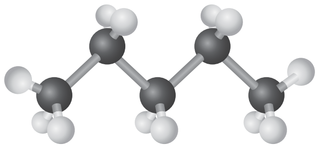
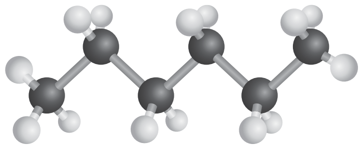
Within the molecules of saturated hydrocarbons, the atoms bonded to each carbon atom are arranged in a roughly tetrahedral shape. The reason for this shape will be explained in Unit 2.

Properties of alkanes

Alkanes are largely unreactive. This makes them good solvents for reactions, especially when water is unsuitable. Examples include reactions in which water reacts violently with the chemicals. Hexane, cyclohexane and petroleum spirits (a mixture of low-boiling alkanes) are all used in the laboratory for this purpose. Kerosene is a mixture of about 70% (by volume) of branched- and straight-chain alkanes with between 10 and 16 carbon atoms, along with some other hydrocarbons. It is used as an industrial solvent and for fuel. The main reaction of alkanes is combustion, which is the reaction with oxygen to produce carbon dioxide, water and heat energy.

The first four alkanes are gases. The next group of alkanes are volatile liquids. The larger alkanes are waxy solids at room temperature, and have stronger intermolecular forces due to their larger electron dense regions.

TABLE 1 Formulas, 3D models and boiling points of the first six alkanes

Name	Molecular formula	3D model	Condensed structural formula	Boiling point (°C)
Methane	CH_4		CH_4	–162
Ethane	C_2H_6		CH_3CH_3	–89
Propane	C_3H_8		$CH_3CH_2CH_3$	–42
Butane	C_4H_{10}		$CH_3(CH_2)_2CH_3$	–0.5
Pentane	C_5H_{12}		$CH_3(CH_2)_3CH_3$	36
Hexane	C_6H_{14}		$CH_3(CH_2)_4CH_3$	69

Worked example 6.3A

Determining the formula of an alkane

Worked example 6.3A: Watch a video that shows how to solve this problem.

Determine the formula of the linear alkane with 12 carbon atoms. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to apply the general formula for an alkane to find its formula. The question is worth 1 mark, so we must correctly apply the formula.
Step 2: Apply the general formula C_nH_{2n+2}, when $n = 12$.	$2n + 2 = (2 \times 12) + 2$ $= 26$
Step 3: Finalise your answer.	$C_{12}H_{26}$ (1 mark)

Your turn

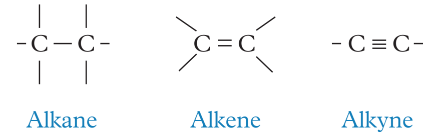
Determine the formula of the linear alkane with 8 carbon atoms. (1 mark)

What are unsaturated hydrocarbons?

Alkenes and **alkynes** are examples of **unsaturated hydrocarbons** because they contain less than the maximum number of hydrogen atoms possible for the number of carbon atoms present.

The general formula for a linear alkene is $C_nH_{2n+2-2P}$ where P is the number of double bonds present. For example, but-1-ene ($H_2C{=}CH{-}CH_2{-}CH_3$) has one double bond and molecular formula C_4H_8, and buta-1,3-diene ($H_2C{=}CH{-}CH{=}CH_2$) has two double bonds and molecular formula C_4H_6 ($n = 4$, so $2 \times 4 + 2 - 2 \times 2 = 6$).

alkene	a hydrocarbon that contains one or more double bonds between carbon atoms
alkyne	a hydrocarbon that contains one or more triple bonds between carbon atoms
unsaturated hydrocarbon	a compound of carbon and hydrogen that contains double or triple bonds and therefore has fewer than the maximum number of hydrogen atoms for a particular number of carbon atoms



$CH_3CH_2CH_3$	$CH_3CH{=}CH_2$	$CH_3C{=}CH$
Propane	Propene	Propyne

FIGURE 4 The general structures of alkanes, alkenes and alkynes and examples with three carbon atoms

Properties of alkenes

Monoalkenes have boiling points close to but slightly lower than those of the alkanes with the same number of carbon atoms. This feature of monoalkenes is because the number of electrons is two fewer than that for the corresponding alkane, due to the two fewer hydrogen atoms. This results in slightly weaker interactions between molecules.

The presence of the double bond in alkenes limits rotation around the double bond, which can lock groups on either side of the double bond into a fixed geometric relationship. Two compounds containing a double bond may have the same molecular formula and atom connectivity, but have groups that are arranged differently in space. These compounds have different physical properties, including boiling points.

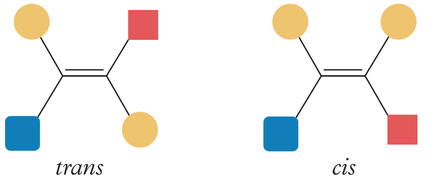


FIGURE 5 The restricted rotation around a double bond results in *trans* and *cis* stereoisomers.

As shown in Figure 5, if two identical groups are on opposite sides of the double bond, it is called the *trans* isomer. If two identical groups are on the same side of the double bond, it is called the *cis* isomer. The groups being compared are attached to the two carbon atoms in the double bond (not the same carbon atom). A different naming system applies if there are four different groups attached to the double-bond carbon atoms.

Alkenes and alkynes show greater chemical reactivity than alkanes and are able to undergo a type of reaction called addition, in which a reactant adds across the double or triple bond to give a final addition product. If a monoalkene reacts, it becomes a saturated molecule. You will learn more about this in Unit 4.

What is benzene?

Benzene is the simplest representative of a class of compounds called **aromatic compounds** (or **arenes**, Figure 6). It is also a hydrocarbon with the formula C_6H_6 and is found in petrol (at around 1% by volume). Alternating double and single bonds are shown around the ring.

benzene
the simplest aromatic compound with the molecular formula C_6H_6, in which the carbon atoms are arranged in a flat (planar) six-membered ring, with one hydrogen atom attached to each carbon atom

aromatic compound/arene
a cyclic unsaturated hydrocarbon compound that has delocalised clouds of electrons

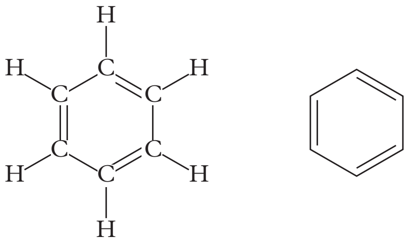


FIGURE 6 Representations of the structure of benzene; the structure on the left shows all of the atoms in benzene, whereas on the right, the carbon atoms are understood to be located at the corners, with one hydrogen atom attached to each carbon atom.

The distances between carbon atoms in compounds change with the type of bond present (Table 2). If benzene actually had alternating single and double bonds, then we would expect there to be three longer C–C single bonds and three shorter C=C double bonds. In reality, when the bond distances are measured, they are the same between all carbon atoms at around 140 pm, which is between the distance expected for single and double C–C bonds.

TABLE 2 Carbon–carbon bond types and typical bond distances

Bond type	Bond distance (pm)
C–C single bond	154
C=C double bond	134
C=C triple bond	109
Bonds between carbon atoms in benzene	140

Chemists explain this with a concept known as resonance. This can be represented by two Lewis structures shown with a double-headed arrow (Figure 7).

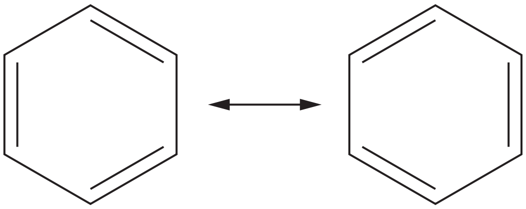


FIGURE 7 Resonance forms of benzenes

However, neither of the two Lewis structures represent the actual benzene molecule. Benzene is not alternating between two different forms. Instead, the real molecule can be described as a hybrid of these two resonance forms, which is more stable than either of them. A more accurate description of the bonding in benzene is that the electrons in the double bonds are delocalised around the ring. These delocalised electrons can be shown as rings or doughnuts above and below the plane that contains the carbon and hydrogen nuclei (Figure 8).

Study tip

To help you remember:

- *trans* isomer: the groups have been *trans* ported apart; *trans* = "cross"
- cis isomer: "*cis*"ters stick together; *cis* = "beside".

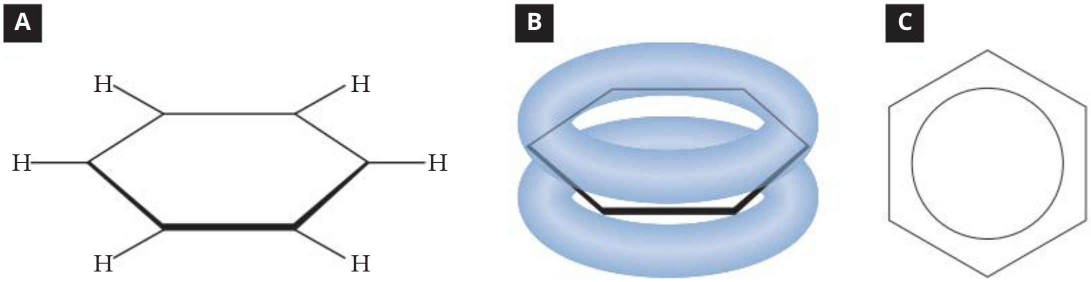


FIGURE 8 (A) The localised bonds in benzene between carbon and hydrogen atoms and (B) the clouds of delocalised electrons. (C) Another way of depicting benzene.

Properties of benzene

The extra stability that comes from these delocalised electrons explains why benzene does not readily undergo the addition reactions that are usually seen for alkenes. If benzene were to undergo an addition reaction, then the extra stability that comes from having resonance would be lost.

Instead, benzene typically undergoes substitution reactions, where a hydrogen atom is replaced by another atom or group of atoms. These substitution reactions often require quite reactive reagents or higher temperatures than are required for substitution reactions of alkenes; this is because of the high stability of benzene. Again, you will learn about this in Unit 4. Benzene is a very important industrial chemical because it can be used to make many other compounds.

Benzene has a boiling point of 80.09°C, which is only 0.66°C lower than that of its fully saturated counterpart cyclohexane (80.75°C). The lower number of electrons in benzene (because benzene has six fewer hydrogen atoms than cyclohexane) should lead to weaker intermolecular forces, but this is compensated for by the clouds of delocalised electrons in benzene, which are easily distorted. This causes the intermolecular forces to be stronger than expected.

Real-world chemistry

Covalent biochemical molecules

All living organisms use large covalent molecules, termed biochemical molecules. Humans obtain these from food; our body's cells use them for growth and repair as well as for energy and energy storage. These molecules all contain carbon and hydrogen. Carbohydrates and lipids also contain oxygen, while proteins contain oxygen, nitrogen and sometimes sulfur and selenium.

The simplest carbohydrates are monosaccharides such as glucose, $C_6H_{12}O_6$. Glucose is transported via the bloodstream to all cells, where it is used in cellular respiration to provide energy. Glucose is largely obtained from the breakdown of complex carbohydrates, such as starch, a polysaccharide made of a long chain of monosaccharides that have been bonded together. Digestion breaks these apart to form monosaccharides. Then, they are temporarily stored as another polysaccharide, glycogen, in the liver.

Proteins are also large molecules, made from many smaller molecules called amino acids. Each amino acid is connected by an amide linkage (CONH, blue in Figure 9). They each have a unique middle section termed a residue (yellow). Humans obtain 20 different amino through protein-containing foods. The digestive system breaks these into the constituent amino acids, so that they can be carried to cells where proteins are built, via the bloodstream.

Lipids are a group of biochemical molecules that are not soluble in water. They largely enter our bodies from a variety of plant-based oils and animal

▶

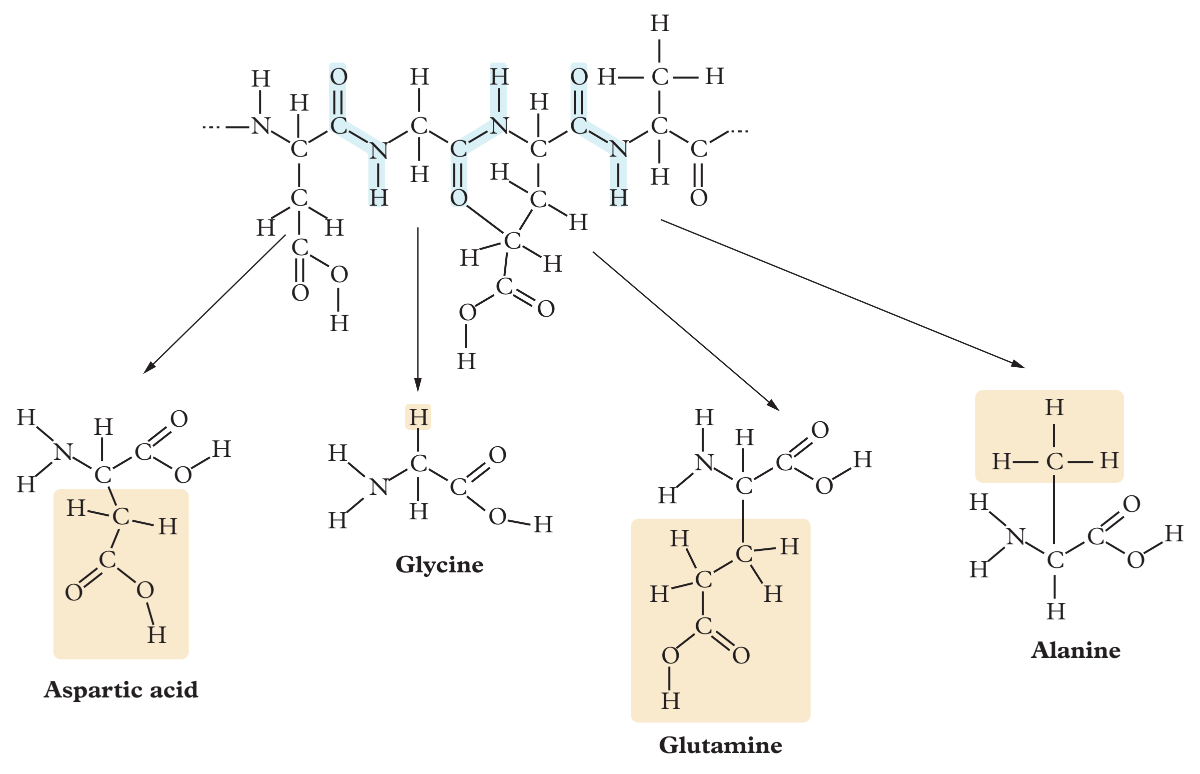


FIGURE 9 A section of a protein showing four amino acids joined together and the constituent amino acids below

fats, which contain triglycerides, three long chain fatty acids bonded together via a glycerol backbone. Our digestive system can break them into fatty acids and package them for absorption into the body. They are used to form the lipid membranes around cells, form deposits within the body that assist with warmth, and can be used as a source of energy.

Apply your understanding

1 When two monosaccharides combine to form a disaccharide, a water molecule is produced. **Use** this fact and the formula of glucose to **determine** the formula of a disaccharide. (1 mark)

2 **Compare** the size and elemental composition of the four amino acid residues in Figure 9. (3 marks)

3 Fatty acids can be saturated, monounsaturated or polyunsaturated. **Use** your knowledge of saturated and unsaturated hydrocarbons to **explain** what this means. (3 marks)

Check your learning 6.3

Check your learning 6.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** why very pure liquid water is a very poor conductor of electricity, but tap water will conduct and seawater will conduct even better. (6 marks)

2 Covalent bonds are extremely strong, yet covalent molecular substances tend to have relatively low melting and boiling points, and are soft. **Explain** why the strong covalent bonds do not influence these properties. (4 marks)

3 **Explain** the differences between linear and branched alkanes by drawing structures of molecules of each type, each containing six carbon atoms. (3 marks)

4 **Explain** why alkanes do not react with hydrochloric acid (HCl) but alkenes do. (2 marks)

Practical

Lesson 6.4

Testing for saturation

Learning intentions and success criteria

Video demonstration

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 6.5

Properties of giant covalent networks

Key ideas

→ Giant covalent networks have high melting points.

→ The conductivity of giant covalent networks can vary.

→ Allotropes are forms of elements that have the same composition but a different structure.

Learning intentions and success criteria

What are giant covalent networks?

Several substances consist of large assemblies of atoms all covalently bonded together to form **giant covalent networks**, with the network extending throughout the solid. These are also called covalent network solids and may contain only one type of atom or more than one type of atom.

giant covalent network
an element or compound in which each atom is covalently bonded to other atoms to form an extended assembly of covalently bonded atoms

Carbon and silicon are elements that form covalent networks made of only one type of atom. These are represented using the symbol for the element, such as C, although diamond and graphite can be further distinguished by a subscript, making them $C_{(diamond)}$ and $C_{(graphite)}$. Giant covalent networks containing two or more different elements are represented with a **formula unit**, which shows the simplest whole number ratio of elements. Examples include SiO_2 and SiC.

Commonly known covalent network compounds include at least one of carbon, silicon, boron or nitrogen; for example, boron nitride, BN. These four elements can each form three or four covalent bonds with neighbouring atoms, enabling them to form large networks.

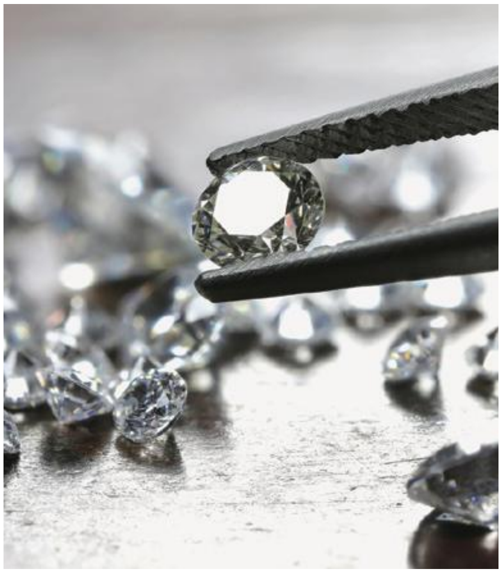


FIGURE 1 Diamond is a giant covalent networks of carbon.

formula unit
the formula of any compound that does not form discrete molecules; the smallest unit of a non-molecular substance, which applies to the formula of covalent network compounds and ionic compounds

What is the structure of giant covalent networks?

All the atoms in giant covalent networks are bonded to their neighbouring atoms with very strong covalent bonds, and this type of bonding extends throughout the entire crystal. In most cases, this is a three-dimensional arrangement. Notably, $C_{(diamond)}$ forms this type of lattice (Figure 2A). The element silicon can form a crystalline lattice of the same form, as does silicon carbide, and crystalline silicon dioxide (quartz, Figure 2B).

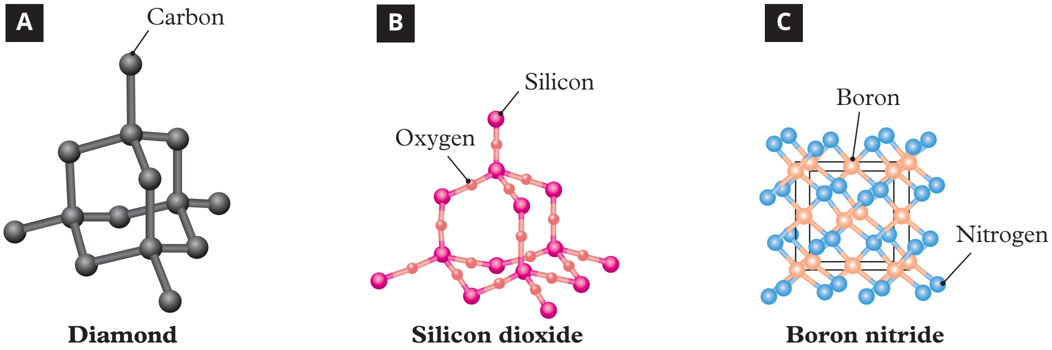


FIGURE 2 Part of the structures of three covalent networks solids. (A) In the structure of diamond, each carbon atom is covalently bonded to four other neighbouring carbon atoms. (B) In quartz (silicon dioxide (SiO_2)), each silicon atom is surrounded by four oxygen atoms in a tetrahedral arrangement and each oxygen atoms bridges two silicon atoms. (C) In cubic boron nitride (BN), each boron atom is bonded to several neighbouring nitrogen atoms.

Some solids form a two-dimensional or single-plane network. The most well-known of these is graphite, made of the element carbon. Each carbon atom in graphite is bonded to three other carbon atoms to form sheets that are stacked on top of each other.

There are no covalent bonds between sheets; they are held together by delocalised electrons as each carbon atom contributes one valence electron not used in the covalent bonds. The dashed lines in Figure 3 represent the attractive forces due to delocalised electrons, and the grey shading emphasises the flat or planar nature of the layers.

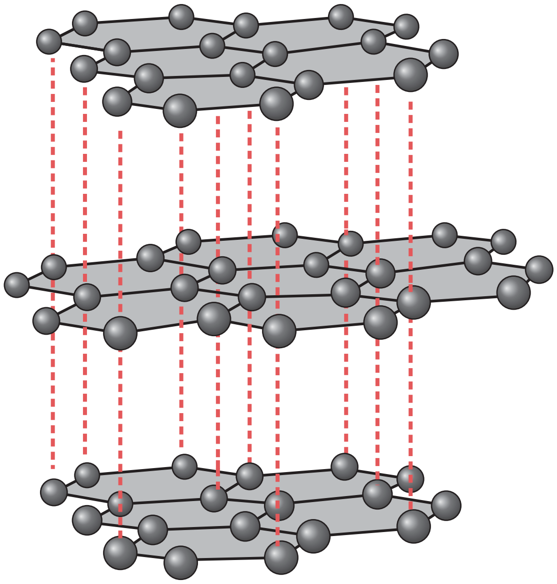


FIGURE 3 Graphite features planes of covalently bonded atoms stacked on top of each other. The dashed lines represent weak intermolecular forces, and the covalent bonds are shown as solid lines.

How does the structure of giant covalent networks explain their physical properties?

In solid form, most giant covalent network elements and compounds share the same physical properties. They owe this to their structures.

Melting points

Covalent bonds are very strong bonds. To melt a solid to a liquid, all the atoms in the lattice must gain a large amount of thermal energy to overcome the strong covalent bonds and move around within the liquid. This requires extremely high temperatures, typically thousands of degrees Celsius.

Thermal and electrical conductivity

Thermal energy can be transmitted by giant covalent networks, as it passes by vibrations through the crystal lattice. Although they are good thermal conductors, they are not very effective conductors of electricity. All electrons in giant covalent network solids are either held with their atoms or within covalent bonds and shared between two atoms. They are not free to move, and so cannot carry electrical current.

Graphite is an exception; it carries current reasonably well parallel to the planes, as the delocalised electrons are able to move. However, graphite is a weak conductor at right angles to its layers.

Hardness

It is difficult to scratch most giant covalent network substances and they are among the hardest natural substances. For this reason, diamond and silicon carbide are used as industrial abrasives, and diamond and various coloured quartzes are prized as precious or semi-precious gemstones for jewellery. The hardness is due to the strong covalent bonding extending throughout the lattice.

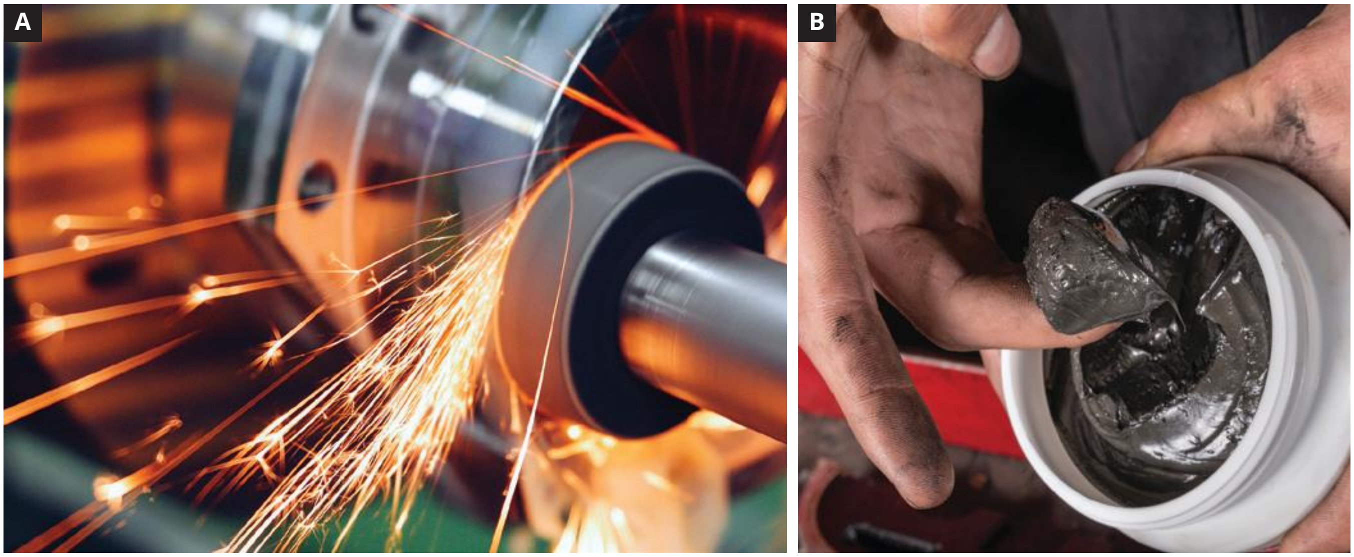


FIGURE 4 (A) Diamonds are added to industrial tools to harden their surfaces and allow better grinding, cutting, drilling and polishing of hard materials. (B) Graphite is used as a lubricant.

Again, graphite is an exception. It is soft due to the two-dimensional layers which can slide over each other easily. It is used as a dry lubricant in industrial applications.

Apart from graphite, any breakages to a giant covalent network can be in any direction within the crystal lattice. The covalent bonds are disrupted and the break will be irregular. For graphite, breaks tends to follow the planes of the two-dimensional layers.

Brittleness

When struck by a force, the atoms are not able to slide past each other. This means that giant covalent networks are not typically malleable. If the force is large enough to break the strong covalent bonds, the crystal will crack or break – that is, they are brittle.

What is the difference between bonding in simple covalent molecules and giant covalent networks?

Covalent molecular substances and giant covalent networks both involve covalent bonding but are otherwise very different. Nearly all non-metallic elements and compounds are covalent molecules, whereas only a few covalent elements and compounds exist as giant covalent networks.

The properties, too, are almost the opposite of each other. The physical properties of covalent molecules are due to the relatively weak intermolecular forces, whereas it is the very strong covalent bonds throughout that influence the properties of giant covalent networks.

What are allotropes of carbon?

Graphite and diamond are both solids at room temperature and both contain only carbon atoms. They are called **allotropes**, because they have the same formula of C, but differ in the way that the atoms are bonded to each other. Allotropes are different forms of the same element, existing in the same state (e.g. solid or gas) under the same conditions of temperature and pressure.

allotrope
one of the different structural forms in which a particular element may exist in a given physical state; exist due to differences in bonding between atoms of that element

Carbon is the element with the greatest number of allotropes (Figure 5). One form is amorphous carbon, which does not form a regular arrangement of its carbon atoms. Graphene has a single-layer arrangement of carbon atoms in a hexagonal arrangement. The last group of carbon allotropes is the fullerenes. Buckminsterfullerene (or a “buckyball”), which has the formula C_{60}, was first synthesised in 1985. Other fullerenes have since been found or synthesised, including nanotubes which resemble a sheet of graphene rolled into a cylinder. Fullerenes have been found in nature and detected in outer space.

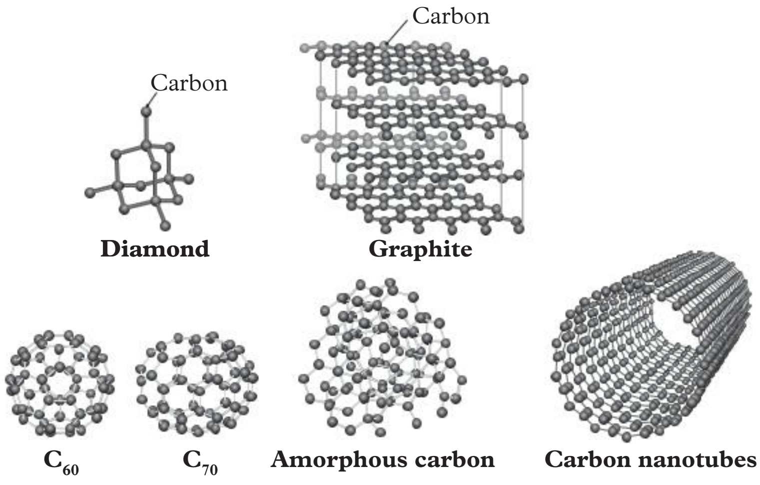


FIGURE 5 Structures of various allotropes of carbon

Real-world chemistry

Nanomaterials

Nanomaterials are substances that contain particles so small that at least one of the dimensions' range (such as length, width or height) is ≤100 nanometres (nm). Nanomaterials can have novel electronic, mechanical or optical properties. For example, UV blockers in sunscreens are now available as nanoparticles of zinc oxide and titanium oxide, forming a less visible layer on the skin than previous zinc creams.

Nanoparticles can be engineered using atoms or simple molecules as building blocks in natural biological processes. This allows scientists to tailor the synthesis of nanostructures.

Currently, silver is the most commonly used nanoparticle, often used for its anti-bacterial properties. Second to this is carbon-based nanomaterials, including graphene and the fullerenes, as well as graphene oxide, nanodiamonds and carbon-based quantum dots. They have wide applications, including energy storage in graphene supercapacitors, use in targeted drug delivery within the body, addition to polymers to form high-strength but low-density materials, and manufacture of composite materials used in sensors for the detection of pollutants.

Nanotechnology is in its early stages yet, and while there are already many uses, research excitingly suggests a great many more uses in the future.

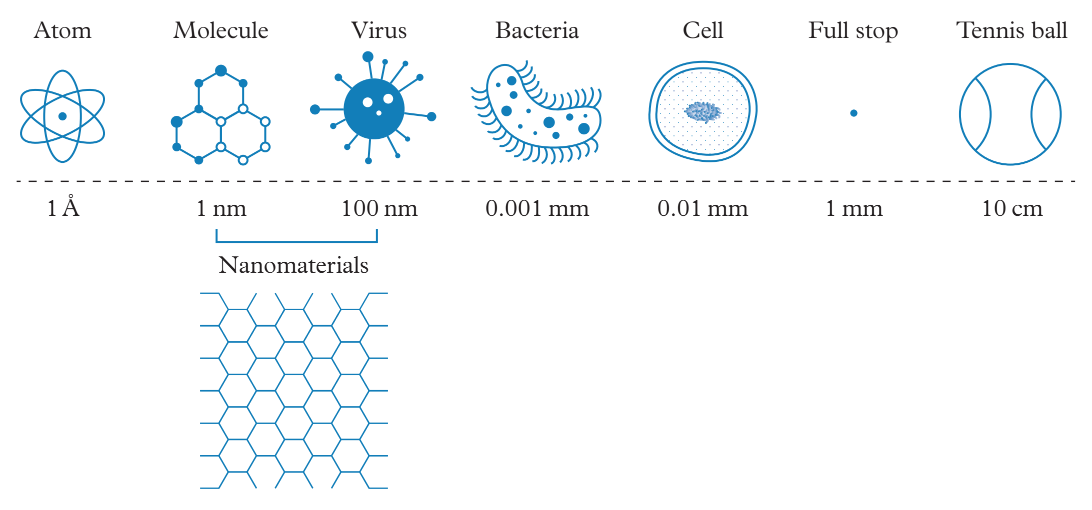


FIGURE 6 Examples of substances across a range of sizes

Apply your understanding

1 The prefix "nano" means 10^{-9}. If a particle is 100 nm long, **calculate** how many would fit end to end in 1 mm. (2 marks)

2 **Identify** the chemical formulas for zinc oxide and titanium oxide. (2 marks)

3 Graphene promises to have multiple uses due to its strength for its size. **Use** your knowledge of the bonding involved to **explain** why. (2 marks)

Challenge

Making diamonds

Explain whether graphite can be converted into diamonds. (2 marks)

FIGURE 7 Diamond

Skill drill

Developing a research question from a claim

Science inquiry skill: Planning investigations (Lesson 1.4)

To complete your research investigation in Unit 4, you will need to write a research question from a claim provided. Practising this skill now will help you with your internal assessment later.

The following claim is made about diamonds: *Synthetic diamonds have more useful properties in industry than silicon.*

To develop a research question, several questions would need to be asked and researched first.

Practise your skills

1 **Describe** the industrial uses of synthetic diamonds and silicon. (2 marks)

2 **Select** an industrial purpose that will be the focus of your question. **Identify** which properties will be considered. (2 marks)

3 **Construct** a suitable research question. (2 marks)

Check your learning 6.5

Check your learning 6.5: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** three different allotropes of carbon and the differences between them. (3 marks)

2 **Explain** why silicon dioxide (quartz) has a high melting point. (3 marks)

Analytical processes

3 **Compare** the structures shown in Figure 2 of diamond and quartz. (3 marks)

4 Covalent molecular substances and giant covalent network substances have some similarities but some key differences. Both have atoms bonded together with very strong covalent bonds, formed by atoms sharing valence electrons.

   a **Construct** labelled diagrams to **compare** covalent molecular and giant covalent network bonding. (2 marks)

   b **Compare** the properties of covalent molecules and giant covalent networks using a Venn diagram. Properties that are shared by both types of substance appear in the section of overlap. The first property has been done for you. (6 marks)

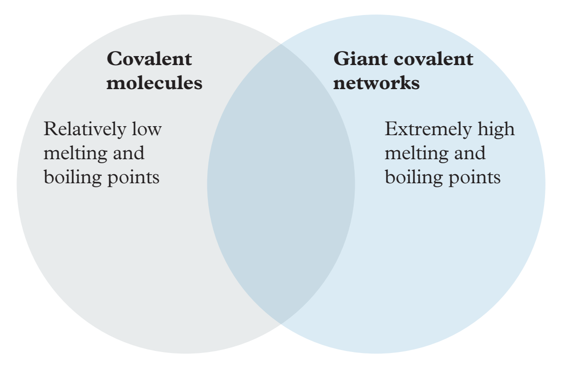


FIGURE 8 Properties of covalent molecules and giant covalent networks

Practical

Learning intentions and success criteria

Video demonstration

Lesson 6.6
Investigating physical properties

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

MODULE 6

Lesson 6.7

Review: Properties and structures of materials

Summary

6.1
- Ionic compounds are held together as solids by the attractions between oppositely charged ions. The ions are held in a repeating three-dimensional arrangement called a crystalline lattice structure.
- Ionic compounds do not conduct electricity in the solid state but can conduct electricity if dissolved in solution or in the molten state.
- Ionic compounds usually have relatively high melting and boiling points, reflecting the strength of the electrostatic interactions between ions.
- Ionic compounds as solids are relatively hard but brittle.

6.2
- Metals in the free electron model of metallic bonding are regarded as positively charged ions in a lattice surrounded by a sea of delocalised electrons. The electrostatic attractions between the cations and negatively charged electrons hold metallic solids together.
- Metals are malleable, ductile and good conductors of electricity and heat. They are lustrous. Generally, they have high boiling points.

6.3
- Covalent molecular substances have strong covalent bonds within their molecules and weak intermolecular forces between their molecules.
- Simple molecules with covalent bonds have low melting and boiling points and are electrically non-conductive as solids. Aqueous solutions are usually non-conductors of electricity unless the molecule reacts with water to form ions.
- Hydrocarbons are compounds composed entirely of carbon and hydrogen atoms. Alkanes (saturated hydrocarbons) contain the maximum number of hydrogen atoms possible for the given number of carbon atoms. Alkenes (unsaturated hydrocarbons) contain one or more double bonds between carbon atoms.
- Benzene (C_6H_6) is composed of six carbon atoms arranged in a ring, with one hydrogen atom attached to each carbon atom.
- Alkenes and alkynes typically undergo addition reactions, as the double and triple bonds can react easily.

6.4
- Practical: Testing for saturation

6.5
- Giant covalent networks consist of many covalent bonds between atoms and generally cannot conduct electricity (although there are exceptions such as graphite). They also have very high melting points. They can conduct heat and are generally very hard, except graphite which is soft.
- Allotropes are forms of the same element that have a different structure. The allotropes of carbon include diamond, graphite, amorphous carbon, the fullerenes and graphene.

6.6
- Practical: Investigating physical properties

Key formulas

Molecular formula of a linear or branched alkane	C_nH_{2n+2}
Molecular formula of a linear or branched alkene	$C_nH_{2n+2-2P}$ where P is the number of double bonds
Molecular formula of a linear or branched alkyne	C_nH_{2n-2}

Review questions 6.7A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Which one of the following shows the order of boiling points correctly from lowest to highest?

A $MgCl_2$, Cl_2, Mg

B Mg, Cl_2, $MgCl_2$

C Cl_2, Mg, $MgCl_2$

D $MgCl_2$, Mg, Cl_2

2 Which of the following substances is hard but brittle, has a high melting point and conducts electricity as a liquid but not as a solid?

A Cu

B CaO

C N_2

D SiO_2

3 Which of the following correctly describes metallic bonding?

A Interactions between pairs of metal atoms with electrons shared between them

B Interactions between positively charged atomic nuclei and outer shell valence electrons

C Electrostatic interactions between electrons and negatively charged metallic anions

D Electrostatic interactions between positively charged metal ions and freely mobile electrons

4 Which of the following is a general characteristic of ionic compounds?

A Low boiling points

B Low electrical conductivity in the solid state

C Soft

D A disorganised arrangement in the solid state

5 Which of the following molecular formulas could represent a linear alkane?

A C_6H_{12}

B C_6H_{14}

C C_6H_{10}

D C_6H_6

6 Which of the following substances does not contain delocalised electrons?

A Silver

B Diamond

C Benzene

D Graphite

7 Which of the following is not a hydrocarbon?

A Ethane

B Propane

C Ethanol

D Methane

8 Which of the following is not an allotrope of carbon?

A Amorphous carbon

B Graphite

C Diamond

D Carbonate

9 Which of the following is not an ionic compound?

A CO_3^{2-}

B $MgCl_2$

C $CaCl_2$

D FeO

10 Metallic and ionic solids

A both contain positively charged ions.

B both have delocalised electrons.

C both have two or more elements.

D both have ions with a 1+ charge.

11 Ionic substances that are soluble in water are conductors of electricity because

A they allow electrons to flow through the solution from the negative electrode to the positive electrode.

B they allow anions to flow through the solution from the negative electrode to the positive electrode.

C they allow cations to flow through the solution from the positive electrode to the negative electrode.

D they allow both cations and anions to flow through the solution in the opposite direction.

12 Which of the following correctly lists properties shared by all covalent molecular substances?

A They are soft and waxy when solid and have low melting and boiling points.

B They do not conduct heat or electricity well as solids.

C They are soft when solid, and dissolve in water to form solutions that do not conduct electricity.

D They dissolve in water easily and have low melting and boiling points.

13 Covalent molecular and covalent network substances have different properties because
  A the covalent bonds in molecules are much weaker than the covalent bonds in covalent networks.
  B covalent molecules are unable to form crystal lattices in the solid state, whereas covalent network substances are present in crystal lattices.
  C covalent molecular substances are soft and have low melting points, while covalent network substances are hard and have high melting points.
  D the intermolecular forces between covalent molecular substances are weak and produce different properties than the strong covalent bonds in covalent networks.

14 Metallic substances are malleable because when a metal is hit
  A the positive and negative ions in metals are able to slide past each other and form new bonds.
  B the metal ions slide past each other and delocalised electrons move around to form new bonds.
  C the delocalised electrons move but the metal ions stay in position.
  D the protons move but are able to form new bonds with the neutrons.

Review questions 6.7B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

15 **Describe** the differences in bonding between potassium iodide and iodine monochloride. (2 marks)

16 **Explain** why benzene typically undergoes substitution reactions rather than the addition reactions usually seen for alkenes, even though its formula would suggest it is an unsaturated hydrocarbon. (2 marks)

17 The properties of ionic compounds depend on their bonding and structure.
  a **Describe** the main features of ionic bonding by drawing a labelled diagram. (2 marks)
  b **Explain** why ionically bonded substances are unable to conduct as solids, despite being made of charged particles. (2 marks)

18 **Explain** why methane has the lowest boiling point of the alkanes. (2 marks)

19 **Explain** why most aqueous solutions of covalent molecular substances such as sucrose do not conduct electricity but some, such as a solution of ethanoic acid, CH_3COOH, will conduct electricity. (4 marks)

20 **Discriminate** between the various allotropes of carbon by comparing their structures. (4 marks)

21 Based on the particles present, and the structure and bonding in each substance, **discriminate** between
  a metals and ionically bonded substances (4 marks)
  b covalent molecular and covalent network substances. (4 marks)

Analytical processes

22 Table 1 shows the results of laboratory tests on four unknown white powders, samples A–D. **Analyse** the data to **determine** the type of bonding present in each of the four substances. (4 marks)

TABLE 1 Results of tests on four unknown white powders

Sample	Melting point	Solubility in water	Electrical conductivity of aqueous solution
A	57°C	soluble	did not conduct
B	did not melt using a Bunsen burner	soluble	conducted
C	126°C	insoluble	unable to test
D	did not melt using a Bunsen burner	insoluble	unable to test

23 The results in Table 2 were obtained in laboratory tests conducted on six different substances, U–Z. The identities of the six substances, not in the same order, are graphite, aluminium, lead, calcium carbonate, silicon dioxide and lithium chloride.

Analyse the data to **determine** the identities of the six substances. (6 marks)

TABLE 2 Results of tests conducted on six different substances

Sample	Electrical conductivity of solid	Electrical conductivity of molten liquid	Electrical conductivity of aqueous solution	Effect of force from a hammer
U	did not conduct	good conductor	good conductor	unable to test, as in powder form
V	good conductor	good conductor	did not dissolve	malleable, bends but does not break
W	good conductor	did not melt with a Bunsen burner	did not dissolve	malleable, bends but does not break
X	did not conduct	did not melt with a Bunsen burner	did not dissolve	very hard, did not break, but a previously broken edge showed an irregular break
Y	did not conduct	did not melt with a Bunsen burner	did not dissolve	brittle, breaks along distinct planes, forming small crystals of the same shape
Z	moderate conductor	did not melt with a Bunsen burner	N/A	brittle, breaks in flakes

Knowledge utilisation

24 Concept maps are useful ways to summarise and communicate information about a topic. Concepts are linked by statements. **Consider** the concept map and **identify** the missing information (grey boxes). Then, **create** a separate, new concept map to communicate information about the concepts involved in metallic bonding. (3 marks)

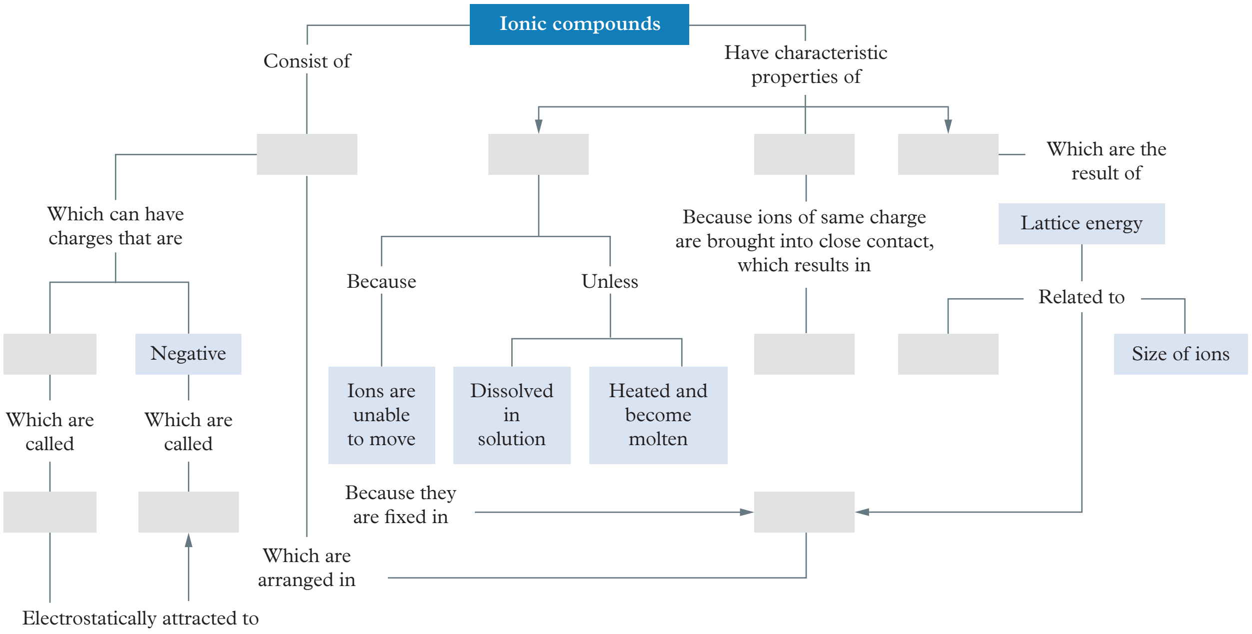


FIGURE 1 Complete this concept map on ionic bonding.

25 **Apply** your understanding of the structures of solids to **propose** the reasoning behind the following observation: Borax ($Na_2B_4O_7{\cdot}10H_2O$), naphthalene ($C_{10}H_8$) and camphor ($C_{10}H_{16}O$) are all compounds used to control insects. Both naphthalene and camphor have distinctive odours and are used in mothballs, while borax does not smell. (6 marks)

Data drill

Identifying substances using melting and boiling points

The melting and boiling points of four substances are shown in the graph in Figure 1. The results of testing for solubility in water are shown in Table 1.

The four substances, not in the same order as A, B, C and D on the graph and table, are ethane (covalent molecule), magnesium (metal), magnesium chloride (ionic compound) and quartz (giant covalent network).

TABLE 1 Solubility of substances A–D

Substance	Solubility in water
A	soluble
B	insoluble
C	not tested
D	insoluble

Apply understanding

1 **Identify** the boiling point of substance C. (1 mark)

2 **Calculate** the difference between the melting point and the boiling point for the substance with the lowest melting point. (2 marks)

Analyse evidence

3 **Organise** the substances A, B, C and D in order of increasing melting points. (1 mark)

Interpret evidence

4 **Determine** the identities of A, B, C and D. **Justify** your choices. (4 marks)

5 The solubility of C in water was not tested. **Predict** whether it would dissolve or not. **Justify** your response. (2 marks)

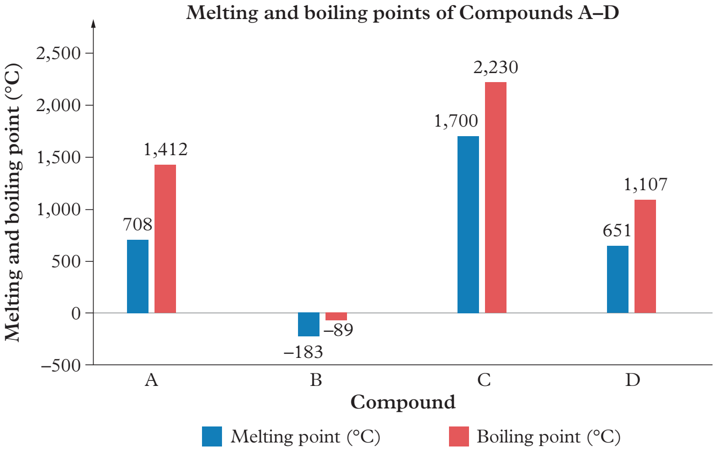


FIGURE 1 Melting and boiling points (°C) of four compounds

Module 6 checklist: Properties and structure of materials

Quizlet: Revise key terms online to test your understanding

UNIT 1

Topic 2 review

Multiple choice

(1 mark each)

1 Which of the following is an example of a homogeneous mixture?
  A Distilled water
  B Clean tap water
  C Seawater sampled from the shoreline
  D Soft drink after it has been first opened

2 The type of bonding present in $CaCO_3$ is
  A metallic.
  B ionic only.
  C covalent only.
  D ionic and covalent.

3 Which of the following correctly lists the properties of all metallic substances?
  A Insoluble in water, conducts electricity when molten, brittle
  B Conducts electricity when solid or liquid (molten), lustrous, ductile
  C Low melting point, conducts electricity when solid and molten (liquid), relatively hard
  D High melting point, insoluble in water, conducts electricity when solid and in solution

4 A soft waxy solid was determined to melt across the range of temperatures from 60°C to 66°C. From this, it can be concluded that the solid
  A is soluble in water.
  B is not a pure substance.
  C is made of very small molecules.
  D consists of one covalent molecular compound.

5 Which of the following descriptions of separation methods is correct?
  A Centrifugation is the separation of a solid from a liquid by pouring the liquid off carefully, leaving the solid behind.
  B Distillation is the purification of a liquid from a mixture by boiling the mixture to form a pure vapour that is cooled and collected.
  C In filtration, a mixture is placed in a container, e.g. a tube, which is spun at high speed, and particles separate according to their size, density and shape.
  D Flotation is where dense solid substances in a liquid mixture will naturally sink to the bottom of a container over time because of gravity, then can be separated using a separating funnel.

6 Which of the following statements applies to compounds?
  A No compound has the same ratio of elements.
  B Compounds consist of two or more elements mixed together.
  C Elements in a compound can be separated by physical or chemical means.
  D Different samples of the same compound will contain elements in the same ratio.

7 Which of the following lists only pure substances?
  A Diamond, glucose, steel, air
  B Sulfur, wood, zinc, ammonia
  C Coffee, petrol, ethanol, silver
  D Nickel, methane, graphite, potassium iodide

8 Brass is a homogeneous alloy made of copper and zinc. This means that
  A copper and zinc are present in a 1 : 1 ratio in brass.
  B brass is a compound, consisting of copper and zinc in a fixed ratio.
  C copper molecules and zinc molecules are mixed evenly throughout the solid.
  D samples from different parts of the same brass object will show identical distributions of copper atoms and zinc atoms.

9 Which of the following statements is true for all allotropes of carbon?
  A They do not conduct electricity.
  B The covalent bonds are single bonds.

C The covalent bonding is arranged in three dimensions.

D Each carbon atom is bonded to four other carbon atoms.

10 Unsaturated hydrocarbons can be distinguished from saturated hydrocarbons by

A determining the density.

B determining the boiling point.

C distilling to see if any water is recovered.

D testing for the rapid decolourisation of Br_2.

11 Which of the following statements best describes the compound benzene, C_6H_6?

A Benzene is unsaturated, as it has three double bonds in the ring.

B Benzene will rapidly decolourise Br_2 due to the double bonds present.

C Benzene will conduct electricity, due to delocalised electrons.

D Benzene has six delocalised electrons present in circular regions above and below the ring-shaped molecule.

12 Covalent molecular substances such as ethanol and oxygen have low boiling points because

A they are soft substances and also have low melting points.

B it does not require much energy to separate the atoms in the molecules to form a vapour.

C covalent bonds are the weakest type of chemical bond, so a small amount of energy is required to boil these substances.

D the attractive forces holding the substances together are relatively weak and require small amounts of energy to overcome them.

13 A solid was struck with a hammer. The result is shown in Figure 1.

FIGURE 1 A solid struck with a hammer

Based on the image shown, it is possible to conclude that the substance

A is a covalent network solid that has a two-dimensional lattice.

B is an ionically bonded solid, that has fractured in a regular fashion.

C that the substance consists of metal ions arranged in a lattice.

D would dissolve in water to produce a solution that conducts electricity.

14 Which of the following is always true of covalent network solids?

A They are extremely hard.

B They may be elements or compounds.

C They are usually good electrical conductors.

D Their atoms are arranged in a three-dimensional lattice.

15 Which of the following is true about elements?

A All pure substances are elements.

B Elements cannot be broken down into simpler substances.

C Elements can only exist as single atoms, not molecules.

D The periodic table lists elements in order of their atomic radius.

Short response

16 **Identify** the separation techniques that could be used to purify the following mixtures.

a Butanol and water. Butanol is completely water soluble and boils at 118°C. (1 mark)

b Sodium sulfate and water. Sodium sulfate is water soluble. (1 mark)

c Ethyl ethanoate and water. Ethyl ethanoate is a liquid at room temperature, is largely insoluble in water, and has a density of $0.902\,g\,cm^{-3}$. (1 mark)

17 **Classify** the following mixtures as heterogeneous or homogeneous. **Justify** your answers.

a A gaseous mixture of halothane ($C_2HBrClF_3$), nitrogen and oxygen used as an inhalable anaesthetic (2 marks)

b Galvanised iron (Note: “galvanised” means zinc-coated) (2 marks)

18 Graphite and diamond are two allotropes of carbon. Both sublime, rather than melt, at temperatures over 3,500°C. Graphite conducts electricity, while diamond is a non-conductor. **Explain** these properties. (4 marks)

19 A student made a series of observations during distillation of a liquid.

- A clear, colourless liquid was distilled.
- The liquid leaving the condenser was clear and colourless.

- The mass of the liquid collected in the receiving flask was 90% of the original mass.
- The starting liquid had a freezing point 5°C lower than that of the condensed liquid.
- A clear, colourless liquid remained in the starting flask at the end of the distillation.
- The liquid remaining in the starting flask had a density 1.2 times that of the liquid which distilled over and condensed in the receiving flask.

Analyse the statements to **identify** which observations support the conclusion that the original liquid was a mixture. **Explain,** in chemical terms, your reasoning. (4 marks)

20 The following statements contain errors. **Identify** the errors and rewrite the statement to **explain** the bonding or property correctly.

a Ionic compounds are brittle, not malleable, because when struck, the metal ions do not slide over each other. (1 mark)

b The electrical conductivity of molten ionic compounds is due to valence electrons being able to move freely through the molten liquid. (1 mark)

c In solid sodium chloride, each sodium ion in the lattice is bonded to the one chloride ion in the lattice that accepted its electron. (1 mark)

d Some ionic compounds, such as potassium chloride, are soluble in water because water molecules are attracted to the ionic potassium chloride molecules and overcome the strong ionic bonding in the lattice. (1 mark)

21 A 1 m^2 sheet of gold with a thickness of 50 nm (230 atoms) can be prepared from a single gram of gold. From 1 g of gold, a wire can be made that is 165 m long and 20 μm thick. **Explain** these properties in terms of the type of bonding in gold. (3 marks)

22 Consider the melting points of a few selected elements as shown in Figure 2.

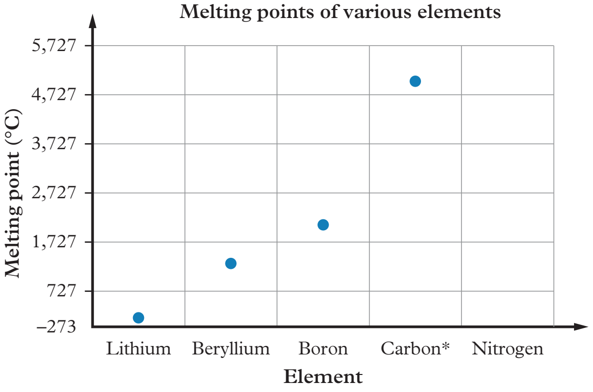


FIGURE 2 Melting points of various elements

*Note: Carbon does not melt at this temperature at 1 bar pressure; instead, it will sublime, that is, go directly to the vapour phase. However, for the purposes of this question, we will refer to the process as “melting” for simplicity of comparison.

a **Explain** why carbon has the highest melting point of the elements shown. (3 marks)

b **Use** your knowledge of bonding to **predict** the melting point of the next element, nitrogen. **Justify** your answer. (2 marks)

23 Students have been provided with four unknown samples of white powdered solids in the laboratory, W, X, Y and Z. The samples are sodium chloride NaCl, calcium carbonate $CaCO_3$, glucose (a simple sugar) $C_6H_{12}O_6$ and citric acid $C_6H_8O_7$, not necessarily in that order. They conduct some tests and record their results in Table 1 to determine which substance is which.

TABLE 1 Results of unknown four substances

Substance	Solubility in water	Electrical conductivity of solution	Ease of melting when heated
W	soluble	does not conduct	melted with a hot Bunsen flame
X	soluble	conducts well	did not melt with a hot Bunsen flame
Y	soluble	conducts slightly	melted with a hot Bunsen flame
Z	insoluble	N/A	did not melt with a hot Bunsen flame

Analyse the data to **determine** the identities of W, X, Y and Z. (4 marks)

24 Many ionic compounds are able to dissolve in water, despite the strength of the bonds between oppositely charged ions in the solid crystal lattice. This is because the attraction between water molecules and ions is sufficiently strong to overcome the strong ionic bonds.

The solubility and other properties of several ionic compounds are shown in Table 2. Note that s = soluble, i = insoluble, and that the ratio of ionic radii is obtained by dividing radius of anion by radius of cation.

TABLE 2 Properties of several ionic compounds

Compound	Solubility in water	Charge of cation	Charge of anion	Radius of cation (nm)	Radius of anion (nm)	Ratio of ionic radii
KI	s	+1	−1	1.37	2.2	1.6
$CuCl_2$	s	+2	−1	0.57	1.81	3.2
CuI_2	i	+2	−1	0.57	2.2	3.9
CuS	i	+2	−2	0.57	1.84	3.2
Al_2O_3	i	+2	−2	0.39	1.35	3.5
BeO	?	+2	−2	0.16	1.35	8.4

Use the data in the table to support your answers to the following questions.

- **a** **Analyse** the data to **determine** the relationship between solubility and the ratio of ionic radii. (1 mark)
- **b** **Analyse** the data to **determine** the relationship between solubility and ionic charge. (1 mark)
- **c** **Predict** whether BeO would be soluble or insoluble in water. **Justify** your prediction. (2 marks)

TOTAL MARKS

/50

MODULE

7 Mole concept and law of conservation of mass

Introduction

Until recently, scientists had not seen an individual atom. This meant that the structure of atoms and the way that they behaved had to be inferred from data gathered from millions of different physical properties and chemical reactions.

Chemists know that atoms are very small, so to make it easier, atoms are counted in large groups, rather than individually, just like we use the term century for a period of one hundred years. Terms such as “mole” are used to describe and compare the extremely large numbers of atoms found in every gram of a substance.

When chemists compared the number of moles of atoms in reactants and products, they found that they were the same. This led to the law of conservation of mass, which states that atoms cannot be created or destroyed in a chemical reaction.

Prior knowledge

Prior knowledge quiz

Check your understanding of the law of conservation of mass and balancing chemical equations before you start.

Subject matter

Science understanding

→ Discriminate between the terms *empirical formula*, *molecular formula* and the *formula unit*.

→ State that a mole is a precisely defined quantity of matter equal to Avogadro’s number of particles.

→ State the law of conservation of mass.

→ Explain that the mole concept relates mass, moles and molar mass.

→ Apply 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. (Formula: moles $(n) = \frac{\text{mass } (m)}{\text{molar mass } (M)}$)

→ Determine 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.

- → Determine limiting reactants.
- → Discriminate between experimental and theoretical yield.
- → Analyse data to determine percentage and theoretical yield. (Formula: percentage yield (%) $= \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$)

Science as a human endeavour

- → Appreciate that chemistry principles can be applied to industrial processes to reduce energy requirements.

Science inquiry

- → Investigate limiting reagent/s and percentage yield.
- → Investigate the empirical formula of a compound from reactions involving mass change.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

These lessons are available on Oxford Digital.

Lesson 7.4 Determining the empirical formula of magnesium oxide
Lesson 7.6 Investigating limiting reactants
Lesson 7.8 Investigating the percentage yield of copper carbonate

Lesson 7.1

The mole

Key ideas

→ A mole is a quantity of matter that is equal to Avogadro's number of particles (6.02×10^{23}).

What is a mole?

When first learning to count, we use special names to refer to groups of numbers, such as, ones, tens, hundreds, thousands and even millions. The names do not always refer to numbers with zeros. A pair usually means two. Trio means three. One dozen is the same as twelve. As the numbers become larger, they can be called billions (10^9, one thousand million) or even googols (10^{100}).

Because molecules and atoms are so small, it is almost impossible to count them individually. Instead, chemists count the number of representative particles in terms of large groups of numbers. The most common of these number groups is 602,000,000,000,000,000,000,000 or 6.02×10^{23}. This number is named after the scientist Amedeo Avogadro, and is often called Avogadro's number, or **Avogadro's constant.**

In the 20th century, scientists linked this large number of atomic particles to the mass of the particles measured in grams and called the concept **mole**. Initially the mole was linked to 16 grams of oxygen atoms. Later, scientists defined the mole as the number of atoms of ^{12}C in 12 grams of the carbon-12 isotope. Careful measurements found this number of atoms to be 6.02×10^{23}. As testing has become more advanced, this number of atoms has varied slightly, but the overall concept remains the same.

Today, the mole is defined by the International Union of Pure and Applied Chemistry (IUPAC) as "the **amount** of substance containing exactly $6.02214076 \times 10^{23}$ elementary entities. This number is the fixed numerical value of the Avogadro's constant, N_A, when expressed in mol^{-1}, and is called the Avogadro number."

- One pair of molecules of water = 2 particles of H_2O
- One dozen molecules of water = 12 particles of H_2O
- One billion molecules of water = 10^9 particles of H_2O
- One mole of water = 6.02×10^{23} particles of H_2O
- One googol of water = 10^{100} particles of H_2O

The number of molecules, amount (in moles) and Avogadro's number are linked by the following equation:

Number of molecules = number of moles of substance × Avogadro's constant

When calculating the number of particles, scientists often use symbols instead of words. In chemistry, the mole is represented by the letter n and is abbreviated to "mol". Instead of writing "the number of moles of water", a chemist may write $n(H_2O)$. In the same way, N is used to represent the number of molecules, ions or atoms, and N_A represents Avogadro's constant. This means we can rewrite the equation as:

$$N = n \times N_A$$

Learning intentions and success criteria

Avogadro's constant
the number of atoms or molecules in one mole of a substance; equivalent to 6.02×10^{23} particles per mol (mol^{-1})

mole
the quantity of a substance containing roughly 6.02×10^{23} particles (Avogadro's number)

amount
the number of moles

Study tip

Review your understanding of scientific notation. Ensure you can accurately enter the values into your calculator and read them in scientific notation from your calculator screen.

Study tip

For covalent substances, N is usually the number of molecules. For ionic substances, it is typically the number of formula units.

Let's consider water (H_2O) as an example. Each water molecule has a single oxygen atom and two hydrogen atoms. This means two molecules of water have a total of two oxygen atoms and four hydrogen atoms. If the number of water molecules was one mole, then there would be one mole of oxygen atoms and two moles of hydrogen atoms. Although the number of water molecules has increased (to 6.02×10^{23} molecules), the ratio of hydrogen atoms to oxygen atoms has not changed.

FIGURE 1 One mole of a solid, a gas and a liquid

Study tip

A mole is such a huge number, it is difficult to imagine. If 1 mol of peanut butter sandwiches could be made and placed on the surface of the earth, the depth of peanut butter sandwiches would be over 300 km stretching well beyond the atmosphere into space!

Worked example 7.1A

Calculating the number of covalent molecules and atoms

▶ **Worked example 7.1A:** Watch a video that shows how to solve this problem.

For a sample of dry ice containing 4.5 mol of carbon dioxide (CO_2), **calculate**:

a the number of carbon dioxide molecules (2 marks)

b the number of oxygen atoms in 4.5 mol of carbon dioxide molecules. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to apply the formula for number of particles. The questions are worth a variety of marks, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required.	$N = n \times N_A$ $N(CO_2) = ?$, $n = 4.5\,\text{mol}$, $N_A = 6.02 \times 10^{23}$ $N(O) = ?$
Step 3: For part **a**, substitute the known values into the formula and solve for the unknown value.	$N(CO_2) = 4.5\,\text{mol} \times 6.02 \times 10^{23}$ (1 mark) $= 2.709 \times 10^{24}$ molecules
Step 4: For part **b**, multiply the number of particles of the molecule by the number of atoms per molecule.	There are 2 × O atoms in each CO_2 molecule, so we multiply by 2. $N(O) = 2 \times N(CO_2)$ $= 5.418 \times 10^{24}$ molecules
Step 5: Finalise your answers and make sure you have expressed them using the correct units and number of significant figures.	**a** 2.7×10^{24} molecules (1 mark) **b** 5.4×10^{24} atoms (1 mark)

Your turn

For a sample containing 8.0 moles of methane (CH_4), **calculate**:

a the number of methane molecules. (2 marks)

b the number of hydrogen atoms. (1 mark)

Worked example 7.1B

Calculating the number of ionic substance and ions

Worked example 7.1B: Watch a video that shows how to solve this problem.

A sample contains 0.25 mol of aluminium sulfate ($Al_2(SO_4)_3$). **Calculate**:

a the number of formula units of aluminium sulfate present. (2 marks)

b the number of aluminium ions and sulfate ions present. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to apply the formula for number of particles. The questions are worth 2 marks each, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required.	$N = n \times N_A$ $N(Al_2(SO_4)_3) = ?$, $n = 0.25\,mol$, $N_A = 6.02 \times 10^{23}$ $N(Al) = ?$, $N(SO_4{}^{2-}) = ?$
Step 3: For part **a**, substitute the known values into the formula and solve for the unknown value.	$N\left(Al_2(SO_4)_3\right) = 0.25\,mol \times 6.02 \times 10^{23}$ (1 mark) $= 1.505 \times 10^{23}$ formula units
Step 4: For part **b**, multiply the number of formula units of the compound by the number of atoms per unit.	There are 2 × Al^{3+} ions in $Al_2(SO_4)_3$, so we multiply by 2. $N(Al^{3+}) = 2 \times N\left(Al_2(SO_4)_3\right)$ $= 3.01 \times 10^{23}$ ions There are 3 × $SO_4{}^{2-}$ ions in $Al_2(SO_4)_3$, so we multiply by 3. $N(SO_4^{2-}) = 3 \times N\left(Al_2(SO_4)_3\right)$ $= 4.515 \times 10^{23}$ ions
Step 5: Finalise your answers and make sure you have expressed them using the correct units and number of significant figures.	**a** 1.5×10^{23} formula units (1 mark) **b** 3.0×10^{23} Al^{3+} ions (1 mark) and 4.5×10^{23} $SO_4{}^{2-}$ ions (1 mark)

Your turn

A sample contains 8.0 mol of iron(III) nitrate ($Fe(NO_3)_3$). **Calculate**:

a the number of formula units of iron(III) nitrate present (2 marks)

b the number of iron(III) ions and nitrate ions present. (2 marks)

Worked example 7.1C

Calculating number of moles from number of particles

Worked example 7.1C: Watch a video that shows how to solve this problem.

There are 1.5×10^{24} molecules of CO_2 in a cube of dry ice. **Calculate** the amount, in mol, of dry ice. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Calculate” means to determine or find a number or answer by using mathematical processes. We need to apply the formula for number of particles. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required. Recall that “amount” means “number of moles”.	$N = n \times N_A$ $N(CO_2) = 1.5 \times 10^{24}$, $n = ?$, $N_A = 6.02 \times 10^{23}$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$1.5 \times 10^{24} = n \times 6.02 \times 10^{23}$ (1 mark) $n = \frac{1.5 \times 10^{24}}{6.02 \times 10^{23}}$ $= 2.492$ mol
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	2.5 mol (1 mark)

Your turn

There are 7.02×10^{23} molecules of water in a glass. **Calculate** the amount, in mol, of water. (2 marks)

Study tip

Revise the index laws before completing these questions.

$$\frac{10^a}{10^b} = 10^{a-b}$$
$$10^a \times 10^b = 10^{a+b}$$
$$10^0 = 1$$

Check your learning 7.1

Check your learning 7.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define** “mole”, with reference to Avogadro’s number. (2 marks)

2 **Identify** how many atoms are in 12.000 g of ^{12}C. (1 mark)

3 **Calculate** the number of hydrogen atoms in:
  - **a** 1.0 mol of hydrogen atoms (2 marks)
  - **b** 1.0 mol of H_2 gas (3 marks)
  - **c** 0.50 mol of HCl (3 marks)
  - **d** 3.5 mol of H_2O. (3 marks)

4 A sample contains 2.00 mol of CH_3COOH. **Calculate** the number of atoms of:
  - **a** carbon (3 marks)
  - **b** hydrogen (1 mark)
  - **c** oxygen. (1 mark)

5 A sample contains 0.200 mol of $Ca(OH)_2$. **Calculate** the number of:
  - **a** calcium ions (3 marks)
  - **b** hydroxide ions. (1 mark)

6 **Calculate** the amount, in mol, of:
  - **a** sugar in a cup of sugar (sucrose) containing 3.52×10^{23} molecules of sucrose (2 marks)
  - **b** NaCl in a spoon of salt containing 2.55×10^{23} formula units of sodium chloride. (2 marks)

Lesson 7.2

The mass of a mole

Key ideas

→ The law of conservation of mass states that mass cannot be created nor destroyed in a chemical reaction.

→ The molar mass is the mass of one mole of a substance.

→ The mole is equivalent to the mass of a substance divided by its molar mass.

What is the law of conservation of mass?

Learning intentions and success criteria

law of conservation of mass states that in an isolated or closed system, mass cannot be created nor destroyed

The **law of conservation of mass** states that mass is neither created nor destroyed in a chemical reaction. Consider the atoms that made up the carbohydrates in prehistoric plants that were eaten by dinosaurs, digested into individual atoms and rearranged into new molecules in the dinosaur's body. These molecules were later broken down into the same number and type of atoms when the dinosaur died. These same atoms were recycled and still exist around us today.

If the number and type of atoms in a closed system do not change, then the mass of atoms in that closed system will not change. For example, one carbon atom and two oxygen atoms have the same mass as one molecule of carbon dioxide.

$$\text{mass } (CO_2) = \text{mass (1 carbon atom)} + \text{mass (2 oxygen atoms)}$$

This can be extended to larger numbers of carbon and oxygen atoms:

$$\text{mass } (1\,\text{mol } CO_2) = \text{mass (1 mol carbon atoms)} + \text{mass (2 mol oxygen atoms)}$$

What is molar mass?

molar mass the mass of one mole of a substance

The mass of one mole of an element is known as the **molar mass**, and is given the symbol *M*. The molar mass is measured in units of g/mol or $g\,mol^{-1}$. If the molar mass of carbon-12 is $12\,g\,mol^{-1}$, then the mass of one mole of carbon-12 atoms is 12 g. This means the mass of 6.02×10^{23} atoms of carbon-12 (1 mol) is also 12 g.

$$\begin{aligned} M(\text{carbon-12}) &= \text{molar mass of carbon-12} \\ &= \text{mass of one mol } (6.02 \times 10^{23} \text{ atoms) of carbon-12} \\ &= 12\ g\,mol^{-1} \end{aligned}$$

The relative atomic masses of each element are known and are generally found on the periodic table or lists of atomic masses. If the chemical formula is known, the molar mass, in $g\,mol^{-1}$, is calculated by adding the relative atomic masses of the constituent atoms.

Worked example 7.2A

Calculating molar masses from the formula

Worked example 7.2A: Watch a video that shows how to solve this problem.

Calculate the molar masses of:

a CO_2 (1 mark)

b $Ba(NO_3)_2$. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to find the relative atomic masses of the constituent atoms to determine the molar mass. The questions are worth 1 mark each, so we must gather the correct information and complete the calculations.
Step 2: Gather any data required.	**a** $M(CO_2)$=? RAM(C) = 12.01 g mol⁻¹, RAM(O) = 16.00 g mol⁻¹ **b** $M(Ba(NO_3)_2)$ = ? RAM(Ba) = 137.33 g mol⁻¹, RAM(N) = 14.01 g mol⁻¹, RAM(O) = 16.00 g mol⁻¹
Step 3: Add the relative atomic masses of the constituent elements. Remember to multiply the RAM of each atom by the number of that atom in the compound.	**a** $M(CO_2) = 12.01\,\text{g mol}^{-1} + (2 \times 16.00\,\text{g mol}^{-1})$ $= 44.01\,\text{g mol}^{-1}$ (1 mark) **b** $M(Ba(NO_3)_2) = 137.33\,\text{g mol}^{-1} + (2 \times 14.01\,\text{g mol}^{-1}) + (2 \times 3 \times 16.00\,\text{g mol}^{-1})$ $= 261.35\,\text{g mol}^{-1}$ (1 mark)

Your turn

Calculate the molar masses of:

a NH_4NO_3 (1 mark)

b C_2H_6. (1 mark)

How are molar mass, mass and number of moles (amount) related?

The molar mass can always be found using relative atomic masses if the formula is known. However, it is more useful for chemists to know the amount, in moles, of a given mass, in grams.

Consider the following: 1 mol of carbon dioxide is 44.01 g (as calculated above). So, 88.02 g of carbon dioxide would be an amount of 2.00 mol, and 4.401 g would be an amount of 0.100 mol. These examples show that the amount, in mol, can be calculated using the formula:

$$\text{moles } (n) = \frac{\text{mass } (m)}{\text{molar mass } (M)}$$

n is the number of moles, m is the mass of a sample in g, and M is the molar mass of the substance in g mol⁻¹.

This formula can be rearranged to find the mass of a given amount, or even the molar mass if the chemical formula is not known. Chemists work using moles rather than single molecules or formula units, as a mole represents a measurable amount. The mass of a mole of a substance can be found by using the molar mass of its elements.

Study tip

The equation $n = \frac{m}{M}$ can be easily rearranged by exchanging the molar mass (M) and the number of moles (n) to form $M = \frac{m}{n}$ and $m = \frac{n}{M}$.

FIGURE 1 These beakers contain one mole of copper(II) sulfate pentahydrate (on the scale) and (from top right clockwise) sucrose, potassium manganate(VII), sodium chloride, iron filings, copper turnings and nickel(II) chloride.

Worked example 7.2B

Calculating *n*, *m* and *M*

Worked example 7.2B: Watch a video that shows how to solve this problem.

Calculate:

a the amount, in mol, of 45.0 g of H_2O (2 marks)

b the mass of 0.500 mol of NaOH (2 marks)

c the molar mass of a 0.156 mol sample of a gas with a mass of 2.50 g. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to apply the formula for number of moles. Each question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required.	$n = \frac{m}{M}$ **a** $n = ?, m = 45.0\,\text{g}, M = 18.02\,\text{g mol}^{-1}$ **b** $n = 0.500\,\text{mol}, m = ?, M = 40.00\,\text{g mol}^{-1}$ **c** $n = 0.156\,\text{mol}, m = 2.50\,\text{g}, M = ?$
Step 3: Substitute the known values into the formulas and solve for the unknown values.	**a** $n = \frac{45.0\,\text{g}}{18.02\,\text{g mol}^{-1}}$ (1 mark) $= 2.497\,\text{mol}$ **b** $0.500\,\text{mol} = \frac{m}{40.00\,\text{g mol}^{-1}}$ (1 mark) $m = 0.500\,\text{mol} \times 40.00\,\text{g mol}^{-1}$ $= 20.00\,\text{g}$ **c** $0.156\,\text{mol} = \frac{2.50\,\text{g}}{M}$ (1 mark) $M = \frac{2.50\,\text{g}}{0.156\,\text{mol}}$ $= 16.03\,\text{g mol}^{-1}$
Step 4: Finalise your answers and make sure you have expressed them using the correct units and number of significant figures.	**a** 2.50 mol (1 mark) **b** 20.0 g (1 mark) **c** $16.0\,\text{g mol}^{-1}$ (1 mark)

Your turn

Calculate:

a the amount, in mol, of 10.0 g of $MgSO_4$ (2 marks)

b the mass of 2.00 mol of NO_2 (2 marks)

c the molar mass of a 0.250 mol sample of a gas with a mass of 11.0 g. (2 marks)

Check your learning 7.2

Check your learning 7.2: Complete these questions online or in your workbook.

Retrieval comprehension

1 **Define:**

  a mass (1 mark)

  b relative atomic mass (1 mark)

  c molar mass. (1 mark)

2 **Calculate** the molar mass of the following, using relative atomic masses:

  a CH_4 (1 mark)

  b $Fe(NO_3)_3$ (1 mark)

  c SiO_2 (1 mark)

  d $CuSO_4{\bullet}5H_2O$. (1 mark)

3 **Calculate** the amount, in moles, in 100.0 g of:
  - **a** CH_4 (2 marks)
  - **b** $Fe(NO_3)_3$ (2 marks)
  - **c** SiO_2 (2 marks)
  - **d** $CuSO_4{\bullet}5H_2O$. (2 marks)

4 **Calculate** the amount, in moles, of:
  - **a** 24 g of O (2 marks)
  - **b** 2.7 g of CO_2 (2 marks)
  - **c** 2.0 g of NaCl (2 marks)
  - **d** 4.6 g of ethanol (C_2H_5OH). (2 marks)

5 **Calculate** the mass of:
  - **a** 0.25 mol of water (2 marks)
  - **b** 0.20 mol of zinc (2 marks)
  - **c** 18×10^{23} atoms of magnesium (4 marks)
  - **d** 18.0×10^{23} formula units of magnesium chloride ($MgCl_2$). (4 marks)

Analytical processes

6 **Determine** the identity of a pure substance if 3.5 mol of the substance has a mass of 59.584 g. The substance contains hydrogen atoms and another element. (3 marks)

Lesson 7.3

Empirical and molecular formulas

Key ideas

→ The percentage composition of a compound is determined from the relative atomic masses of each element.

→ The empirical formula as the simplest whole-number ratio of elements in a compound. It is determined from percentage composition by mass.

→ The molecular formula of a compound can be determined from its empirical formula and molar mass.

What is percentage composition?

Learning intentions and success criteria

Over 200 years ago, electricity was used to separate the atoms in water to form two different gases, oxygen and hydrogen. This process, known as electrolysis, produced exactly twice as much hydrogen gas as oxygen gas. This suggested that the ratio of the elements was two hydrogen atoms to one oxygen atom.

But oxygen atoms are much larger and have more mass than hydrogen atoms. The single oxygen atom contributes more to the mass of the water than the two hydrogen atoms. The amount of mass each element contributes to a molecule can be expressed as a **percentage composition**. For example, the percentage composition of water is 11.21% hydrogen and 88.81% oxygen.

percentage composition
the percentage that each element contributes to the mass of a compound

The percentage composition of an element (x) in a compound can be calculated by the following formula:

$$\%\text{ of } x \text{ in a compound} = \frac{\text{mass of } x \text{ in 1 mol of a compound}}{\text{molar mass of compound}} \times \frac{100}{1}\%$$

Worked example 7.3A

Determining percentage of an element in a compound

Worked example 7.3A: Watch a video that shows how to solve this problem.

Determine the percentage of carbon by mass in ethanol, CH_3CH_2OH. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to compare the mass of carbon in 1 mol of ethanol with the mass of 1 mol of ethanol. The question is worth 2 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Select the correct formulas and gather any data required. Remember to multiply the A_r of each atom by the number of that atom in the compound.	$\%\text{ of } x \text{ in a compound} = \frac{\text{mass of } x \text{ in 1 mol of a compound}}{\text{molar mass of compound}} \times \frac{100}{1}$ $\%(C) = ?$ $M(\text{C in 1 mol of } CH_3CO_2OH) = 2 \times 12.01\,g\,mol^{-1} = 24.02\,g\,mol^{-1}$ $M(CH_3CH_2OH) = 46.08\,g\,mol^{-1}$
Step 3: Substitute the known values into the formulas and solve for the unknown value.	$\%(C) = \frac{24.02\,g\,mol^{-1}}{46.08\,g\,mol^{-1}} \times \frac{100}{1}$ (1 mark) $= 52.127\%$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	52.13% (1 mark)

Your turn

Determine the percentage of hydrogen by mass in water. (2 marks)

How do you determine the empirical formula?

To determine the chemical formula of a new compound, the percentage composition of all the elements in the compound are determined, and then the relative proportions of each element are identified. This simplest whole-number ratio of elements in a compound is called the **empirical formula**. The empirical formula is like the simplest fraction in mathematics – it cannot be subdivided any further at present.

An empirical formula can be determined experimentally by establishing the ratio of the mass of each element in a compound. The empirical formula is calculated by:

1. Identifying all the elements present in the sample
2. Determining the percentage mass of each element
3. Using the percentage mass to determine the number of moles of each element
4. Determining the whole-number ratio by dividing all numbers of moles by the smallest number of moles
5. If necessary, using a multiplication factor to achieve whole numbers.

empirical formula
the simplest whole-number ratio of the elements present in a compound

Study tip

We tend to use "empirical formula" to describe the simplest whole-number ratio of elements in a covalent compound. When talking about the simplest whole-number ratio of elements in an ionic compound, the term "formula unit" is more commonly used.

Worked example 7.3B

Determining the empirical formula from percentage mass

Worked example 7.3B: Watch a video that shows how to solve this problem.

A compound contains 40.00% carbon, 6.70% hydrogen and 53.30% oxygen by mass. **Determine** the empirical formula of the compound. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to find the empirical formula of the compound. The question is worth 4 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Select the correct formulas and gather any data required. Let's assume there is 100 g of the compound in total. To determine the mass of each element, multiply the percentage by 100 g.	$n = \frac{m}{M}$ $n(\text{C}) = ?$, $n(\text{H}) = ?$, $n(\text{O}) = ?$ $\%(\text{C}) = 40.00\%$ so $m(\text{C}) = 40.00\,\text{g}$, $M(\text{C}) = 12.01\,\text{g}\,\text{mol}^{-1}$ $\%(\text{H}) = 6.70\%$ so $m(\text{H}) = 6.70\,\text{g}$, $M(\text{H}) = 1.01\,\text{g}\,\text{mol}^{-1}$ $\%(\text{O}) = 53.30\%$ so $m(\text{O}) = 53.30\,\text{g}$, $M(\text{O}) = 16.00\,\text{g}\,\text{mol}^{-1}$
Step 3: Substitute the known values into the formula and solve for the number of moles of each element.	$n(\text{C}) = \frac{40.00\,\text{g}\,\text{mol}^{-1}}{12.01\,\text{g}\,\text{mol}^{-1}} = 3.331\,\text{mol}$ $n(\text{H}) = \frac{6.70\,\text{g}\,\text{mol}^{-1}}{1.01\,\text{g}\,\text{mol}^{-1}} = 6.634\,\text{mol}$ $n(\text{O}) = \frac{53.30\,\text{g}\,\text{mol}^{-1}}{16.00\,\text{g}\,\text{mol}^{-1}} = 3.331\,\text{mol}$ (1 mark)
Step 4: Express these as a mole ratio.	C : H : O 3.331 mol:6.634 mol:3.331 mol (1 mark)
Step 5: Determine the whole number ratio of moles by dividing all numbers of moles by the smallest number of moles. If necessary, use a multiplication factor to achieve whole numbers. In this case, all three numbers are already whole numbers. Slight rounding may be required, depending on accuracy of initial measurements.	C : H : O $\frac{3.331\,\text{g}\,\text{mol}^{-1}}{3.331\,\text{g}\,\text{mol}^{-1}} : \frac{6.634\,\text{g}\,\text{mol}^{-1}}{3.331\,\text{g}\,\text{mol}^{-1}} : \frac{3.331\,\text{g}\,\text{mol}^{-1}}{3.331\,\text{g}\,\text{mol}^{-1}}$ (1 mark) 1 : 2 : 1
Step 6: Finalise your answer. Remember that this is the simplest ratio of elements, and the molecular formula may be a multiple of this.	CH_2O (1 mark)

Your turn

The compound known as laughing gas contains 63.7% nitrogen by mass and the remainder is oxygen. **Determine** the empirical formula of the compound. (4 marks)

Skill drill

Designing an experiment to determine empirical formula

Science inquiry skill(s): Planning investigations (Lesson 1.4)

What must be considered when designing an investigation? The syllabus requires you to:

"design investigations, including the procedure/s to be followed, the materials required, and the type and amount of primary and/or secondary data required to obtain valid and reliable evidence, e.g.

→ consider replicates, number of data points, and quality of sources

→ identify the types of errors, extraneous variables or confounding factors that are likely to influence results and implement strategies to minimise systematic and random error."

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

In Practical 7.4, you would have made a single set of measurements with the aim of finding the empirical formula of magnesium oxide. Possible sources of error could have included incomplete combustion of magnesium and loss of product through "smoke" escaping.

Practise your skills

1 **Describe** how you would modify the procedure to ensure you have sufficient and relevant data points. **Consider** replicates. (2 marks)

2 **Comment** on any sources of error that noticeably impacted the results in the original practical. **Describe** how you would modify the procedure to overcome these sources of error. (2 marks)

3 **Construct** a research question suitable for this investigation. (2 marks)

How do you determine the molecular formula?

molecular formula
the chemical formula of a compound that indicates the number of atoms of each element present in one molecule

The **molecular formula** is the complete chemical formula of a covalent molecular compound that shows how many atoms of each element are in one molecule of the compound. The empirical formula of water is the same as the molecular formula, H_2O. For other molecular compounds, the molecular formula may be a whole-number multiple of the empirical formula. For example, the empirical formula of butene is CH_2, but its molecular formula is C_4H_8 (four times the number of atoms found in the empirical formula).

$$\text{Molecular formula } (C_4H_8) = 4 \times \text{empirical formula } (CH_2)$$

Some compounds have the same empirical formula as each other, but different molecular formulas. If a compound has a different molecular formula, it will have different chemical properties. For example, ethene, propene and butene have the empirical formula CH_2 (Table 1). Ethene is used in the manufacture of plastics, antifreeze and solvents. Propene and butene are useful raw materials for a variety of important products.

TABLE 1 A comparison of empirical and molecular formula

Name	Empirical formula	Molecular formula
Ethene	CH_2	C_2H_4
Propene	CH_2	C_3H_6
Butene	CH_2	C_4H_8

Study tip

Always write the calculation you are completing in the equation. For example, distinguishing between the empirical and the molecular formula will help to prevent errors.

The molecular formula can be obtained from the empirical formula if the compound's molar mass is known, using the following steps:

1 Find the empirical formula.

2 Find the molar mass of the empirical formula.

3 Calculate the molecular formula by multiplying the empirical formula by

$$\frac{M(\text{molecular formula})}{M(\text{empirical formula})}$$

Worked example 7.3C

Determining the molecular formula

Worked example 7.3C: Watch a video that shows how to solve this problem.

The empirical formula of compound is CH_2 and the molar mass is 28.054 g mol^{-1}. **Determine** the molecular formula of the compound. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to find the molecular formula of the compound. The question is worth 4 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Select the correct formulas and gather any data required.	Empirical formula = CH_2 $M(\text{C}) = 12.01\,\text{g mol}^{-1}$ $M(\text{H}) = 1.01\,\text{g mol}^{-1}$ $M(\text{CH}_2) = ?$, $M(\text{compound}) = 28.054\,\text{g mol}^{-1}$ Molecular formula = ?
Step 3: Calculate the molar mass of the empirical formula using the relative atomic masses.	$M(\text{CH}_2) = 12.01\,\text{g mol}^{-1} \times (2 \times 1.01\,\text{g mol}^{-1})$ (1 mark) $= 14.03\,\text{g mol}^{-1}$ (1 mark)
Step 4: Divide the molar mass of compound by the molar mass of empirical formula to find the multiple.	$\frac{M(\text{molecular formula})}{M(\text{empirical formula})} = \frac{28.054\,\text{g mol}^{-1}}{14.03\,\text{g mol}^{-1}}$ $= 2$ (1 mark)
Step 5: Finalise your answer by multiplying each subscript in the empirical formula by the multiple.	$C_{1\times2}H_{2\times2} = C_2H_4$ (1 mark)

Your turn

A compound is found to have an empirical formula of CH_2O and a molar mass of 60.06 g mol^{-1}. **Determine** the molecular formula of the compound. (4 marks)

Check your learning 7.3

Check your learning 7.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** what is meant by the percentage composition of a compound. (1 mark)

2 **Explain** how the empirical formula of a compound can be different from its molecular formula. (2 marks)

Analytical processes

3 **Determine** the percentage composition of each element in glucose ($C_6H_{12}O_6$). (6 marks)

4 A new compound which may have antibacterial properties has been identified in a plant in the Amazon rainforest. It contains 47.37% carbon, 10.59% hydrogen and 42.04% oxygen. **Determine** the empirical formula for this new compound. (4 marks)

5 A compound is found to contain only boron and hydrogen. If it contains 11.55% hydrogen, **determine** its empirical formula. (5 marks)

6 A compound with a molar mass of 34.02 g mol^{-1} is known to contain 6.3% of hydrogen and 93.7% of oxygen. **Determine** the empirical and molecular formula for this compound. (8 marks)

Practical

Lesson 7.4

Determining the empirical formula of magnesium oxide

Learning intentions and success criteria

Video demonstration

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 7.5

Stoichiometric ratios

Learning intentions and success criteria

Key ideas

→ The mole concept and the mole ratio in chemical equations can be used to calculate the mass of reactants and products in a chemical reaction. These types of calculations are called stoichiometry.

→ Stoichiometric ratios can be used to determine limiting or excess reactants, and then predict the mass or number of moles of reactants required or products formed.

What is stoichiometry?

For a chemical reaction to occur, the reactants must combine in the correct ratios. The use of mathematical calculations to determine the relative quantities of reactants or products formed in a chemical reaction is called **stoichiometry**. These quantities are usually presented as a ratio of whole numbers (the **stoichiometric ratio**).

stoichiometry
the numerical relationship between the amount of reactants used and the products produced in a reaction; the application of the numerical relationship in calculations

stoichiometric ratio
the whole number ratio between each reactant and product in a reaction

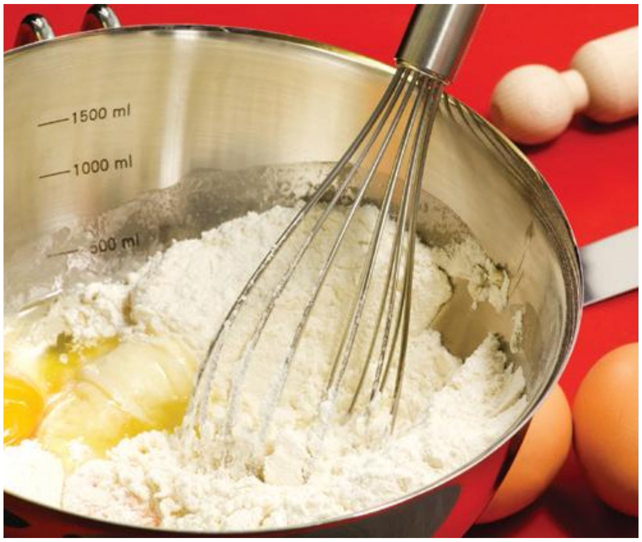


FIGURE 1 The correct ratio of flour, butter, eggs and milk (the reactants) is required for a cake (the product) to be edible.

How is stoichiometry used to calculate mass of reactants and products?

Being able to calculate the mass of a given reactant or product when the mass of another reactant or product is known is very important in chemistry. It helps guide decisions such as how much of a certain reactant is required in a mixture. It also tells us how much product can be formed.

The key to these calculations is the ratio between each reactant and product in a balanced equation, call the stoichiometric ratio. This is the ratio of molecules/formula units, and also as the ratio of moles of each reactant and product.

Let's consider the burning of propane, used in gas barbecues. It can be thought of in terms of molecules, but also in terms of moles.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

- **One molecule** of propane reacts with **5 molecules** of oxygen to produce **3 molecules** of carbon dioxide and **4 molecules** of water.
- **One mole** of propane reacts with **5 moles** of oxygen to produce **3 moles** of carbon dioxide and **4 moles** of water.

The stoichiometric ratio is $C_3H_8 : O_2 : CO_2 : H_2O = 1 : 5 : 3 : 4$. This is also called the **mole ratio**.

mole ratio
the ratio of the amounts in moles of the reactants and products in a reaction

Regardless of whether you are calculating the quantity (amount or mass) of reactant or product, the same general steps are followed:

1 Start with a balanced chemical equation for the reaction.

2 Extract all of the information from the question, i.e. what values are provided and what do you need to calculate?

3 Use the given mass of reactant/product to convert to moles.

4 Use the mole ratio from the balanced equation to calculate the amount (moles) of any other reactant or product as required:

$$\text{mole ratio} = \frac{\text{coefficient in front of unknown substance}}{\text{coefficient in front of known substance}}$$

$$n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$$

5 If required, convert back to mass for the desired reactant or product.

In QCE Chemistry, you will also work with quantities such as volumes of solution of known or unknown concentrations, and with gases, whose amounts can be found from pressure, volume and temperature.

Study tip

When using the coefficients of a chemical equation to find mole ratio, make sure the chemical equation is balanced.

Worked example 7.4A

Calculating the mass of a substance using stoichiometry

Worked example 7.4A: Watch a video that shows how to solve this problem.

Ethanol burns with oxygen to form carbon dioxide and water vapour according to the balanced equation:

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

Calculate the mass of carbon dioxide produced when 5.00 g of ethanol is burnt. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Calculate” means to determine or find a number or answer by using mathematical processes. We need to use stoichiometry to find the mass of a product. The question is worth 4 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Start with the balanced chemical equation for the reaction. Show the mole ratio above it.	1 : 3 : 2 : 3 $CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
Step 3: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ For CH_3CH_2OH: $n = ?$, $m = 5.00\,g$, $M = 46.08\,g\,mol^{-1}$ For CO_2: $n = ?$, $m = ?$, $M = 44.01\,g\,mol^{-1}$
Step 4: Substitute the known values into the formula and solve for the amount (in mol) of ethanol.	$n(CH_3CH_2OH) = \frac{5.00\,g}{46.08\,g\,mol^{-1}}$ (1 mark) $= 0.1085\,mol$ (1 mark)
Step 5: Use the mole ratio to find the amount (in mol) of carbon dioxide. mole ratio $= \frac{\text{coefficient in front of unknown substance}}{\text{coefficient in front of known substance}}$	$n(CO_2) = \frac{2}{1} \times n(CH_3CH_2OH) = \frac{2}{1} \times 0.105$ $= 0.2170\,mol$
Step 6: The question asks for mass, so we must convert the amount (in mol) of carbon dioxide to mass.	$0.2170\,mol = \frac{m(CO_2)}{44.01\,g\,mol^{-1}}$ (1 mark) $m(CO_2) = 9.55\,g$ (1 mark)

Your turn

Calculate the mass of oxygen required to burn 50.0 g of ethanol according to the chemical reaction equation above. (4 marks)

Study tip

It helps to include chemical formulas in brackets when performing calculations. This communicates your work clearly, and helps you keep track of what you have found at each step. It is highly recommended that you show units in your calculations to check that they will give you the correct units for your answers.

How is stoichiometry used to determine limiting reactants?

Consider the cake in Figure 1. It is important that there is enough of each ingredient. If there is a lot of flour, but limited amounts of butter, then you cannot produce a full-sized cake. Having more flour available will not help make the cake bigger. Instead, the excess flour will be left over at the end. This cake mixture analogy is similar to the limitations that occur in chemical reactions. When one reactant is limited, the amount of product will also be limited.

If one reactant is limited and used completely (the **limiting reactant**), then the reaction will cease and no more product will be produced. Any left-over reactants are called **excess reactants**.

limiting reactant
the reactant that is totally consumed when a chemical reaction is complete; this limits the amount of product that can be produced

excess reactant
the reactant that is not totally consumed by a chemical reaction

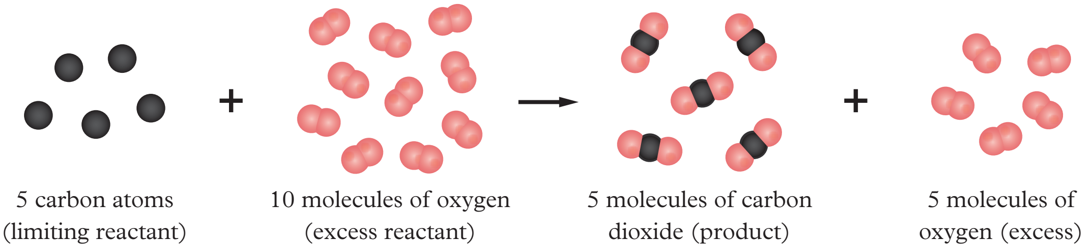


FIGURE 2 The five black C atoms in this reaction are the limiting reactant, because no more product can be produced once the limiting C atoms have all been used in the reaction.

To find the limiting reactant:

1. Start with a balanced chemical equation for the reaction.
2. Extract all of the information from the question, i.e. what values are provided and what do you need to calculate?
3. Use the given mass of reactant/product to convert to moles.
4. Use the stoichiometric coefficients to calculate $\frac{n}{\text{coefficient}}$ for each reactant. The reactant with the smallest $\frac{n}{\text{coefficient}}$ is the limiting reactant.

Once you have determined the limiting reactant, you can then calculate the quantity of the desired reactant or product. For these calculations, you must use the amount (in mol) of the limiting reactant.

Worked example 7.4B

Determining the limiting reactant in a reaction

Worked example 7.4B: Watch a video that shows how to solve this problem.

Nitrogen reacts with hydrogen to produce ammonia, as in the reaction equation shown.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

100.0 g of nitrogen and 25.0 g of hydrogen are mixed together.

a **Determine** which is the limiting reactant. (5 marks)

b **Calculate** the mass of ammonia formed. (3 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. "Calculate" means to determine or find a number or answer by using mathematical processes. We need to use stoichiometry to find the limiting reactant, then calculate the quantity of product formed. The questions are worth a variety of marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Start with the balanced chemical equation for the reaction. Show the mole ratio above it.	1 : 3 : 2 $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
Step 3: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ For N_2: n = ?, m = 100.0 g, M = 28.02 g mol^{-1} For H_2, n = ?, m = 25.0 g, M = 2.02 g mol^{-1} For NH_3, n = ?, m = ?, M = 17.04 g mol^{-1}
Step 4: Substitute the known values into the formulas and solve for the amounts (in mol) of the reactants.	$n(N_2) = \frac{100.0\,\text{g}}{28.02\,\text{g mol}^{-1}}$ (1 mark) = 3.5689 mol (1 mark) $n(H_2) = \frac{25.0\,\text{g}}{2.02\,\text{g mol}^{-1}}$ (1 mark) = 12.376 mol (1 mark)
Step 5: For part **a**, use $\frac{n}{\text{coefficient}}$ to find the limiting reactant (the smallest $\frac{n}{\text{coefficient}}$).	$\frac{n}{\text{coefficient}}$ for $N_2 = \frac{3.5689\,\text{mol}}{1}$ = 3.5689 mol $\frac{n}{\text{coefficient}}$ for $H_2 = \frac{12.376\,\text{mol}}{3}$ = 4.1254 mol The $\frac{n}{\text{coefficient}}$ for N_2 is smaller, so it is the limiting reactant. (1 mark) H_2 is in excess.

Think	Do
Step 6: For part **b**, use the amount (in mol) of the limiting reactant and the mole ratio to find the amount (in mol) of ammonia.	The mole ratio uses the coefficients for N_2 (the limiting reactant) and NH_3 (the product). $n(NH_3) = \frac{2}{1} \times n(N_2)$ $= 7.1378\,mol$
Step 7: The question asks for mass, so we must convert the amount (in mol) of ammonia to mass.	$7.1378\,mol = \frac{m}{17.04\,g\,mol^{-1}}$ (1 mark) $m = 122\,g$ (1 mark)

Your turn

1.00 g of potassium iodide and 1.00 g of lead(II) nitrate are separately dissolved in water, and then mixed, reacting to form insoluble lead(II) iodide according to the equation:

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

a **Determine** which is the limiting reactant. (5 marks)

b **Calculate** the mass of lead iodide formed. (2 marks)

Study tip

Store your calculated values in your calculator and do not round until the very last step.

Study tip

Not all questions will tell you to find the limiting reactant first. Recognise questions that involve limiting reactants by the fact that quantities of two reactants are given, rather than one reactant.

Real-world chemistry

Combustion reactions in Formula One

The chemical reaction that occurs between the fuel and oxygen in a Formula One car engine demonstrates the idea that if one reactant is limited, the amount of product will also be limited. The amounts of fuel and oxygen must be in the correct ratio for the chemical reaction to occur efficiently. Too little of either reactant results in a smaller reaction and less energy. The balanced combustion reaction of octane can be used as an example.

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

The reaction can be described as 2 mol of octane reacting with 25 mol of oxygen to produce 16 mol of carbon dioxide and 18 mol of water. The mole ratio is 2 : 25 : 16 : 18.

If there is a limited amount of one reactant, then only limited amounts of the product can be produced. For example, if there is a lot of oxygen available but only a limited amount of octane (0.5 mol), then only limited amounts of carbon dioxide (4 mol) or water (4.5 mol) can be produced. This results in less energy being produced in the reaction. Adding more oxygen into the mixture will not change the amount of carbon dioxide or water being produced or produce more power in the motor. The oxygen is already in excess.

Injecting more octane fuel with the oxygen so that both reactants are in the correct ratio will allow the reaction to occur most efficiently.

Over the course of decades, designers of Formula One vehicles have improved the ability for these high-powered vehicles to go faster in many ways. One strategy involved replacing the older naturally aspirated engines with turbo-charged engines. In a naturally aspirated engine, air direct from the atmosphere enters the cylinder during the engine stroke. In a turbo-charged engine, air is forced into the engine under pressure, so has the potential to provide a larger and better supply of oxygen throughout the cylinder.

Apply your understanding

1 **Describe** the types of tests that designers could carry out on the combustion of fuel in a Formula One vehicle to optimise energy production. (1 mark)

2 **Propose** two hypotheses that designers could have considered when trying to improve the older naturally-aspirated engines. (6 marks)

3 **Consider** the role of fuel injectors and **propose** reasons why the fineness of the fuel mist, and the precision of timing would be important. (2 marks)

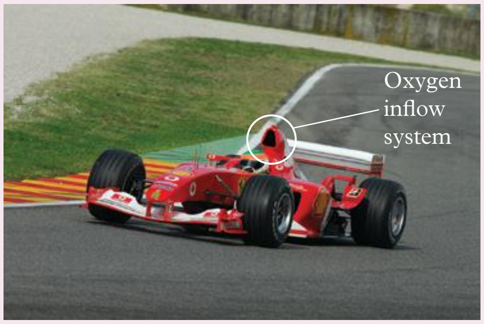


FIGURE 3 Many Formula One vehicles have systems that increase air flow into the engine to ensure the correct ratio of fuel to oxygen for an efficient combustion reaction.

Check your learning 7.5

Check your learning 7.5: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define**:
  a mole ratio (1 mark)
  b limiting reactant (1 mark)
  c excess reactant. (1 mark)

2 **Consider** the reaction:

$$Ba_3N_2 + 6H_2O \rightarrow 3Ba(OH)_2 + 2NH_3$$

Calculate the mass of NH_3 produced if 220 g of barium nitride (Ba_3N_2) reacts with excess water. (4 marks)

3 If 120 g of propane (C_3H_8) is burnt in excess oxygen, **calculate** the mass of water produced. (4 marks)

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Analytical processes

4 90.00 g of iron(II) chloride ($FeCl_2$) reacts with 52.00 g of hydrogen sulfide (H_2S) to produce solid iron(III) sulfide (Fe_2S_3) and hydrochloric acid (HCl).
  a **Construct** a balanced chemical equation for the reaction. (1 mark)
  b **Determine** the limiting reactant. (5 marks)
  c **Calculate** the mass of iron(II) sulfide produced in the reaction. (2 marks)

5 Rescue workers use self-contained breathing apparatus to remove carbon dioxide and water from their exhaled breath. First, potassium oxide removes water from the exhaled breath:

$$4KO_2 + 2H_2O \rightarrow 3O_2 + 4KOH$$

The potassium hydroxide (KOH) produced is then available to react with the carbon dioxide:

$$KOH + CO_2 \rightarrow KHCO_3$$

  a **Calculate** the mass of potassium super oxide needed to produce 235 g of oxygen. (4 marks)
  b If 123 g of KO_2 is available in the breathing apparatus, **determine** what mass of carbon dioxide can be removed from the exhaled air. (4 marks)

FIGURE 4 Rescue workers use breathing apparatus.

Learning intentions and success criteria

Video demonstration

Practical

Lesson 7.6

Investigating limiting reactants

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 7.7

Yield

Learning intentions and success criteria

Key ideas

→ Experimental (actual) yield can vary from the theoretical (expected) yield.

→ The percentage yield is calculated from experimental or given data.

theoretical yield the amount of product predicted from stoichiometric calculations

experimental yield the actual amount of product that is produced in a chemical reaction

What is yield?

Chemists measure the efficiency of chemical reactions by comparing the predicted amount of product from their stoichiometric calculations (**theoretical yield**) to the actual amount of product they achieve from the chemical reaction (**experimental yield**).

If the theoretical yield and the experimental yield are identical, then the experiment is considered 100% efficient. But chemical reactions rarely work perfectly to give 100% yield. This is because many factors can affect the reaction and reduce the yield, such as impure reactants, other side reactions occurring or lost products through reactants or products remaining on the sides of containers. An added complication is that very few reactions go to completion.

FIGURE 1 Laboratory scales being used to weigh a product.

The percentage yield can be calculated by the following equation:

$$\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

The percentage yield can never be higher than 100% because it would mean new atoms had been created in the chemical reaction. This would break the law of conservation of mass.

Worked example 7.5A

Determining experimental and percentage yield

Worked example 7.5A: Watch a video that shows how to solve this problem.

When 39.75 g of magnesium carbonate decomposes in an experiment, 15.00 g of magnesium oxide is formed. **Determine** the percentage yield of magnesium oxide. (6 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to find the theoretical yield, and compare this to the experimental yield to find the percentage yield. The question is worth 6 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Start with a balanced chemical equation for the reaction. Show the mole ratio above it.	1 : 1 : 1 $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$
Step 3: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ For $MgCO_3$: n = ?, m = 39.75 g, M = 84.32 g mol^{-1} For MgO: n = ?, m (or experimental yield) = 15.00 g, M = 40.31 g mol^{-1} $\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$ percentage yield = ?, theoretical yield = ?
Step 4: Substitute the known values into the formula and solve for the amount of magnesium carbonate.	$n(MgCO_3) = \frac{39.75\,g}{84.32\,g\,mol^{-1}}$ (1 mark) = 0.4714 mol (1 mark)
Step 5: Use the mole ratio to calculate the theoretical number of moles of magnesium oxide.	$n(MgO) = \frac{1}{1} \times n(MgCO_3)$ = 0.4714 mol
Step 6: Substitute the known values into the formula and solve for the theoretical mass of magnesium oxide.	$0.4714\,mol = \frac{m}{40.31\,g\,mol^{-1}}$ (1 mark) $m = 0.4714\,mol \times 40.31\,g\,mol^{-1}$ = 19.00 g (1 mark)
Step 7: Substitute the known values into the formula and solve for the percentage yield.	$\text{percentage yield} = \frac{15.00\,g}{19.00\,g} \times \frac{100}{1}$ (1 mark) = 78.94% (1 mark)

Your turn

Sulfuric acid is manufactured using the contact process. The first step involves reacting sulfur dioxide gas with oxygen gas to produce sulfur trioxide. For every 1,000 g of sulfur dioxide used, 1,200 g of sulfur trioxide is produced. **Determine** the percentage yield of sulfur trioxide. (6 marks)

Check your learning 7.7

Check your learning 7.7: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** the difference between theoretical yield and experimental yield. (2 marks)

2 **Explain** why a percentage yield of 109% is impossible. (2 marks)

Analytical processes

3 Magnesium oxide is formed when magnesium ribbon burns in the presence of oxygen:

$$2Mg + O_2 \rightarrow 2MgO$$

  a **Determine** the theoretical yield of MgO formed when 0.12 g of Mg reacts with excess oxygen. (4 marks)

  b When a student performed this experiment, they found 0.18 g of MgO was formed. **Determine** the percentage yield of this reaction. (2 marks)

4 Heptane undergoes a combustion reaction with oxygen to form carbon dioxide and water:

$$C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$$

  a **Determine** the theoretical yield of CO_2 formed when 10.0 g of C_7H_{16} is burnt with excess O_2. (4 marks)

  b A chemist analysing this reaction produced 25 g of carbon dioxide. **Determine** the percentage yield for this reaction. (2 marks)

Practical

Lesson 7.8

Investigating the percentage yield of copper carbonate

Learning intentions and success criteria

Video demonstration

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

MODULE 7

Lesson 7.9

Review: Mole concept and law of conservation of mass

Summary

7.1
- A mole is a quantity of matter that is equal to Avogadro's number of particles (6.02×10^{23}).

7.2
- The law of conservation of mass states that mass is neither created nor destroyed in a chemical reaction.
- The molar mass is the mass of one mole of a substance.
- The mole is equivalent to the mass of a substance divided by its molar mass.

7.3
- The percentage composition of a compound is determined from the relative atomic masses of each element.
- The empirical formula is the simplest whole-number ratio of elements in a compound. It is determined from percentage composition.
- The molecular formula of a compound can be determined from its empirical formula and the molar mass.

7.4
- Practical: Determining the empirical formula of magnesium oxide.

7.5
- The mole concept and the mole ratio in chemical equations can be used to calculate the mass of reactants/products involved in a reaction. These types of calculations are called stoichiometry.
- Stoichiometric ratios can be used to determine the limiting or excess reactants, and then predict the mass or number of moles of reactants required or products formed.

7.6
- Practical: Investigating limiting reactants.

7.7
- Experimental (actual) yield can vary from the theoretical (expected) yield.
- The percentage yield is calculated from experimental or given data.

7.8
- Practical: Investigating the percentage yield of copper carbonate.

Key formulas

Number of molecules	$n = \frac{N}{N_A}$
Number of moles	$n = \frac{m}{M}$
Percentage composition	$\%\text{ of } x \text{ in a compound} = \frac{\text{mass of } x \text{ in 1 mol of a compound}}{\text{molar mass of compound}} \times \frac{100}{1}$
Multiple of empirical formula	$\text{Multiple of empirical formula} = \frac{M(\text{molecular formula})}{M(\text{empirical formula})}$
Mole ratio	$\text{mole ratio} = \frac{\text{coefficient in front of unknown substance}}{\text{coefficient in front of known substance}}$
Limiting reactant	$\frac{n}{\text{coefficient}}$
Percentage yield	$\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$

Review questions 7.9A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 1 mol of CH_4O contains

A 1 mol of hydrogen.

B 2 mol of hydrogen.

C 3 mol of hydrogen.

D 4 mol of hydrogen.

2 0.123 mol of methane (CH_4) contains

A 5 molecules.

B 2.04×10^{-25} molecules.

C 2.46×10^{-2} molecules.

D 7.40×10^{22} molecules.

3 The combustion of propane (C_3H_8) in excess oxygen produces carbon dioxide and water.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

If 3.4 mol of oxygen is used in this reaction, the number of moles of carbon dioxide produced would be

A 2.0 mol.

B 3.4 mol.

C 0.68 mol.

D 2.7 mol.

4 The smallest number of molecules would be found in

A 0.5 g of CH_3OH.

B 0.5 g of N_2.

C 0.5 g of NH_3.

D 0.5 g of C_2H_2.

5 A forensic scientist analyses an unknown compound and finds that it contains 72.7% oxygen and 27.3% carbon. The empirical formula of the compound could be

A CO.

B CO_2.

C CO_3.

D C_7O_2.

6 Sulfuric acid reacts with potassium hydroxide according to the equation:

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

Which of the following is correct?

A $n(H_2SO_4) = \frac{2}{1} \times n(KOH)$

B $n(H_2SO_4) = \frac{1}{1} \times n(KOH)$

C $n(H_2SO_4) = \frac{1}{2} \times n(KOH)$

D $n(H_2SO_4) = \frac{2}{1} \times n(K_2SO_4)$

7 The molar mass of glucose ($C_6H_{12}O_6$) is

A 180.18 g mol^{-1}.

B 246.78 g mol^{-1}.

C 29.07 g mol^{-1}.

D 180.00 g mol^{-1}.

8 Hydrogen and oxygen react to form water as shown:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

The mass ratio closest to the actual stoichiometric ratio is shown in which of the following mixtures?

A 2 g of hydrogen and 1 g of oxygen

B 2 g of hydrogen and 16 g of oxygen

C 2 g of hydrogen and 18 g of oxygen

D 2 g of hydrogen and 32 g of oxygen

9 Which one of the following contains the most atoms?

A 1 mol of $CuSO_4{\cdot}5H_2O$

B 1 mol of $(NH_4)_2CO_3$

C 1 mol of $CH_3(CH_2)_2CH_2OH$

D 1 mol of $C_6H_{12}O_6$

10 The empirical formula of a compound is CH_2O. Which of the following statements must be true.

A The molecular formula is $C_2H_4O_2$.

B There will be a minimum of 2 carbon atoms in the molecule.

C The molar mass is a whole number multiple of 4.

D The molar mass is a whole number multiple of 30.03.

Review questions 7.9B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

11 **Identify** the symbol used to write the number of moles of carbon dioxide. (1 mark)

12 **Identify** the factors that can prevent the experimental yield being equivalent to the theoretical yield. (2 marks)

13 **Calculate** the number of atoms of magnesium in 24.31 g of Mg. (4 marks)

14 **Calculate** the molar mass of:

- **a** caffeine ($C_8H_{10}N_4O_2$) (1 mark)
- **b** dinitrogen tetroxide (N_2O_4) (1 mark)
- **c** calcium hydroxide $Ca(OH)_2$ (1 mark)
- **d** sodium hydrogen carbonate $NaHCO_3$. (1 mark)

15 **Calculate** the number of hydrogen atoms present in 43.8 g of urea ($(NH_2)_2CO$). (3 marks)

16 **Identify** an example of a substance that has the same empirical formula and molecular formula. **Explain** your answer. (2 marks)

17 **Calculate** each of the following found in 3.5 moles of calcium ethanoate ($Ca(CH_3COO)_2$).

- **a** Moles of calcium atoms (1 mark)
- **b** Moles of carbon atoms (2 marks)
- **c** Moles of hydrogen atoms (2 marks)
- **d** Moles of oxygen atoms (2 marks)
- **e** Mass of hydrogen (2 marks)
- **f** Mass of oxygen (2 marks)

18 Most living things use glucose ($C_6H_{12}O_6$) and oxygen to produce carbon dioxide, water and energy. This reaction can be represented by the equation:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

- **a** If 968 g of glucose was used, **calculate** the mass of carbon dioxide produced. (4 marks)
- **b** **Calculate** the number of grams of oxygen needed to make 550 g of carbon dioxide. (4 marks)

19 Copper(I) sulfide is roasted in air to react with oxygen, in a series of steps which can be summarised into the following chemical equation. Consider 1.000 kg (or 1,000 g) of copper(I) sulfide.

$$Cu_s2(s) + O_2(g) \rightarrow 2Cu(s) + SO_2(g)$$

- **a** **Calculate** the mass of copper that can be obtained. (4 marks)
- **b** **Calculate** the mass of sulfur dioxide produced. (4 marks)

20 Urea ($(NH_2)_2CO$) is a fertiliser used by many farmers to replace nitrogen in the soil. It can be produced by reacting ammonia (NH_3) with carbon dioxide:

$$2NH_3 + CO_2 \rightarrow (NH_2)_2CO + H_2O$$

If 849.2 g of ammonia is mixed with 1,223 g of carbon dioxide and allowed to react to completion, **calculate** how much urea is formed. (7 marks)

FIGURE 1 Farmer applying poultry manure to pasture, which provides urea fertiliser

Analytical processes

21 **Discriminate** between empirical and molecular formulae of compounds and give an example to **justify** your response. (2 marks)

22 The amino acid serine has the percentage composition 34.29% carbon, 6.71% hydrogen, 45.67% oxygen and 13.33% nitrogen. If the molar mass is 105.9 g mol^{-1}, **determine** the:

- **a** empirical formula of serine. (4 marks)
- **b** molecular formula of serine. (4 marks)

23 Phosphorus reacts with a limited amount of oxygen to form tetraphosphorus hexoxide:

$$P_4 + 3O_2 \rightarrow P_4O_6$$

If enough oxygen is available, then the tetraphosphorus hexoxide reacts further to produce phosphorus pentoxide:

$$P_4O_6 + 2O_2 \rightarrow P_4O_{10}$$

a If a container holds 5.77 g of phosphorus and 57.7 g of oxygen, **determine** the amount, in mol, of oxygen that will be in excess in the first reaction. (5 marks)

b **Use** your answer from part **a** to **determine** the mass of tetra phosphorus decaoxide that will be produced. (2 marks)

24 **Determine** the:

a empirical formula of a molecule containing 41.40% carbon, 3.40% hydrogen and 55.20% oxygen (4 marks)

b molecular formula of a compound with a molar mass of 136 g mol^{-1} and an empirical formula of C_4H_4O (4 marks)

c empirical formula of ethanoic acid given that it has a percentage composition of 39.9% C, 6.7% H and 53.4% O (4 marks)

d percentage composition by mass of carbon, nitrogen and oxygen in ethanamide (C_2H_5NO). (6 marks)

25 The strength of eggshell depends on the amount of calcium carbonate ($CaCO_3$) in the eggshell. The percentage of calcium carbonate in the eggshell can be determined by gravimetric analysis. A total of 0.412 g of clean dry eggshell is completely dissolved in a minimum volume of dilute hydrochloric acid:

$$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$$

An excess of a basic solution of ammonium oxalate ($(NH_4)_2C_2O_4$) is then added to form crystals of calcium oxalate monohydrate ($CaC_2O_4{\cdot}H_2O$). The suspension is filtered and the crystals are dried to constant mass. A total of 0.523 g of $CaC_2O_4{\cdot}H_2O$ is collected.

$$Ca^{2+}(aq) + (NH_4)_2C_2O_4(aq) + H_2O(l) \rightarrow CaC_2O_4{\cdot}H_2O(s) + 2NH_4^+(aq)$$

Determine the percentage, by mass, of calcium carbonate in the eggshell. (6 marks)

26 Fool's gold is a common iron ore that contains the mineral iron pyrite (FeS_2). Typically, the percentage by mass of FeS_2 in a sample of fool's gold is 90–95%. The actual percentage in a sample can be determined experimentally. The sulfur in FeS_2 is converted to sulfate (SO_4^{2-}), which is then mixed with an excess of barium chloride ($BaCl_2$) to form barium sulfate ($BaSO_4$) according to the equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

When the reaction has gone to completion, the $BaSO_4$ precipitate is collected in a filter paper and carefully washed. The filter paper and its contents are then transferred to a crucible. The crucible and its contents are heated until constant mass is achieved. The data for an analysis of a mineral sample is shown in Table 1.

TABLE 1 Results from analysis of minerals

Initial mass of mineral sample (g)	18.88
Mass of crucible and filter paper (g)	123.40
Mass of crucible, filter paper and dry $BaSO_4$ (g)	174.99
$M(FeS_2)$ (g mol^{-1})	120.0
$M(BaCl_2)$ (g mol^{-1})	208.23
$M(BaSO_4)$ (g mol^{-1})	233.39

Analyse the data to **determine** the percentage by mass of FeS_2 in this mineral sample. (9 marks)

27 Antimony (Sb) is used in the electronics industry to make semiconductor devices, such as infrared detectors. It is obtained by heating stibnite (Sb_2S_3) with scrap iron and drawing off the molten antimony:

$$2Sb_2S_3(s) + 3Fe(s) \rightarrow 2Sb(s) + 3FeS(s)$$

If 300 g of stibnite and 125 g of iron are heated together to give 100 g of Sb, **determine** the:

a limiting reactant (5 marks)

b mass of the excess reactant remaining (3 marks)

c theoretical yield of antimony (3 marks)

d percentage yield of antimony. (2 marks)

Data drill

Limiting reactant graphs

The reaction for zinc (Zn) reacting with hydrochloric acid (HCl) is:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2\,(aq) + H_2\,(g)$$

The mass of hydrogen gas is graphed against the volume of HCl added, as shown in Figure 1.

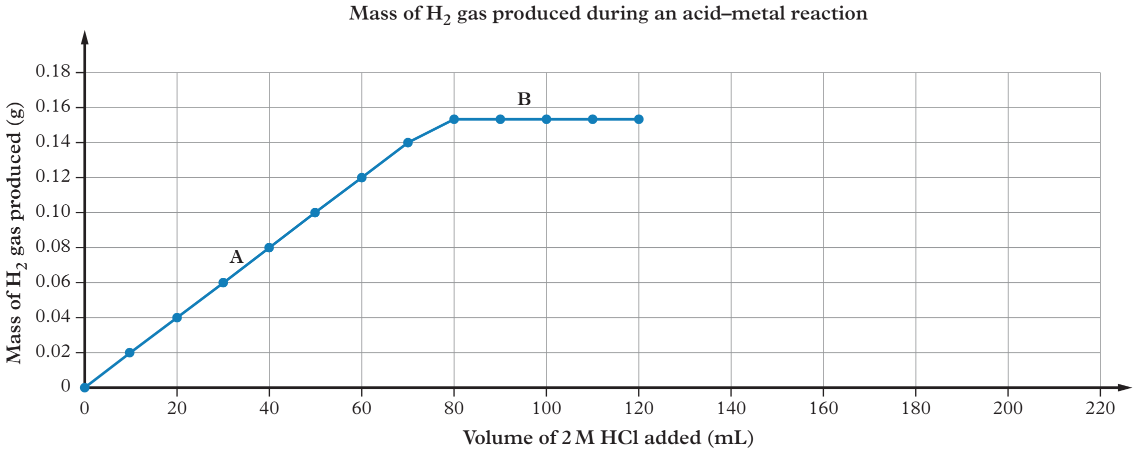


FIGURE 1 Mass of H_2 gas produced when 5 g of zinc metal is reacted with 2 M HCl.

Apply understanding

1 **Identify** which reactant is the limiting reagent and which reagent is in excess in each region (A and B) of the graph in Figure 1. (2 marks)

2 **Identify** the trend in region A of the graph. (1 mark)

Analyse data

3 **Compare** the mass of hydrogen produced after the addition of 40 mL of HCl with the mass of hydrogen after 80 mL. **Identify** any limitation or uncertainty. (3 marks)

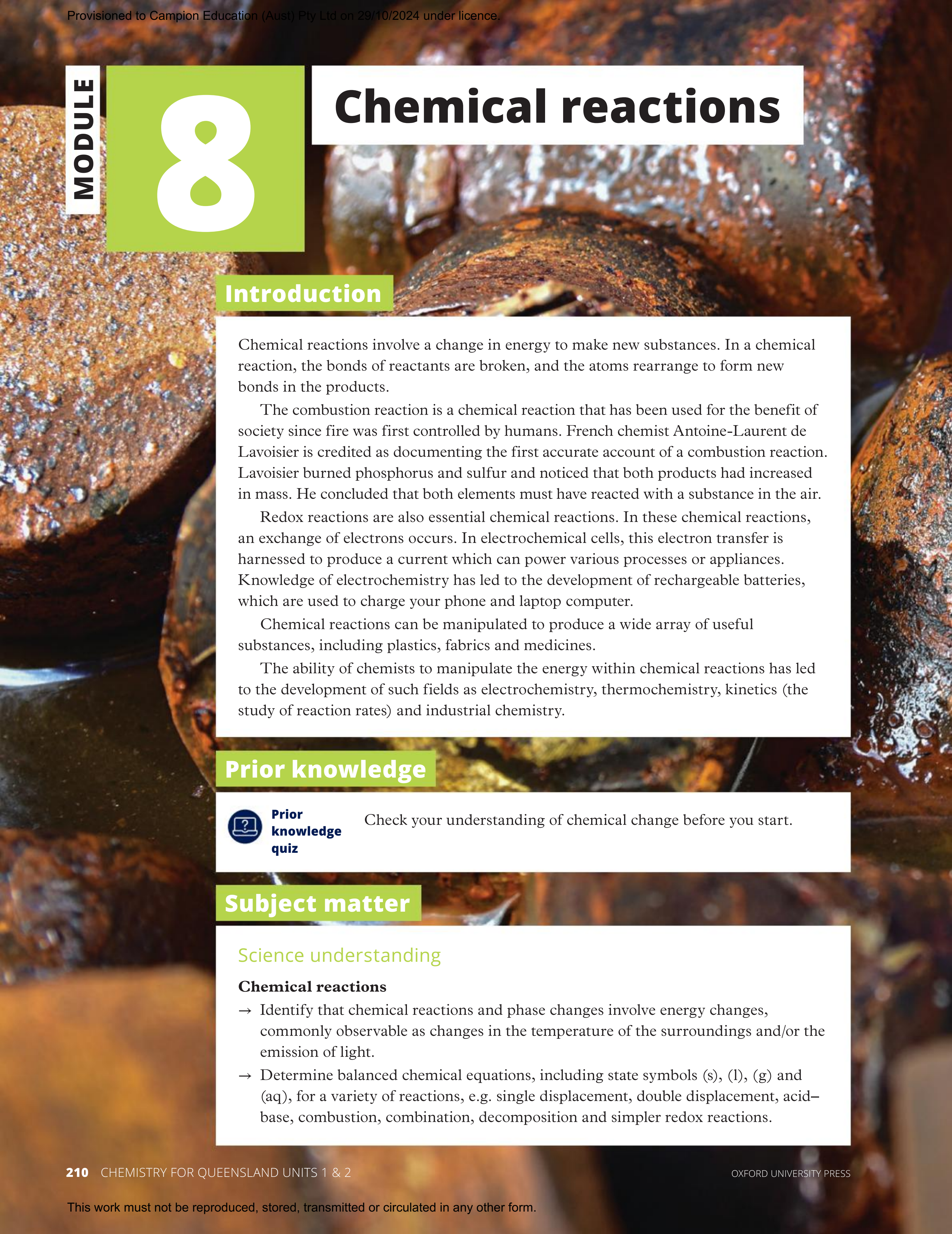
Evaluate evidence

4 **Extrapolate** the data in Figure 1 and **determine** how many grams of H_2 will be produced when 200 mL of HCl is added to 5 g of zinc. (2 marks)

5 **Predict** the trend if 2.5 g of zinc had been used, instead of 5 g, and represent this by adding a curve to the graph. (3 marks)

Module 7 checklist: Mole concept and law of conservation of mass

Quizlet: Revise key terms online to test your understanding



MODULE 8

Chemical reactions

Introduction

Chemical reactions involve a change in energy to make new substances. In a chemical reaction, the bonds of reactants are broken, and the atoms rearrange to form new bonds in the products.

The combustion reaction is a chemical reaction that has been used for the benefit of society since fire was first controlled by humans. French chemist Antoine-Laurent de Lavoisier is credited as documenting the first accurate account of a combustion reaction. Lavoisier burned phosphorus and sulfur and noticed that both products had increased in mass. He concluded that both elements must have reacted with a substance in the air.

Redox reactions are also essential chemical reactions. In these chemical reactions, an exchange of electrons occurs. In electrochemical cells, this electron transfer is harnessed to produce a current which can power various processes or appliances. Knowledge of electrochemistry has led to the development of rechargeable batteries, which are used to charge your phone and laptop computer.

Chemical reactions can be manipulated to produce a wide array of useful substances, including plastics, fabrics and medicines.

The ability of chemists to manipulate the energy within chemical reactions has led to the development of such fields as electrochemistry, thermochemistry, kinetics (the study of reaction rates) and industrial chemistry.

Prior knowledge

Prior knowledge quiz Check your understanding of chemical change before you start.

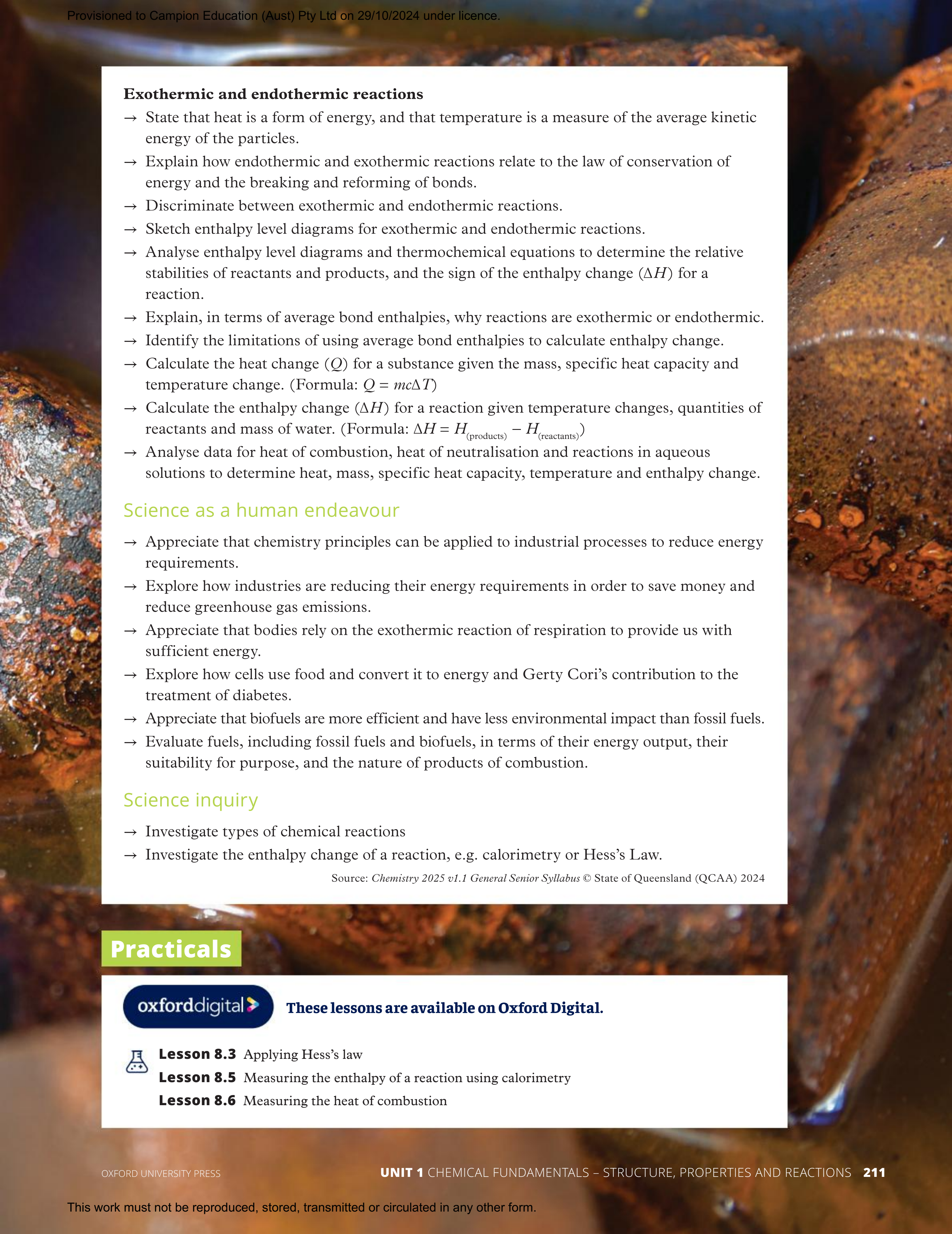
Subject matter

Science understanding

Chemical reactions

→ Identify 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.

→ Determine balanced chemical equations, including state symbols (s), (l), (g) and (aq), for a variety of reactions, e.g. single displacement, double displacement, acid–base, combustion, combination, decomposition and simpler redox reactions.



Exothermic and endothermic reactions

- → State that heat is a form of energy, and that temperature is a measure of the average kinetic energy of the particles.
- → Explain how endothermic and exothermic reactions relate to the law of conservation of energy and the breaking and reforming of bonds.
- → Discriminate between exothermic and endothermic reactions.
- → Sketch enthalpy level diagrams for exothermic and endothermic reactions.
- → Analyse enthalpy level diagrams and thermochemical equations to determine 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.
- → Identify the limitations of using average bond enthalpies to calculate enthalpy change.
- → Calculate the heat change (Q) for a substance given the mass, specific heat capacity and temperature change. (Formula: $Q = mc\Delta T$)
- → Calculate the enthalpy change (ΔH) for a reaction given temperature changes, quantities of reactants and mass of water. (Formula: $\Delta H = H_{(\text{products})} - H_{(\text{reactants})}$)
- → Analyse data for heat of combustion, heat of neutralisation and reactions in aqueous solutions to determine heat, mass, specific heat capacity, temperature and enthalpy change.

Science as a human endeavour

- → Appreciate that chemistry principles can be applied to industrial processes to reduce energy requirements.
- → Explore how industries are reducing their energy requirements in order to save money and reduce greenhouse gas emissions.
- → Appreciate that bodies rely on the exothermic reaction of respiration to provide us with sufficient energy.
- → Explore how cells use food and convert it to energy and Gerty Cori's contribution to the treatment of diabetes.
- → Appreciate that biofuels are more efficient and have less environmental impact than fossil fuels.
- → Evaluate fuels, including fossil fuels and biofuels, in terms of their energy output, their suitability for purpose, and the nature of products of combustion.

Science inquiry

- → Investigate types of chemical reactions
- → Investigate the enthalpy change of a reaction, e.g. calorimetry or Hess's Law.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

oxforddigital **These lessons are available on Oxford Digital.**

Lesson 8.3 Applying Hess's law

Lesson 8.5 Measuring the enthalpy of a reaction using calorimetry

Lesson 8.6 Measuring the heat of combustion

Lesson 8.1
Physical and chemical changes

Key ideas

→ Phase changes are physical changes that involve changes in energy.

→ Chemical reactions also involve energy changes, and include reactions like single displacement, double displacement, acid–base, combustion, combination, decomposition and simple redox reactions.

→ Chemical reactions can be represented using balanced chemical equations, including states.

Learning intentions and success criteria

What are physical and chemical changes?

Physical and chemical changes are influenced by changes in energy. The energy causes a change within a substance, which results in the formation of a new substance or a phase change. The most common form of energy involved in chemical reactions is thermal energy (heat), which is associated with the movement (kinetic energy) of particles.

FIGURE 1 Physical changes (such as the melting of ice) and chemical changes (such as the formation of nitrogen dioxide gas bubbles when iron reacts with nitric acid) involve changes in energy.

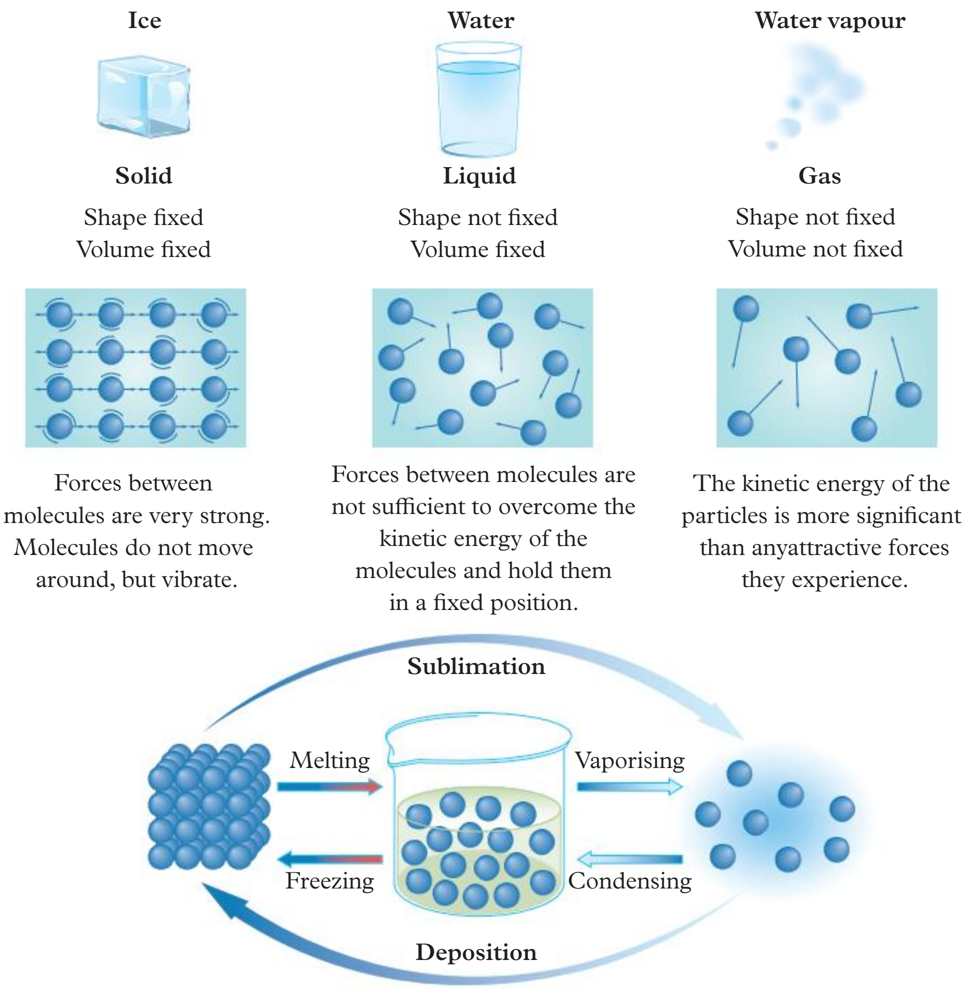
What happens during phase changes?

thermal energy the energy any substance has at a particular temperature

physical change when the molecules remain the same but the substance is in a different state that has different properties

Substances change between phases or states when **thermal energy** (heat) is added or removed. Phase changes are **physical changes**, not chemical changes, because molecules do not break apart to form new molecules. Instead, the molecules move in different ways to change states (or phases) between solid, liquid and gas.

Kinetic molecular theory helps to explain phase or state changes when substances are heated or cooled. This theory states that particles are in constant random motion. During changes of state, added heat energy causes the motion of particles to change significantly, but without any increase or decrease in temperature. Water is often used to demonstrate state changes because it transforms from solid to liquid to gas over a narrow temperature range, which can be easily observed and measured (Figure 2).



Forces between molecules are very strong. Molecules do not move around, but vibrate.

Forces between molecules are not sufficient to overcome the kinetic energy of the molecules and hold them in a fixed position.

The kinetic energy of the particles is more significant than anyattractive forces they experience.

FIGURE 2 The changes of state and movement of water molecules when energy is added or removed

Solid to liquid – melting

Water molecules have relatively strong forces attracting them to one another. In solid water (ice), water molecules form a lattice held together by attractive forces. The molecules vibrate within the lattice because they do not have enough energy to overcome the forces of attraction between the molecules. As thermal energy is added, in a process called **melting**, the molecules vibrate faster and start to move around. They move away from their lattice positions while still remaining close together, to become a liquid.

melting
the process of turning a solid into a liquid by applying thermal energy; particles move away from fixed positions, but are still close together

Liquid to gas – vaporisation

In liquid water, the molecules vibrate, rotate and move around, but are still close together. They fill the bottom of the container that they occupy and remain there until more energy is added. As thermal energy is added, in a process called **vaporisation**, the molecules move faster and start to overcome the attractive forces between molecules. The particles move apart, with empty spaces between them, to become a gas called water vapour. However, the molecules stay intact as no covalent bonds are broken.

vaporisation
the process of turning a liquid into a gas by applying thermal energy; particles move faster and separate from one another

Vaporisation includes both **boiling** and **evaporation**. Boiling occurs at the temperature where thermal energy overcomes the attractive forces between the molecules so that they can form bubbles of gas below the surface. Evaporation occurs at much lower temperatures and involves individual molecules escaping from the surface. It occurs much more slowly. As thermal energy increases, so does the rate of evaporation.

Gas to liquid – condensation

In the gaseous state (water vapour), the energy of movement of the water molecules is greater than the forces that attract the water molecules to each other. The molecules move about freely in a random motion. When heat is removed, the gaseous water molecules transfer their energy to the surroundings. Their energy decreases and the attractive forces between water molecules dominates again. The molecules come very close to each other and form a liquid. This process, known as **condensation**, occurs in clouds to form rain or precipitation.

Liquid to solid – freezing

As more thermal energy is removed from water molecules, they no longer have the energy to overcome the attractive forces between one another. They align back into the solid ice lattice. This process is known as **freezing**.

Solid to gas – sublimation

During the process of **sublimation**, molecules change from the solid state to the gaseous state, without going through the liquid phase. A few substances sublime directly from solid to gas, notably frozen carbon dioxide, known as dry ice. Iodine also sublimes to a purple vapour, which is highly toxic.

Gas to solid – deposition

Deposition occurs when molecules change straight from a gas to a solid, skipping the liquid phase. It is the opposite of sublimation. On cold mornings, you may notice that frost forms on the windows of a car, when there has been no rain overnight. When the outside temperature falls below the freezing point of water (0°C), water vapour in the air loses thermal energy and converts from a gas to a solid, forming ice on the car.

boiling
vaporisation at a substance's boiling point, which is the temperature at which a liquid forms bubbles of its own gas beneath the surface

evaporation
the phase change from a liquid to a gas

condensation
the process of turning a gas into a liquid by removing thermal energy; particles move slower and come close together

freezing
the process of turning a liquid into a solid by removing thermal energy; particles stop moving around and stay in fixed positions where they can only vibrate

sublimation
the process of turning a solid into a gas by applying thermal energy; particles move and separate from one another

deposition
the process of turning a gas into a solid by removing thermal energy; particles stop moving and become more attracted to one another

FIGURE 3 When water vapour comes into contact with a cool glass surface, it condenses into droplets on the glass.

FIGURE 4 Pellets of dry ice undergoing sublimation

How is thermal energy related to radiation?

Thermal energy is stored by the substance as motion of particles: vibrational (and occasionally rotational) energy in solids, and vibrational, rotational and translational energy in liquids and gases. Increasing the temperature results in more rapid motion.

During melting, added energy causes particles to move from fixed positions as they gain translational energy without an increase in temperature. During vaporisation, added energy causes particles to separate from each other, leaving empty space between them, without an increase in temperature. Condensation and freezing involve the loss of energy to the surroundings, and a decrease in particle motion. The changes in kinetic energy occur when thermal energy is transferred by physical contact between particles. This occurs through processes of **conduction** and **convection**.

Changes of state and chemical reactions all involve the addition or removal of thermal energy. Each change of state can be symbolised in an equation where energy is either added or removed. For example, the following equation represents energy being absorbed during vaporisation:

$$H_2O(l) + \text{energy} \rightarrow H_2O(g)$$

As well as changes in temperature, thermal energy can be transferred to or from substances by **thermal radiation**. It is a form of **electromagnetic radiation** which uses electromagnetic waves to transfer energy. Thermal radiation is emitted to the surroundings by all objects with a temperature above absolute zero (0 K) and can include visible light and infrared light. For example, when hot metal glows bright yellow–red, it emits visible light. Human beings emit infrared radiation, which can only be seen by infrared or thermal imaging.

An example of infrared imaging is shown in Figure 5. When the radiator heater is turned on, it generates thermal energy and emits more thermal radiation. Thermal radiation is then absorbed by the air particles within the room, they gain kinetic energy and their motion increases.

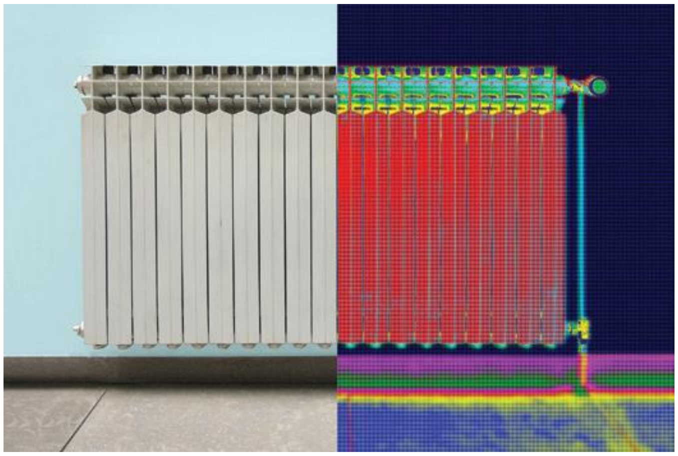


FIGURE 5 Compare the infrared image of the thermal radiation (heat energy) emitted from a radiator heater (right) with the real image of the radiator (left). Red regions are the warmest and blue regions are the coolest.

conduction
a process in which thermal energy is transferred between particles when they make direct contact with each other

convection
a process in which thermal energy is transferred between particles via the movement of liquid or gas particles

thermal radiation
a process in which thermal energy is transferred in the form of electromagnetic waves

electromagnetic radiation
energy consisting of electromagnetic waves

What happens during chemical reactions?

Chemical reactions occur when energy breaks apart the bonds within chemical reactants. The atoms then rearrange to form products. The energy that breaks the bonds may be in the form of thermal or light energy. Slow reactions speed up when heated, as the bonds are more easily broken. If electromagnetic radiation is involved in reactions, this usually requires high energy radiation, such as ultraviolet radiation, to provide enough energy to break the strong chemical bonds.

chemical reaction
when the bonds in a molecule break apart and atoms rearrange to form a new substance

balanced
having the same numbers of atoms of each element on either side of a reaction arrow

How do you write balanced chemical equations?

As the atoms rearrange, there must be the same number and type of atoms in the products as there are in the reactants. When writing chemical equations, it is essential that they are **balanced** and include the correct formulas and states. This helps you to correctly communicate chemical concepts and principles.

Study tip

You will need to write balanced chemical equations throughout Units 1–4 in QCE Chemistry.

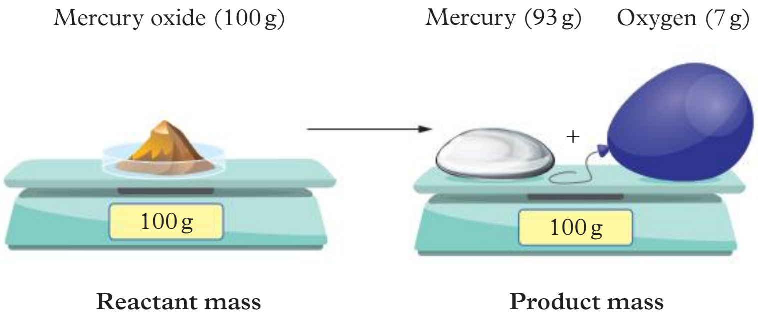


FIGURE 6 The atoms of mercury oxide (HgO) rearrange to form liquid mercury (Hg) and oxygen gas (O_2).

Law of conservation of mass

The law of conservation of mass states that matter cannot be created or destroyed. According to this law, all chemical equations must be balanced. A balanced chemical equation has the same number of each type of atom on either side of the reaction arrow.

Consider the example in Figure 6. When heated, solid mercury oxide (HgO) decomposes into liquid mercury (Hg) and oxygen gas (O_2). We can write this as a chemical equation using the following steps.

1 Write the reactant/s on the left, product/s on the right, and a reaction arrow in between:

$$HgO \rightarrow Hg + O_2$$

2 Tally the atoms on either side of the reaction arrow to determine whether it is balanced.

Left side		Right side	
Hg	O	Hg	O
1	1	1	2

There are two oxygen atoms on the product side of the reaction and only one on the reactant side. The equation is not balanced.

3 Add coefficients in front of the molecules to add more molecules. These numbers multiply all the atoms in the chemical formula. For example, putting a 2 in front of the HgO doubles the number of Hg and O atoms. Balance one atom at a time. Let's look at O first: there are 2 on the right but only 1 on the left, so we can add a coefficient of 2 to HgO.

$$2HgO \rightarrow Hg + O_2$$

Left side		Right side	
Hg	O	Hg	O
2	1	1	2

Now, the O atoms are balanced, but Hg is not. There are 2 on the left and 1 on the right, so we can add a coefficient of 2 to Hg.

$$2HgO \rightarrow 2Hg + O_2$$

Left side		Right side	
Hg	O	Hg	O
2	2	2	2

Now, both Hg and O atoms are balanced.

Study tip

Questions involving balancing chemical equation are typically worth one mark. If the chemical equation is balanced incorrectly or the formulas and states are incorrect, no mark is allocated.

Study tip

If a polyatomic ion (an ion that is made up of more than one atom, such as hydroxide (OH^-) or nitrate (NO_3^-)) remains unchanged in a chemical reaction, it can be treated as one entity, rather than breaking it up into individual atoms.

Study tip

Only the coefficients in front of the chemical formulas can be changed to balance an equation. The chemical formulas themselves, i.e. the subscripts, cannot be altered.

Using solubility rules to assign states

States are used to indicate the nature of reactants or products. They must be assigned to all substances in chemical equations. The four states that are assigned in chemical reactions are solid, liquid, gas and aqueous.

- Solid (s) – all metals, other than mercury, are solids if they are unreactive at room temperature. Some non-metals, such as carbon and sulfur, also form solids. In precipitation reactions, a solid is formed. To determine which product has formed the solid, a solubility table can be used. This is a list of ions that indicates whether they form precipitates or remain in solution when combined with other ions. In the table, if a substance is listed as insoluble, it does not dissolve in water and is therefore a solid.

Some simple rules for solubility of salts in water are:

1 All nitrate (NO_3^-) salts are soluble.
2 Most salts of Na^+, K^+ and NH_4^+ are soluble.
3 Most chloride salts are soluble. Notable exceptions are $AgCl$ and Hg_2Cl_2.
4 Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$ and $CaSO_4$.
5 Most hydroxides are only slightly soluble. The important soluble hydroxides are $NaOH$, KOH and $Ca(OH)_2$.
6 Most sulfide (S^{2-}), carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}) salts are only slightly soluble.

Liquid (l) – water is almost always a liquid, unless other stated. For example, water vapour is gaseous water and ice is solid water. Ionic substances can also be liquids, but only when they are melted, or **molten**. For example, sodium chloride melts at approximately 800°C, at which point it becomes a liquid.

molten
a melted substance

- Gas (g) – common gases at room temperature include H_2, F_2, Cl_2, N_2, O_2 and the noble gases.
- Aqueous (aq) – Solutions formed from water-soluble substances and water are **aqueous** solutions.

aqueous
a substance that is dissolved in water or the state given to a substance that is dissolved in water

What are single displacement reactions?

A **single displacement reaction** occurs when a more reactive element displaces a less reactive element from a compound. The abbreviated reactivity series for metals and non-metals are outlined in Tables 1 and 2.

single displacement reaction
a chemical reaction in which a more reactive element takes the place of a less reactive element in a compound

TABLE 1 Abbreviated reactivity series for metals

K	Most reactive
Na	↑
Li	
Ba	
Sr	
Ca	
Mg	
Al	
I	
Mn	
Zn	
Cr	
Fe	
Cd	
Co	
Ni	
Sn	
Pb	
(H)	
Cu	
Ag	
Au	Least reactive

TABLE 2 Abbreviated reactivity series for non-metals

F	Most reactive
Cl	↑
Br	
O	
I	
S	Least reactive

Study tip
To determine if an ionic substance is solid or aqueous, use the "Solubility of selected compounds" table in the Formula and data book.

Consider the reaction of aqueous silver(I) nitrate and copper metal. Copper is more reactive than silver (Table 1), so it displaces Ag^+ ions from solution to form aqueous copper(II) nitrate and silver metal, as the following unbalanced equation shows:

$$Ag(NO_3)(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

We will balance this equation in Worked example 8.1A.

We can also consider the reaction of sodium bromide and a non-metal compound, chlorine gas. Chlorine is more reactive than bromine, so it displaces bromine from the sodium ion to form sodium chloride and bromine gas, as the following unbalanced equation shows.

$$NaBr(s) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$$

We will balance this equation in Worked example 8.1B.

Study tip

In Worked example 8.1A, the equation has been rewritten at each step to walk you through the thinking process. When you complete these questions yourself, you do not need to do this – just build on the same equation and leave spaces in front of each compound so you can fill in the coefficients.

Worked example 8.1A

Writing balanced equations for single displacement reactions involving metals

Worked example 8.1A: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the single displacement reaction between silver nitrate solution and copper metal. (1 mark)

$$Ag(NO_3)(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to express the reaction as a balanced chemical equation. The question is worth 1 mark, so we must recall the theory and apply it to the reaction.
Step 2: Recall the formulas for the species involved in the reaction, including their states. This is a single displacement reaction, so we need to switch the ions. In some cases, you will need to deduce whether a reaction will occur using the reactivity series. The reaction will proceed if the metal being added is more reactive than the metal in the compound. In this case, Cu is more reactive than Ag, so the reaction will occur. We have also already been given the unbalanced equation for this particular reaction, so we can just copy it.	$AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$
Step 3: Tally the atoms on either side of the reaction arrow to determine whether it is balanced. We can group the polyatomic ions to make things easier.	(see tally table below) The equation is not balanced.

Left side			Right side		
Ag	NO_3	Cu	Ag	NO_3	Cu
1	1	1	1	2	1

The equation is not balanced.

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 4: Add coefficients in front of the reactants and products to add more as required. Balance one atom (or groups of atoms in a polyatomic ion) at a time. Let's look at NO_3 first.</td><td>There are 2 NO_3 on the right but only 1 on the left, so we need to add a 2 in front of $AgNO_3$ on the left.
<table><tr><th colspan="3">Left side</th><th colspan="3">Right side</th></tr><tr><td>Ag</td><td>NO_3</td><td>Cu</td><td>Ag</td><td>NO_3</td><td>Cu</td></tr><tr><td>2</td><td>2</td><td>1</td><td>1</td><td>2</td><td>1</td></tr></table>The equation is still not balanced.</td></tr>
<tr><td>Step 5: Now, the NO_3 are balanced, but the Ag are not.</td><td>There are 2 Ag on the left but only 1 on the right, so we need to add a 2 in front of Ag on the right.
$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$
<table><tr><th colspan="3">Left side</th><th colspan="3">Right side</th></tr><tr><td>Ag</td><td>NO_3</td><td>Cu</td><td>Ag</td><td>NO_3</td><td>Cu</td></tr><tr><td>2</td><td>2</td><td>1</td><td>2</td><td>2</td><td>1</td></tr></table>The equation is now balanced.</td></tr>
<tr><td>Step 6: Finalise your answer by checking that you have the correct formulas, coefficients and states.</td><td>$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$ (1 mark)</td></tr>
</table>

Your turn

Determine the balanced chemical equation for the reaction between aluminium and a solution of copper sulfate. (1 mark)

$$Al(s) + CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + Cu(s)$$

Worked example 8.1B

Writing balanced equations for single displacement reactions involving non-metals

Worked example 8.1B: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the single displacement reaction between sodium bromide and chlorine gas. (1 mark)

$$NaBr(aq) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$$

<table>
<tr><th>Think</th><th>Do</th></tr>
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<tr><td>Step 2: Tally the atoms on either side of the reaction arrow to determine whether it is balanced. We can group the polyatomic ions to make things easier.</td><td>$Br(aq) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$
<table><tr><th colspan="3">Left side</th><th colspan="3">Right side</th></tr><tr><td>Na</td><td>Br</td><td>Cl</td><td>Na</td><td>Br</td><td>Cl</td></tr><tr><td>1</td><td>1</td><td>2</td><td>1</td><td>2</td><td>1</td></tr></table>The equation is not balanced.</td></tr>
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<tr><td>Step 3: Add coefficients in front of the reactants and products to add more. Balance one atom (or groups of atoms in a polyatomic ion) at a time. Let’s look at Br first.</td><td>There are 2 Br on the right but only 1 on the left, so we need to add a 2 in front of NaBr on the left.
<table>
<tr><th colspan="3">Left side</th><th colspan="3">Right side</th></tr>
<tr><td>Na</td><td>Br</td><td>Cl</td><td>Na</td><td>Br</td><td>Cl</td></tr>
<tr><td>2</td><td>2</td><td>2</td><td>1</td><td>2</td><td>1</td></tr>
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$2NaBr(aq) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$
The equation is still not balanced.</td></tr>
<tr><td>Step 4: Now, the Br are balanced, but the Cl are not.</td><td>There are 2 Cl on the left but only 1 on the right, so we need to add a 2 in front of NaCl on the right.
$2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(g)$
<table>
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<tr><td>2</td><td>2</td><td>2</td><td>2</td><td>2</td><td>2</td></tr>
</table>
The equation is now balanced. This also balanced the Na on both sides.</td></tr>
<tr><td>Step 5: Finalise your answer by checking that you have the correct formulas, coefficients and states.</td><td>$2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(aq) + Br_2(g)$ (1 mark)</td></tr>
</table>

Your turn

Determine the balanced chemical equation for the single displacement reaction between solutions of potassium iodide and bromine. (1 mark)

$$KI(aq) + Br_2(aq) \rightarrow KBr(aq) + I_2(aq)$$

Study tip

For double displacement reactions, take extra care to check the solubility of the ions formed, as these reactions are often precipitation reactions. You will learn more about this in Unit 2.

double displacement reaction

a chemical reaction in which both the cation and anion of one reactant swaps with the cation and anion of the second reactant, forming new products

What are double displacement reactions?

A **double displacement reaction** occurs when the anion (negatively charged ion) and cation (positively charged ion) swap partners. For example, lead and sodium swap partners to form new products.

lead(II) nitrate + sodium hydroxide → lead(II) hydroxide + sodium nitrate

According to the solubility table, all nitrates are soluble in water, and are therefore aqueous. Sodium hydroxide is also soluble in water and therefore aqueous, but lead(II) hydroxide is insoluble and therefore a solid. This is an example of a precipitation reaction because a solid forms, as the following unbalanced equation shows.

$$Pb(NO_3)_2(aq) + NaOH(aq) \rightarrow Pb(OH)_2(s) + NaNO_3(aq)$$

We will balance this in Worked example 8.1C.

Worked example 8.1C

Writing balanced equations for double displacement reactions

Worked example 8.1C: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the reaction between lead (II) nitrate and sodium hydroxide. (1 mark)

$$Pb(NO_3)_2(aq) + NaOH(aq) \rightarrow Pb(OH)_2(s) + NaNO_3(aq)$$

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to express the reaction as a balanced chemical equation. The question is worth 1 mark, so we must recall the theory and apply it to the reaction.
Step 2: Tally the atoms on either side of the reaction arrow to determine whether it is balanced. We can group the polyatomic ions to make things easier.	$NaBr(aq) + Cl_2(g) \rightarrow NaCl(aq) + Br_2(g)$
Step 3: Add coefficients in front of the reactants and products to add more. Balance one atom (or groups of atoms in a polyatomic ion) at a time. Let's look at OH first.	There are 2 OH on the right but only 1 on the left, so we need to add a 2 in front of NaOH on the left. $Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + NaNO_3(aq)$
Step 4: Now, the OH are balanced, but the Na are not.	There are 2 Na on the left but only 1 on the right, so we need to add a 2 in front of the $NaNO_3$ on the right. $Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$
Step 5: Finalise your answer by checking that you have the correct formulas, coefficients and states.	$Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$ (1 mark)

Step 2 tally:

Left side				Right side			
Pb	NO_3	Na	OH	Pb	NO_3	Na	OH
1	2	1	1	1	1	1	2

The equation is not balanced.

Step 3 tally:

Left side				Right side			
Pb	NO_3	Na	OH	Pb	NO_3	Na	OH
1	2	2	2	1	1	1	2

The equation is still not balanced.

Step 4 tally:

Left side				Right side			
Pb	NO_3	Na	OH	Pb	NO_3	Na	OH
1	2	2	2	1	2	2	2

The equation is now balanced. This also balanced the NO_3 on both sides.

Your turn

Determine the balanced chemical equation for the double displacement reaction between sodium sulfate and barium chloride solutions. (1 mark)

$$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + NaCl(aq)$$

What are acid–base reactions?

The reaction of an acid with a base forms water and a salt. You will learn more about acids, bases and the reactions between them in Unit 2. For now, we will just focus on writing balanced equations for the reactions.

The formulas of some common acids and bases can be found in Table 3.

Study tip

Once you learn about acids and bases, it will be assumed that you know the common names and formula for some of them as listed in Table 3. Spend time learning the formulas for hydrochloric acid, nitric acid, sulfuric acid, ethanoic acid, and ammonia.

TABLE 3 Some common acids and bases

Formula	Name of acid	Formula	Name of base
CH_3COOH	Acetic acid (ethanoic acid)	KOH	Potassium hydroxide
HCl	Hydrochloric acid	NH_3	Ammonia
HNO_3	Nitric acid	MgO	Magnesium oxide
H_2CO_3	Carbonic acid	$NaOH$	Sodium hydroxide
H_2SO_4	Sulfuric acid	Na_2O	Sodium oxide
H_3PO_4	Phosphoric acid	$Ca(OH)_2$	Calcium hydroxide

Sulfuric acid reacts with potassium hydroxide (a base) to form a salt (potassium sulfate) and water, as the following unbalanced equation shows.

$$H_2SO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$$

Acid–base reactions are also a type of double displacement reaction. Can you see the similarities?

Worked example 8.1D

Writing balanced equations for acid–base reactions

Worked example 8.1D: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the reaction between sulfuric acid and potassium hydroxide. (1 mark)

$$H_2SO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$$

<table>
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<tr><td>Step 2: Tally the atoms on either side of the reaction arrow to determine whether it is balanced. We can group the polyatomic ions to make things easier. In acid–base reactions, the H and OH combine to form H_2O, but we will separate them into H and O.</td><td>$H_2SO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$
<table>
<tr><th colspan="4">Left side</th><th colspan="4">Right side</th></tr>
<tr><td>H</td><td>SO_4</td><td>K</td><td>O</td><td>H</td><td>SO_4</td><td>K</td><td>O</td></tr>
<tr><td>3</td><td>1</td><td>1</td><td>1</td><td>2</td><td>1</td><td>2</td><td>1</td></tr>
</table>
The equation is not balanced.</td></tr>
<tr><td>Step 3: Add coefficients in front of the reactants and products to add more. Balance one atom (or groups of atoms in a polyatomic ion) at a time. Let’s look at K first.</td><td>There are 2 K on the right but only 1 on the left, so we need to add a 2 in front of KOH on the left.
$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$
<table>
<tr><th colspan="4">Left side</th><th colspan="4">Right side</th></tr>
<tr><td>H</td><td>SO_4</td><td>K</td><td>O</td><td>H</td><td>SO_4</td><td>K</td><td>O</td></tr>
<tr><td>4</td><td>1</td><td>2</td><td>2</td><td>2</td><td>1</td><td>2</td><td>1</td></tr>
</table>
The equation is still not balanced.</td></tr>
</table>

Think	Do
Step 4: Now, the K are balanced, but the O are not.	There are 2 O on the left but only 1 on the right, so we need to add 2 in front of H_2O on the right. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$ Left side: H 4, SO_4 1, K 2, O 2; Right side: H 4, SO_4 1, K 2, O 2 The equation is now balanced. This also balanced the H on both sides.
Step 5: Finalise your answer by checking that you have the correct formulas, coefficients and states.	$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$ (1 mark)

Left side				Right side			
H	SO_4	K	O	H	SO_4	K	O
4	1	2	2	4	1	2	2

Your turn

Determine the balanced chemical equation for the reaction between hydrochloric acid and calcium hydroxide. (1 mark)

$$HCl\ (aq) + Ca(OH)_2\ (aq) \rightarrow CaCl_2(aq) + H_2O(l)$$

What are combustion reactions?

A **combustion reaction** is a reaction between oxygen and an element or a molecule. There are two primary types of combustion reactions: those involving the combustion of metals and non-metals and those involving the combustion of hydrocarbons.

combustion reaction
a chemical reaction in which a reactant is burnt with oxygen to form a metal oxide, covalent compound or carbon dioxide and water

Combustion of metals and non-metals

A metal combusts (reacts with oxygen) to form a metal oxide:

metal + oxygen → metal oxide

For example: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

A non-metal combusts to form a covalent compound:

non-metal + oxygen → covalent compound

For example: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Complete combustion of hydrocarbons

When a hydrocarbon is combusted with sufficient oxygen, it usually forms carbon dioxide and water. This is called **complete combustion**. For example, the combustion of the fuel methane produces carbon dioxide and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

This chemical equation is a simple one to balance as long as carbon and hydrogen are balanced first. Other combustion reactions of hydrocarbons can be balanced by applying the following rules.

1 Balance oxygen last. Oxygen is in three of the four molecules, so balancing carbon and hydrogen first makes it easier to balance oxygen as the last step.

2 After carbon and hydrogen are balanced, there is sometimes an odd number of oxygen atoms on the product side. If this occurs, double everything.

complete combustion
when a fuel reacts with sufficient oxygen (an excess) to produce carbon dioxide, water and heat energy

Study tip

For combustion reactions, always balance the carbons first, then the hydrogens, and finally, the oxygens.

Study tip

Fractions can only be used as coefficients in thermochemical equations.

For example, this equation shows $\frac{13}{2}$ moles of oxygen.

$$C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$$

You will need to double the coefficients to achieve the final equation of:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

Incomplete combustion of hydrocarbons

There are two types of combustion reactions where carbon dioxide does not form. These occur when the amount of oxygen is low or limiting, and are called **incomplete combustion** reactions. A reduced amount of oxygen means there is less oxygen bonding to the carbon atom and the products are either carbon monoxide (CO) and water (H_2O) or carbon soot (C) and water (H_2O). Often, some carbon dioxide (CO_2) is also formed.

incomplete combustion
when a hydrocarbon reacts with a limited supply of oxygen

Worked example 8.1E

Writing balanced equations for complete combustion of hydrocarbons

Worked example 8.1E: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the combustion of ethane gas. (1 mark)

$$C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to express the reaction as a balanced chemical equation. The question is worth 1 mark, so we must recall the theory and apply it to the reaction.
Step 2: Tally the atoms on either side of the reaction arrow to determine whether it is balanced.	$C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ Left side: C 2, H 6, O 2; Right side: C 1, H 2, O 3 The equation is not balanced.
Step 3: Add coefficients in front of the molecules to add more molecules. Balance one atom at a time. For combustion reactions, balance the C first.	There are 2 C on the left but only 1 on the right, so we need to add a 2 in front of CO_2 on the right. $C_2H_6(g) + O_2(g) \rightarrow 2CO_2(g) + H_2O(g)$ Left side: C 2, H 6, O 2; Right side: C 2, H 2, O 5 The equation is still not balanced.
Step 4: Now the C are balanced, so we should move on to the H.	There are 6 H on the left but only 2 on the right, so we need to add a 3 in front of H_2O on the right. Left side: C 2, H 6, O 2; Right side: C 2, H 6, O 7 $C_2H_6(g) + O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$ The equation is still not balanced.

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 5: Now the H are balanced, so we should move on to the O.</td><td>There are 7 O on the right but only 2 on the left, so we can use a fraction of $\frac{7}{2}$ in front of O_2 on the left.
$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
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<tr><td>Step 6: Finalise your answer by checking that you have the correct formulas, coefficients and states. You may need to undo the fraction by multiplying everything by 2.</td><td>$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
or $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$ (1 mark)</td></tr>
</table>

Your turn

Determine the balanced chemical equation for the combustion of octane. (1 mark)

$$C_8H_{18}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Worked example 8.1F

Writing balanced equations for combustion in limited oxygen

Worked example 8.1F: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the combustion of methane gas in limited oxygen. (1 mark)

$$CH_4(g) + O_2(g) \rightarrow CO(g) + H_2O(g)$$

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<tr><td>Step 3: Add coefficients in front of the molecules to add more molecules. Balance one atom (or groups of atoms in a polyatomic ion) at a time. The C are balanced first in a combustion reaction, but since they are already balanced in this one, we move on to H.</td><td>There are 4 H on the left but only 2 on the right, so we need to add a 2 in front of H_2O on the right.
$CH_4(g) + O_2(g) \rightarrow CO(g) + 2H_2O(g)$
<table><tr><th colspan="3">Left side</th><th colspan="3">Right side</th></tr><tr><td>C</td><td>H</td><td>O</td><td>C</td><td>H</td><td>O</td></tr><tr><td>1</td><td>4</td><td>2</td><td>1</td><td>4</td><td>3</td></tr></table>The equation is still not balanced.</td></tr>
</table>

Think	Do
Step 5: Now the H are balanced, so we should move on to the O.	There are 3 O on the right but only 2 on the left, so we can use a fraction of $\frac{3}{2}$ in front of O_2 on the left. $CH_4(g) + \frac{3}{2}O_2(g) \rightarrow CO(g) + 2H_2O(g)$ Left side: C 1, H 4, O 3; Right side: C 1, H 4, O 3 The equation is now balanced.
Step 6: Finalise your answer by checking that you have the correct formulas, coefficients and states. You will need to undo the fraction by multiplying everything by 2.	$CH_4(g) + \frac{3}{2}O_2(g) \rightarrow CO(g) + 2H_2O(g)$ or $2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(g)$ (1 mark)

Left side			Right side		
C	H	O	C	H	O
1	4	3	1	4	3

Your turn

Determine the balanced chemical equation for the combustion of propane gas in limited oxygen. (1 mark)

$$C_3H_8\,(g) + O_2(g) \rightarrow CO(g) + H_2O(g)$$

What are combination and decomposition reactions?

combination reaction
a chemical reaction in which two or more elements and/or compounds react to form a single compound; also called a synthesis or addition reaction

decomposition
a chemical reaction in which a compound breaks down into simpler compounds or elements on the application of heat or electricity

A **combination reaction**, sometimes called synthesis or addition, involves reacting two substances to form one product; for example, the reaction between sodium and chlorine to form sodium chloride.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl_2(s)$$

The combustion of metals and non-metals are also examples of combination reactions.

The opposite of this is a **decomposition** reaction, which occurs when a compound breaks down into smaller compounds, elements, or a mixture of both. This happens when heat or electricity is applied. For example, water decomposes when electricity is applied using a power pack, wires and platinum electrodes, according to the chemical equation:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

Calcium carbonate can also be decomposed by thermal decomposition or heat according to the chemical equation:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Dehydration is another example of decomposition. A hydrated salt is a crystalline salt that contains a certain number of water molecules loosely held in the crystal lattice. The number of water molecules is signified by putting "$\cdot xH_2O$" after the salt. For example, copper(II) sulfate pentahydrate, $CuSO_4{\cdot}5H_2O(s)$ contains five water molecules for every $CuSO_4$.

$$CuSO_4{\cdot}5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$$

What are simple redox reactions?

redox reaction
a chemical reaction involving the transfer of electrons from one reactant to another

Simple **redox reactions** involve the transfer of electrons from one reactant to another. Several of the examples we have already discussed, such as single displacement reactions, combustion reactions, and some but not all decomposition reactions, are also redox reactions. Single displacement reactions are always redox reactions.

The reaction of zinc metal with hydrochloric acid is shown below. In this single displacement reaction, zinc is more reactive than hydrogen and will therefore form zinc chloride and hydrogen gas. This is a redox reaction because the **oxidation states** (valency or assigned charge) of certain atoms have changed (Table 4).

$$\mathrm{Zn(s) + HCl \rightarrow ZnCl_2(aq) + H_2(g)}$$

TABLE 4 Oxidation states of reactants and products

Zn	H^+	Cl^-	→	Zn^{2+}	Cl^-	H^2
0	+1	–1	→	+2	–1	0

The oxidation state of zinc increases from 0 to +2, the oxidation state of hydrogen decreases from +1 to 0 and the oxidation state of the chloride ion is unchanged. The oxidation state of elements is always 0.

oxidation state
the charge assigned to substances in a chemical reaction

Study tip

You will learn more about redox reactions and oxidation states in Unit 3.

Study tip

When balancing redox reactions, start by balancing the metals, then non-metals, then oxygen and lastly hydrogen.

Worked example 8.1G

Writing balanced equations for redox reactions

Worked example 8.1G: Watch a video that shows how to solve this problem.

Determine the balanced chemical equation for the reaction between zinc metal and hydrochloric acid. (1 mark)

$$\mathrm{Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)}$$

<table>
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<tr><td>Zn</td><td>H</td><td>CI</td><td>Zn</td><td>H</td><td>Cl</td></tr>
<tr><td>1</td><td>1</td><td>1</td><td>1</td><td>2</td><td>2</td></tr>
</table>
The equation is not balanced.</td></tr>
<tr><td>Step 3: Add coefficients in front of the reactants and products to add more. Balance one atom (or groups of atoms in a polyatomic ion) at a time. The metals are balanced first in a redox reaction, but since they are already balanced in this one, we move on to Cl.</td><td>There are 2 Cl on the right but only 1 on the left, so we need to add a 2 in front of HCl on the right.
$\mathrm{Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)}$
<table>
<tr><th colspan="3">Left side</th><th colspan="3">Right side</th></tr>
<tr><td>Zn</td><td>H</td><td>CI</td><td>Zn</td><td>H</td><td>Cl</td></tr>
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The equation is now balanced. This also balanced the H on both sides.</td></tr>
<tr><td>Step 4: Finalise your answer by checking that you have the correct formulas, coefficients and states.</td><td>$\mathrm{Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)}$ (1 mark)</td></tr>
</table>

Your turn

Determine the balanced chemical equation for the reaction between aluminium metal and hydrochloric acid. (1 mark)

$$\mathrm{Al(s) + HCl \rightarrow AlCl_3(aq) + H_2(g)}$$

Real-world chemistry

Fuels: Where to in the future?

The internal combustion engine was invented in the late 1700s. Every aspect of society has benefited from its use, from transport to agriculture, industry and power generation. They have since been re-designed to be smaller, lighter, safer to use, and start very quickly. Combustion engines have become highly efficient, losing little energy in the form of heat and sound compared to earlier models.

Most internal combustion engines use petrol or diesel as fuels. They are hydrocarbons derived from fossil fuels and have profound negative environmental impacts. As carbon dioxide (CO_2) is produced during combustion of these fuels, this has resulted in enormous CO_2 emissions, which have flow-on effects on the global temperatures and sea level. For example, the concentration of CO_2 has experienced an increase of more than 200 ppm over the last 20,000 years. This has been associated with global temperatures increasing by more than 4°C and sea level increasing more than 120 m.

One way to curb further increases to CO_2 emissions is by switching from fossil fuels to biofuels. Biofuels are fuels sourced from biological sources, like plants, rather than fossil fuels. Although they produce CO_2 during combustion, the growth of the plants in which they are sourced takes in a significant amount of carbon dioxide. Biofuels therefore contribute far less to any extra atmospheric CO_2.

A common biofuel is ethanol made from fermentation of plant material. It is added to some petrol and sold as E-10. Many modern cars can use E-10 in their engines. Biodiesel can also be manufactured from oils from crops such as soybean, sunflower or canola, or even used cooking oil. Both fuels have the potential to help the transport industry reduce CO_2 emissions.

FIGURE 7 Fossil fuels have contributed enormously to CO_2 emissions.

Apply your understanding

1 Diesel made from fossil fuels is a mixture of hydrocarbons with 10 to 15 C atoms per molecule. An average formula is $C_{12}H_{23}$. **Determine** the balanced chemical equation to represent the combustion of diesel. (1 mark)

2 Biodiesel can have various formulas depending on the source of oil used to make it. One example is $C_{17}H_{34}O_2$. **Determine** the balanced chemical equation for the combustion of this substance. (1 mark)

3 **Determine** the mass of oxygen required to burn 100 g of diesel (average formula) and 100 g of biodiesel (average formula) from the equations in questions **1** and **2**. (8 marks)

Check your learning 8.1

Check your learning 8.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** why it is necessary to balance chemical equations. (2 marks)

2 **Describe** how the state (aq) differs from the state (l). (2 marks)

3 **Explain** how energy changes during phase change, using an example. (3 marks)

Analytical processes

4 **Determine** balanced equations for each of the following reactions.

- **a** $C_3H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$ (1 mark)
- **b** $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$ (1 mark)
- **c** $N_2(g) + O_2(g) \rightarrow NO(g)$ (1 mark)
- **d** $C_4H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$ (1 mark)
- **e** $C_6H_{12}O_6(aq) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ (1 mark)
- **f** $KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + H_2O(l)$ (1 mark)
- **g** $Zn(s) + Pb(NO_3)_4(aq) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$ (1 mark)
- **h** $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow NaCl(aq) + BaSO_4(s)$ (1 mark)
- **i** $NaOH(aq) + Cu(NO_3)_2(aq) \rightarrow NaNO_3(aq) + Cu(OH)_2(s)$ (1 mark)
- **j** $H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g)$. (1 mark)

5 **Determine** a balanced chemical equation for the decomposition of nickel(III) oxide. (1 mark)

$$Ni_2O_3(s) \rightarrow Ni(s) + O_2(g)$$

6 **Determine** a balanced chemical equation for the combustion of methane gas in very limited oxygen. (1 mark)

$$CH_4(g) + O_2(g) \rightarrow C(s) + H_2O(g)$$

7 **Contrast** phase changes and chemical reactions. (3 marks)

8 **Determine** the correct formula and the state symbol of the following substances at room temperature.

- **a** Silver bromide (1 mark)
- **b** Potassium nitrate (1 mark)
- **c** Barium sulfate (1 mark)
- **d** Oxygen (1 mark)
- **e** Water (1 mark)
- **f** Sulfur (1 mark)
- **g** Carbon monoxide (1 mark)
- **h** Calcium carbonate (1 mark)
- **i** Magnesium metal (1 mark)
- **j** Chlorine (1 mark)
- **k** Bromine (1 mark)
- **l** Nitrogen (1 mark)

Lesson 8.2

Chemical energy and thermochemistry

Key ideas

- → Endothermic reactions involve the absorption of heat from the environment and a decrease in temperature. Exothermic reactions involve the release of heat into the environment and an increase in temperature.
- → Enthalpy is the total heat content of a substance.
- → Enthalpy level diagrams are used to show the change in heat energy before, during and after a reaction.
- → Thermochemical equations are balanced chemical equations that also show enthalpy.

Learning intentions and success criteria

What happens to chemical energy in a reaction?

The **law of conservation of energy** states that the total amount of energy cannot be created or destroyed: energy is instead transferred from one place to another or transformed from one type of energy into another. For example, when a television is switched on, electrical energy is transformed into sound energy, light energy and a small amount of heat (thermal) energy.

Chemical energy is the potential energy stored within the bonds of substances. Different molecules are made up of different atoms bonded in different ways, so the reactants and products in chemical equations have different amounts of potential energy. When a chemical reaction takes place and atoms rearrange to form products, the energy level within the reaction system must change. One of two situations will occur:

1 If the reactants have more energy than the products, excess energy is released from the system (the reaction) into the environment. This is called an **exothermic** reaction.

2 If the products have more energy than the reactants, more energy is needed to form them. The extra energy is absorbed from the environment into the system. This is called an **endothermic** reaction.

law of conservation of energy
states energy cannot be created or destroyed

exothermic
a reaction in which the reactants have more energy than the products; excess energy is released to the environment as heat

endothermic
a reaction in which the reactants have less energy than products; the required energy is absorbed from the environment as heat

What is enthalpy?

The amount of energy stored within a substance is called its **enthalpy**, or heat content, and is represented by the symbol H. Enthalpy is composed of both kinetic and potential energy. It is not possible to measure or calculate enthalpy; however, it is possible to measure changes in enthalpy. Therefore, for a reaction, we typically refer to the change in enthalpy (ΔH):

$$\Delta H = H_{\text{(products)}} - H_{\text{(reactants)}}$$

One change that can be measured is the amount of energy required to break a bond, which is the same as the energy released when that type of bond is formed. Different bond types require a different amount of energy to break them. This quantity is known as the **bond enthalpy**. Specifically, it is the amount of energy required to break 1 mol of that type of bond in the gaseous phase. The bond enthalpies in your formula and data book are averaged values for bonds of that type.

Bond enthalpy determines **bond stability**. If the bond enthalpy is high, more energy is required to break the bond and therefore, the bond is more stable. In general, the more electrons shared between bonded atoms, the more energy required to break the bond.

Study tip

Exo means "out", so energy leaves (or is released from) the system in an exothermic reaction. *Endo* means "in", so energy enters (or is absorbed into) the system in an endothermic reaction.

enthalpy
the total energy or heat content of chemical substances or a system

bond enthalpy
the amount of energy required to break 1 mol of a particular bond in the gaseous phase to give uncharged fragments

bond stability
a measure of how difficult it is to break a bond; depends on the enthalpy of the bond

The bond enthalpy of a substance

Bond enthalpies are useful to determine the energy required to break a substance into its individual atoms, or the energy released when the substance forms. This requires you to know the structure of a molecule and add the bond enthalpies for each bond. Bond enthalpies are not the total enthalpy of a substance.

There are limitations to using average bond enthalpies to calculate enthalpy change. These limitations are that bond energies:

- are average values and do not take into consideration that the same bond can have different energies depending on their reaction environment
- do not take into consideration any attractive forces that may exist in the reaction system
- were measured using gaseous reactants and products and therefore do not take into account the actual state of the reactants and products.

Study tip

Remember the limitations of using bond energy in calculations by remembering the three "A"s: Average values, Attractive forces and Actual states.

Worked example 8.2A

Calculating the bond enthalpy of a substance

Worked example 8.2A: Watch a video that shows how to solve this problem.

Calculate the bond enthalpy of 1 mol of propane. This is the energy required to break this amount of substance into its constituent atoms. (1 mark)

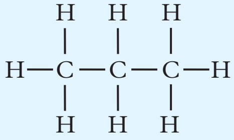


FIGURE 1 Structure of propane

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the bond enthalpy. The question is worth 1 mark, so we must gather the correct information and complete the calculation.
Step 2: Identify the bonds in the substance and list the enthalpy of each of the bonds.	Propane contains 2 C–C bonds and 8 C–H bonds. C–C: 346 kJ mol^{-1} C–H: 414 kJ mol^{-1}
Step 3: Multiply the bond enthalpy values by the number of bonds of each type.	$H(bonds) = (2 \times 346\,\text{kJ mol}^{-1}) + (8 \times 414\,\text{kJ mol}^{-1})$ $= 4{,}004\,\text{kJ mol}^{-1}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	4,000 kJ mol^{-1} (1 mark)

Your turn

Calculate the bond enthalpy of 1 mol of propanol. This is the energy required to break this amount of substance into its constituent atoms. (1 mark)

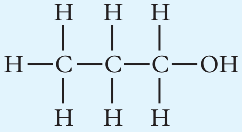


FIGURE 2 Structure of propanol

How are enthalpy, heat and temperature related?

It is important to understand the distinction between heat and temperature. Heat is a form of energy referred to as "thermal energy". Temperature is a measure of the average kinetic energy of the particles. A temperature increase of a system occurs due to thermal energy being added, so temperature and thermal energy are related; however, temperature alone does not measure the total thermal energy content. This is the kinetic energy component of enthalpy. In the next section, temperature changes due to reactions will be used to calculate the change in enthalpy.

Study tip

When calculating the bond enthalpy of a substance, remember to count each individual bond only once. If it helps, draw a structural formula, showing all bonds and cross them out as you go.

As an example, a full kettle of hot water contains much more thermal energy than a mug of water at the same temperature. A 100 g mass of water vapour or steam at 100°C contains more enthalpy than the same mass of liquid water at the same temperature; thermal energy has been added when water is converted to steam, without any corresponding increase in temperature. A 100 g mass of water and a 100 g mass of a different substance at the same temperature will contain different amounts of thermal energy and potential energy, and thus, have different enthalpies.

How is change in enthalpy calculated from bond enthalpies?

It is more useful to be able to calculate the change in enthalpy during a reaction. The difference in enthalpy values between reactants and products is represented as **ΔH**, where the Greek letter capital Δ is used to show a change. This is represented by:

$$\Delta H = H_{(\text{products})} - H_{(\text{reactants})}$$

However, this should not be used in calculations. Instead, the change in enthalpy can be calculated using the energy absorbed when reactants are broken apart minus the energy released when new bonds are formed to make products:

ΔH = energy required to break bonds – energy released when bonds are formed

$$\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma H_{\text{bonds formed}}$$

Consistent with the law of conservation of energy, any energy put into the system to break reactant bonds may be released when product bonds form and so this energy cancels out and only the difference in enthalpy between reactants and products determine the magnitude and sign of ΔH. Using bond enthalpies this way results in an approximate value for ΔH.

A positive ΔH value represents an endothermic reaction, whereas a negative ΔH value represents an exothermic reaction.

ΔH
the difference in enthalpy between reactants and products in a chemical equation

Study tip

When estimating enthalpy change of a reaction from bond enthalpies, use $\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma_{\text{Hbonds formed}}$.

Worked example 8.2B

Determining enthalpy change of a reaction from average bond enthalpies

Worked example 8.2B: Watch a video that shows how to solve this problem.

Ethyne (H–C≡C–H) reacts with hydrogen (H–H) to form ethane (CH_3–CH_3).

a **Use** bond enthalpies to **calculate** the ΔH of the reaction. (4 marks)

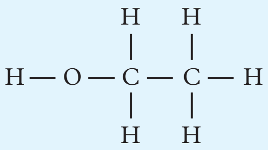
b **Deduce** whether the reaction is endothermic or exothermic. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Use" means to apply knowledge or rules to put theory into practice. "Calculate" means to determine or find a number or answer by using mathematical processes. "Deduce" means to draw a logical conclusion. We need to find the enthalpy using the bond enthalpy values, then decide whether the reaction is endothermic or exothermic. The questions are worth a variety of marks, so we must gather the correct information, correctly apply the formula, complete the calculation and state the answer.
Step 2: Write a balanced equation for the reaction.	$C_2H_2(g) + 2H_2(g) \rightarrow C_2H_6(g)$ (1 mark)

Think	Do
Step 3: Identify the bonds in each substance and list the enthalpy of each of the bonds. Separate reactants from products.	Reactants: Ethyne contains 2 C–H bonds and 1 C≡C bond. Hydrogen contains 1 H–H bond. C≡C: 839 kJ mol^{-1} C–H: 414 kJ mol^{-1} H–H: 436 kJ mol^{-1} Products: Ethane contains 1 C–C bond and 6 C–H bonds. C–C: 346 kJ mol^{-1} C–H: 414 kJ mol^{-1}
Step 4: Multiply the bond enthalpy values by the number of bonds of each type.	$H(\text{bonds})_{\text{ethyne}} = (2 \times 414\ \text{kJ mol}^{-1}) + 839\ \text{kJ mol}^{-1}$ $= 1{,}667\ \text{kJ mol}^{-1}$ $H(\text{bonds})_{\text{hydrogen}} = 2 \times 436\ \text{kJ mol}^{-1}$ $= 872\ \text{kJ mol}^{-1}$ (1 mark for $H_{\text{bonds broken}}$) $H(\text{bonds})_{\text{ethane}} = 346\ \text{kJ mol}^{-1} + (6 \times 414\ \text{kJ mol}^{-1})$ $= 2{,}830\ \text{kJ mol}^{-1}$ (1 mark for $H_{\text{bonds formed}}$)
Step 5: For part **a**, sum up the enthalpies of all reactants, then sum up the enthalpies of all products. Calculate ΔH.	$\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma H_{\text{bonds formed}}$ $= (1{,}667\ \text{kJ mol}^{-1} + 872\ \text{kJ mol}^{-1}) - 2{,}830\ \text{kJ mol}^{-1}$ $= -291\ \text{kJ mol}^{-1}$ (1 mark)
Step 6: For part **b**, inspect the sign (+ or –) of ΔH to draw a conclusion.	ΔH is negative, so $H_{\text{bonds formed}} > H_{\text{bonds broken}}$. Energy has been released to the surroundings and the reaction is exothermic. (1 mark)

Your turn

Ethanol has the structural formula shown.



$$\begin{array}{ccccccccc} & & & & H & & H & & \\ & & & & | & & | & & \\ H & - & O & - & C & - & C & - & H \\ & & & & | & & | & & \\ & & & & H & & H & & \end{array}$$

FIGURE 3 Structure of ethanol

Ethanol reacts with oxygen (O=O) to produce water vapour (H–O–H) and carbon dioxide (O=C=O).

a **Use** bond enthalpies to **calculate** the ΔH of the reaction. (4 marks)

b **Deduce** whether the reaction is endothermic or exothermic. (1 mark)

How are enthalpy level diagrams used in thermochemistry?

Thermochemistry is the study of heat that has been absorbed (endothermic) or released (exothermic) from a chemical reaction. This can be represented on an enthalpy level diagram.

An enthalpy level diagram shows the difference in enthalpies between reactants and products. Enthalpy, H, is represented on the y-axis, and the progress of the reaction, or time, is represented on the x-axis (Figure 4). As actual enthalpy values cannot be determined, the y-axis is not labelled with a scale or units (kJ mol^{-1}).

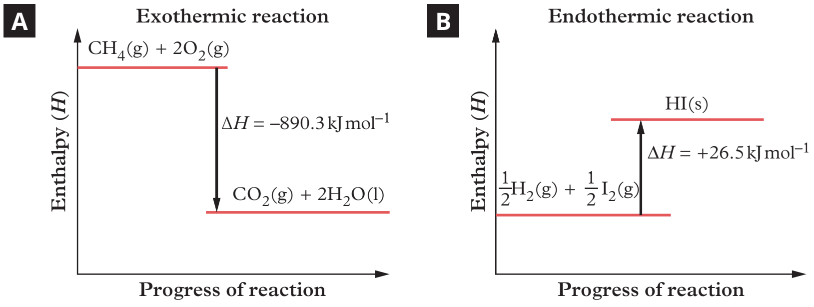


FIGURE 4 Enthalpy level diagrams of an (A) exothermic and (B) endothermic reaction.

The difference in enthalpy values between products and reactants represents the amount of energy that is absorbed or released by the system. This difference indicates whether the reaction is endothermic or exothermic.

Any energy absorbed or released during a reaction cannot disappear; it is conserved. In endothermic reactions, this energy is usually absorbed from the surroundings, making the surroundings cooler. In exothermic reactions, this energy is usually released to the surroundings, making them hotter. In some cases, the energy absorbed or released is not thermal energy but can be light or even sound energy.

Enthalpy level diagrams are also useful in determining the relative stability of reactants and products. All of this information is summarised in Table 1.

TABLE 1 Comparing exothermic and endothermic reactions

Type of reaction	Exothermic reaction	Endothermic reaction
Enthalpy of reactants ($H_{(reactants)}$) vs products ($H_{(products)}$)	$H_{(reactants)} > H_{(products)}$	$H_{(reactants)} < H_{(products)}$
Sign in front of ΔH	– (negative)	+ (positive)
Heat movement	When products form, residual energy transforms into heat and moves from the system to the environment, leaving the reaction vessel feeling hot.	For products to form, energy (generally in the form of heat) must be absorbed from the environment. This leaves the reaction vessel feeling cold.
Evidence of the law of conservation of energy	The energy released is the difference between the enthalpies of reactants and products.	The energy absorbed is the difference between the enthalpies of reactants and products.
Stability of reactants vs products	The products are more stable than the reactants. An input of energy would be required to reverse the reaction, i.e. turn the products back into reactants.	The reactants are more stable than the products. It will take an input of energy to change the reactants into products.

How do you write thermochemical equations?

thermochemical equation
a balanced chemical equation, including states, with a ΔH value that is positive or negative

Thermochemical equations are balanced chemical equations that always include a ΔH value. The ΔH value represents the energy difference in kilojoules (kJ) between the reactants and products for the number of mol of the reactants in the equation. Thus, the units are kJ.

The most common type of thermochemical equation you will encounter is the combustion of fuels. For these reactions, ΔH is the enthalpy or heat of combustion. To write these equations:

1 Start with a balanced chemical equation, including states. Table 2 lists the molar enthalpies ΔH and states of some common fuels. Because these are combustion reactions that generate a large amount of heat, water will typically vaporise and so should be represented as a gas.

2 Write the ΔH value next to the equation, in kJ. This must be negative for an exothermic reaction, which all combustion reactions are. If the coefficient in front of the fuel in the balanced equation is greater than 1, ΔH (in kJ) must be multiplied by the coefficient of the fuel.

For example, for the combustion of ethanol:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \ \Delta H = -1{,}360\,kJ$$

The units of kJ are used, rather than $kJ\,mol^{-1}$, as the equation already shows the number of moles of the substance being burnt; 1 mol in this case.

Study tip

Another reaction you will look at is when a solid substance dissolves. In this situation, ΔH is the heat of dissolution. Similarly, for acid–base neutralisation reactions, ΔH is the heat of neutralisation. You will learn more about these and how to calculate them in Lesson 8.4.

You may also see energy being included as a reactant (endothermic reactions) or product (exothermic reactions), within the equation. No positive or negative sign is required when thermochemical equations are written this way; energy is on the left if endothermic and on the right if exothermic.

$$\text{Endothermic: } N_2(g) + O_2(g) + 180.5\,kJ \rightarrow 2NO(g)$$
$$\text{Exothermic: } C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) + 1{,}360\,kJ$$

Now, consider the combustion of octane:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)\ \Delta H = -10{,}920\ kJ$$

From Table 2, the molar enthalpy of combustion is 5,460 kJ mol^{-1}. From the balanced chemical equation, there is a 2 in front of the C_8H_{18}, i.e. 2 moles of octane. Therefore, the molar enthalpy of combustion is multiplied by 2, giving 10,920 kJ.

TABLE 2 Enthalpy of combustion of some common fuels

Fuel	Formula	State	Molar enthalpy of combustion (kJ mol^{-1})	Enthalpy of combustion (kJ g^{-1})
Hydrogen	H_2	gas	–282	–141
Methane	CH_4	gas	–890	–55.6
Ethane	C_2H_6	gas	–1,560	–51.9
Propane	C_3H_8	gas	–2,220	–50.5
Butane	C_4H_{10}	gas	–2,880	–49.7
Octane	C_8H_{18}	liquid	–5,460	–47.9
Ethyne (acetylene)	C_2H_2	gas	–1,300	–49.9
Methanol	CH_3OH	liquid	–726	–22.7
Ethanol	C_2H_5OH	liquid	–1,360	–29.6
1-propanol	C_3H_7OH	liquid	–2,018	–33.6
1-butanol	C_4H_9OH	liquid	–2,676	–36.1

Study tip

Thermochemical equations are always written with a balanced chemical equation and a ΔH value in kJ. ΔH is the total enthalpy or heat released from the reaction, so when you refer to a table to obtain the molar ΔH value (in kJ mol^{-1}), you must multiply it by the coefficient in front of the fuel.

Study tip

Bond enthalpies and enthalpies (H) of a substance are presented as kJ mol^{-1} but ΔH in thermochemical equations for combustion are presented as kJ, since they take into account the number of moles of the fuel.

Worked example 8.2C

Writing thermochemical equations for combustion reactions

Worked example 8.2C: Watch a video that shows how to solve this problem.

Determine the thermochemical equation for the combustion of 1-propanol. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Determine” means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to write a balanced chemical equation and provide the enthalpy of combustion. The question is worth 2 marks, so we must express the reaction correctly and complete a calculation.
Step 2: Write a balanced chemical equation, including states. Remember that combustion is a reaction with oxygen to form carbon dioxide and water vapour. Refer to Worked example 8.1E for a refresher.	$2C_3H_7OH + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(g)$
Step 3: Find the molar enthalpy of combustion using Table 2 and multiply it by the number of moles of the fuel as indicated by its coefficient. Remember to show the negative sign, since it is an exothermic combustion reaction.	There is 2 mol of 1-propanol, so $\Delta H = -2 \times 2{,}018\,kJ\,mol^{-1} = -4{,}036\,kJ$.

Think	Do
Step 4: Finalise your answer by combining the balanced equation and the ΔH value, in kJ.	$2C_3H_7OH + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(g)$ $\Delta H = -4{,}036\,kJ$ (1 mark for balanced equation; 1 mark for correct ΔH)

Your turn

Determine the thermochemical equation for the combustion of methanol. (2 marks)

What happens when you double or reverse a thermochemical equation?

When the coefficients of a chemical equation are doubled, the molar ΔH value is also doubled because it is a measurement of the energy per mole of reactant. When a chemical equation is reversed, the sign of the ΔH value is reversed because the reactants and products in the enthalpy equation switch sides.

Worked example 8.2D

Writing thermochemical equations for reversed reactions

Worked example 8.2D: Watch a video that shows how to solve this problem.

Determine the thermochemical equation for the reverse reaction of the combustion of butane. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to write a balanced chemical equation and provide the enthalpy of combustion. The question is worth 2 marks, so we must express the reaction correctly and complete a calculation.
Step 2: Write a thermochemical equation for the combustion reaction, including states. Remember that combustion is a reaction with oxygen to form carbon dioxide and water vapour. Refer to Worked example 8.2C for a refresher.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$ $\Delta H = -5{,}760\,kJ$
Step 3: Finalise your answer by reversing the reaction. Switch the reactants and products. For the ΔH value, change the sign from – to +, since the reverse equation will be endothermic.	$8CO_2(g) + 10H_2O(g) \rightarrow 2C_4H_{10}(g) + 13O_2(g)$ $\Delta H = +5{,}760\,kJ$ (1 mark for balanced equation; 1 mark for correct ΔH)

Your turn

Determine the thermochemical equation for the reverse reaction of the combustion of octane. (2 marks)

Challenge

Manipulating enthalpy values

The combustion of iron has a ΔH value of 3,926 kJ for every 1 mol of iron. **Determine** the combustion equation for the reverse reaction, i.e. the decomposition of iron(III) oxide. (2 marks)

How are thermochemical equations used in stoichiometry?

Thermochemical equations are useful to determine the amount of heat released by any quantity of reacting substance.

Worked example 8.2E

Calculating the energy released in thermochemical reactions

Worked example 8.2E: Watch a video that shows how to solve this problem.

For the combustion of octane:

a **Determine** the balanced thermochemical equation. (2 marks)

b **Calculate** the amount of energy that would be released by combustion of 703 g of octane, which is 1 L of this fuel. (3 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. "Calculate" means to determine or find a number or answer by using mathematical processes. We need to write a balanced thermochemical equation and find the energy released from the reaction. The question is worth a variety of marks, so we must express the reaction correctly and complete a calculation.
Step 2: For part **a**, write a thermochemical equation for the combustion reaction, including states. Remember that combustion is a reaction with oxygen to form carbon dioxide and water vapour. Refer to Worked example 8.2C for a refresher.	$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)\ \Delta H = -10{,}920\,kJ$ (1 mark for balanced equation; 1 mark for correct ΔH)
Step 3: For part **b**, select the correct formula and gather any data required.	$n = \frac{m}{M}$ $n = ?,\ m = 703\,g,\ M = 114.26\,g\,mol^{-1}$
Step 4: Substitute the known values into the formula and solve for the unknown value.	$n = \frac{703\,g}{114.26\,g\,mol^{-1}}$ (1 mark) $= 6.1526\,mol$ (1 mark)
Step 5: Use the ratio of moles of fuel to the coefficient in the equation and the ΔH to calculate the energy released in the reaction.	$\Delta H = -10{,}920$ kJ (for 2 mol) $\Delta H = \frac{6.1526\,mol}{2} \times 10{,}920$ $= 33{,}593.38\,kJ$
Step 6: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	703 g of octane will release 33,593 kJ of energy. (1 mark)

Your turn

For the combustion of ethanol:

a **Determine** a balanced thermochemical equation. (2 marks)

b **Calculate** the amount of energy that would be released by combustion of 790 g of ethanol, which is 1 L of this fuel. (3 marks)

How is light energy involved in reactions?

For most endothermic reactions, the energy required is provided by heat from the surroundings. In some cases, electromagnetic radiation – such as light – provides the necessary energy to break bonds. Typically, this light is in the ultraviolet (UV) range of the electromagnetic spectrum.

An example of this is the breakdown of the chlorofluorocarbon (CFC) gas, CF_2Cl_2, also known as freon-12, in the stratosphere. It does not break down in the lower atmosphere; however, in the upper atmosphere, the molecules are exposed to the higher energy types of UV light, and the C–Cl bonds can be broken, releasing highly reactive neutral Cl atoms.

$$CF_2Cl_2(g) + \text{UV radiation} \rightarrow CF_2Cl(g) + Cl(g)$$

The neutral Cl atoms then catalyse the destruction of ozone in the stratosphere. This is why worldwide agreement saw the use of CFCs phased out.

We can also consider this in terms of reactions that release energy. Energy released in chemical reactions is generally released as thermal energy (exothermic) but can also be in the form of electromagnetic radiation, such as the combustion of the wax in a candle which provides sufficient visible light to see things.

Real-world chemistry

TNT: It's not dynamite

Nitroglycerin was first developed in 1847 by Ascanio Sobrero. When he first synthesised the compound, it was so unstable that it exploded in his face, leaving him scarred for the rest of his life. Nitroglycerin is a highly unstable contact explosive. This means it does not require a flame or spark to detonate; it needs only a small physical shock. It combusts according to the following thermochemical equation:

$$4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g) \quad \Delta H = -6{,}092\,\text{kJ}$$

It was Alfred Nobel, after whom the Nobel Prize is named, who started commercially producing the substance as an explosive for blast mining in the 1840s. He turned nitroglycerin into a less reactive paste, which could be moulded like plasticine. He added an ignition source, and dynamite was born. Dynamite is nitroglycerin with absorbents and stabilisers to ensure it does not combust by the smallest jolt.

Trinitrotoluene (TNT) was first synthesised as a yellow dye in 1863 by Julius Wilbrand. It took a very long time for people to realise its destructive powers. Therefore, TNT was not used as an explosive until some years later. It combusts according to the following thermochemical equation:

$$2C_6H_2(NO_2)_3CH_3(s) \rightarrow 7C(s) + 7CO(g) + 3N_2(g) + 5H_2O(g) \quad \Delta H = -6{,}815\,\text{kJ}$$

TNT is primarily used in military, industrial and mining applications. It is more stable than dynamite as it is insensitive to shock or friction. A large amount of energy needs to be added to start the reaction.

The combustion reaction of TNT has a more negative enthalpy than that of dynamite. This makes TNT the preferred explosive because it generates more energy, is more stable and easier to handle.

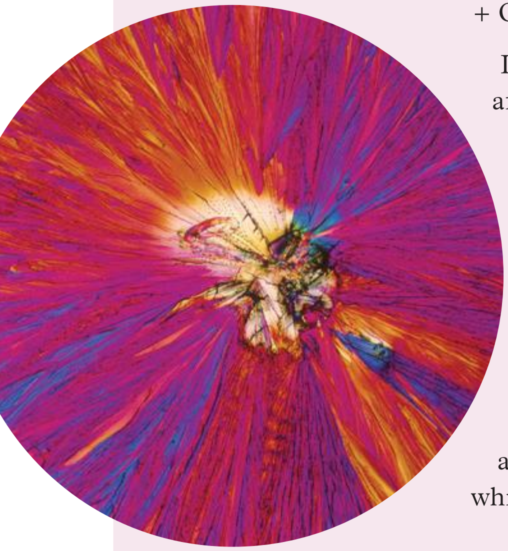


FIGURE 5 Cross-polarised light micrograph of TNT, magnified ×10

Apply your understanding

1 **Use** the thermochemical equations to **calculate** the amount of energy produced by 1 mol of nitroglycerin and 1 mol of TNT. **Use** these values to **determine** how many times more energy is produced per mol of TNT compared to nitroglycerin. (3 marks)

2 **Use** the chemical equations to calculate the amount (in mol) of gases released by 1 mol of nitroglycerin and 1 mol of TNT. **Use** this to **determine** how many times more gas is produced per mol of TNT. (3 marks)

3 The greater enthalpy change when TNT explodes is a big reason why it is preferred over nitroglycerin. It produces hotter gases, which take up much more space than the solid reactant. **Explain** how both the production of more and hotter gases may make TNT a more useful explosive in industry. (2 marks)

Check your learning 8.2

Check your learning 8.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** what enthalpy is. (1 mark)

2 **Recall** the parts of a thermochemical equation. (2 marks)

3 **Recall** the law of conservation of energy. (1 mark)

Analytical processes

4 **Distinguish** between heat and temperature. (1 mark)

5 **Discriminate** between an endothermic and an exothermic reaction. (1 mark)

6 **Contrast** enthalpy diagrams for endothermic and exothermic reactions. (6 marks)

7 **Use** the bond enthalpies from your Formula and data book to answer the following questions about the complete combustion of propene (C_3H_6):

  a **Determine** ΔH for the reaction. (3 marks)

  b **Construct** a fully labelled enthalpy level diagram. (1 mark)

  c **Determine** whether the reaction is endothermic or exothermic. (1 mark)

  d **Determine** the thermochemical equation for this reaction. (2 marks)

8 **Use** the molar enthalpies of combustion from Table 2. **Determine** the thermochemical equations for the combustion reactions of:

  a methane (2 marks)

  b ethyne (2 marks)

  c butane. (2 marks)

9 During an exothermic reaction, the surroundings warm up. **Consider** how this apparent extra energy obeys the law of conservation of energy. (2 marks)

Practical

Lesson 8.3

Applying Hess's law

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Video demonstration

Lesson 8.4

Specific heat capacity

Key ideas

→ Specific heat capacity is a measure of the amount of heat energy it takes to increase the temperature of a specific mass of a substance by 1 K or 1°C.

→ Solution calorimetry is the process of measuring heat absorbed or released during a reaction. It can be used to experimentally determine the specific heat capacity of a substance.

Learning intentions and success criteria

What is specific heat capacity?

Many industries require fuels to operate. When these chemicals are burned or combusted, they release energy into the environment. Chemists and engineers need to be able to calculate the amount of energy released to maximise efficiency of processes used. If the energy released to the environment is used to heat a specific volume of water, the change in temperature of the water can be measured and the energy required to heat the water (which is the energy generated by the fuel when it combusts) can be calculated.

These calculations use a value called **specific heat capacity** (*c*). It is a measure of the amount of heat energy, in joules (J), that it takes to increase the temperature of a specific mass of a substance by 1 kelvin (K) or 1 degree Celsius (°C). A change of 1°C is equivalent to a change of 1 K.

FIGURE 1 Water has a high specific heat capacity, meaning that a relatively large amount of energy is required to increase its temperature by 1°C. This helps to reduce the immediate effects of global warming in terms of temperature increases, as much of the excess heat produced is being absorbed by the oceans.

specific heat capacity
the amount of heat energy, in joules (J), that it takes to increase the temperature of 1 g of a substance by 1 K or 1°C

How do you calculate the energy released from a reaction?

In chemistry, water is a convenient choice as many reactions are carried out in aqueous solution and other reactions can be used to heat a volume of water. This means that reactions occurring in water result in an increase in temperature of the water. The specific heat capacity of water, under ideal conditions, is $4.18\,\text{J}\,\text{g}^{-1}\,\text{K}^{-1}$ (rounded to 2 decimal places). This means that to heat 1 gram of water by 1 K, it requires 4.18 J of energy.

Generally, larger masses of water are used, and the temperature change is more than 1 K. So we can calculate the quantity of heat energy using the formula:

$$Q = mc\Delta T$$

where Q is heat energy in J (not kJ), m is the mass of the water (1 mL = 1 g, for water), c is the specific heat capacity of water ($4.18\,\text{J}\,\text{g}^{-1}\,\text{K}^{-1}$) and ΔT is the change in temperature in K.

Study tip

Since 1 K is equal in magnitude to 1°C, ΔT can be calculated using either unit. An increase in temperature of 25°C to 27°C is the same as an increase in temperature of 298 K to 300 K. ΔT is 2°C or 2 K, and this does not affect the calculation of Q.

Worked example 8.4A

Calculating the quantity of heat energy

Worked example 8.4A: Watch a video that shows how to solve this problem.

A 200.0 mL volume of water was heated in a beaker. The initial temperature of the water was 16°C and the final temperature was 62°C. **Calculate** the energy required to heat the water, in kJ. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to find the heat energy released from the reaction. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required.	$\Delta T = T_f - T_i$ $\Delta T = ?, T_f = 62°C, T_i = 16°C$ or $16\,K$ $Q = mc\Delta T$ $Q = ?, m = 200.0\,g\ (1\,mL = 1\,g), c = 4.18\,J\,g^{-1}\,K^{-1}$
Step 3: Substitute the known values into the formulas and solve for the unknown values.	$\Delta T = 62°C - 16°C$ $= 46°C$ $Q = 200.0\,g \times 4.18\,J\,g^{-1}\,K^{-1} \times 46°C$ (1 mark) $= 38{,}456\,J \times \frac{1\,kJ}{1{,}000\,J}$ $= 38.456\,kJ$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	38 kJ (2 marks)

Your turn

A 40.0 mL volume of water was heated in a beaker. The initial temperature of the water was 22°C and the final temperature was 35°C. **Calculate** the energy required to heat the water, in kJ. (2 marks)

Study tip

Remember that a value of specific heat capacity, $c = 4.18\,J\,g^{-1}\,K^{-1}$, refers specifically to water. Other substances have a different specific heat capacity.

How do you calculate heat of dissolution?

In Lesson 8.2, we talked about the enthalpy or heat of combustion. The heat produced from other reactions can also be calculated. This includes the **heat of dissolution** – that is, the heat produced when a solid is dissolved in water to form an aqueous solution.

It can be assumed that dilute aqueous solutions have the same specific heat capacity as water, since dilute aqueous solutions mostly contain water as the solvent. Therefore, the enthalpy change involved in the temperature change of an aqueous solution can be calculated in the same way, using the formula for Q and the specific heat capacity of water. This will be the energy required to raise the temperature of the solution by a given temperature.

If the amount of the solid being dissolved, n, is known, the heat of solution (ΔH) can be calculated using:

$$\Delta H = \frac{Q}{n}$$

where Q is heat energy in kJ and n is the amount in moles.

heat of dissolution
the enthalpy of a solution when a solid substance dissolves in water

Study tip

Q calculated from $Q = mc\Delta T$ is in J. Since ΔH is typically presented as $kJ\,mol^{-1}$, make sure that you convert Q from J to kJ 1,000 J = 1 kJ.

Coffee cup calorimetry

solution calorimetry
an analytical technique in which the energy generated by a sample is determined by measuring the change in the temperature of an aqueous solution or water

solution calorimeter
a device that measures the change in temperature of water when an endothermic or exothermic reaction takes place

Solution calorimetry is a technique in which the energy (Q) generated by a sample is experimentally determined by measuring the change in the temperature (ΔT) of the solution. This can then be used to find the ΔH.

Calorimetry can be conducted using a **solution calorimeter**. A solution calorimeter is a device that is designed to measure the temperature rise or fall of an aqueous solution or water when an endothermic or exothermic reaction takes place.

A simple solution calorimeter can be made from two insulated foam coffee cups and a foam lid, for maximum heat retention (Figure 2A). There must be a space in the lid to fit a thermometer, to measure temperature change, and a stirrer for even heat distribution. However, since the coffee cups do not retain 100% of the heat of the reaction, some heat is lost to or absorbed from the surroundings. This is a systematic error as long as the same coffee cup is used for all experiments and affects all solutions in the same way.

One way to reduce – but not necessarily eliminate – the systematic error associated with poor insulation is to use a specially designed calorimeter with an insulated container (Figure 2B). Still, the experimental value for ΔH obtained in a school laboratory will always be lower than the theoretical value because simple calorimeters do not retain all of the heat energy in the solution.

A calorimeter can be calibrated using a reaction with a known value of ΔH to measure the amount of heat energy being lost.

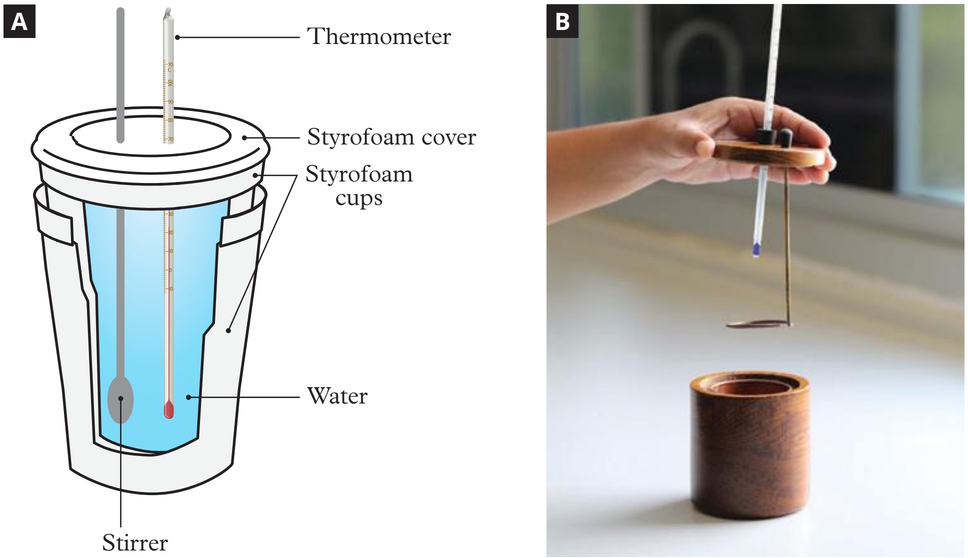


FIGURE 2 The experimental equipment involved in solution calorimetry: (A) a simple calorimeter made using foam cups and (B) a simple calorimeter with a specially designed insulated container

Worked example 8.4B

Calculating the heat of dissolution

Worked example 8.4B: Watch a video that shows how to solve this problem.

When a 3.85 g sample of caesium fluoride (CsF) is added to 40.0 mL of water in a coffee cup calorimeter, a temperature rise of 5.4 K is measured. **Calculate** the heat of dissolution for 1 mol of CsF. (6 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to find the heat of dissolution. The question is worth 6 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Select the correct formula and gather any data required. The *m* in the formula for *Q* is the mass of the water. The *m* in the formula for amount in moles is the mass of the sample.	$Q = mc\Delta T$ $Q = ?$, $m_{\text{water}} = 40.0\,\text{g}$ (1 mL = 1 g), $c = 4.18\,\text{J g}^{-1}\text{K}^{-1}$, $\Delta T = 5.4\,\text{K}$ $n = \frac{m}{M}$ $n = ?$, $m_{\text{CsF}} = 3.85\,\text{g}$, $M = 151.91\,\text{g mol}^{-1}$ $\Delta H = \frac{Q}{n}$ $\Delta H = ?$
Step 3: Substitute the known values into the formula and solve for the heat energy released into the water. The answer is in J.	$Q = 40.0\,\text{g} \times 4.18\,\text{J g}^{-1}\text{K}^{-1} \times 5.4\,\text{K}$ (1 mark) $= 902.88\,\text{J}$ (1 mark)
Step 4: Substitute the known values into the formula and solve for the amount of sample in moles.	$n = \frac{3.85\,\text{g}}{151.91\,\text{g mol}^{-1}}$ (1 mark) $= 0.02534\,\text{mol}$ (1 mark)
Step 5: Substitute the known values into the formula and solve for the heat of dissolution. Remember that *Q* needs to be in kJ, so complete the required conversion.	$\Delta H = \frac{\frac{902.88\,\text{J}}{1{,}000}}{0.2534\,\text{mol}}$ (1 mark) $= 35.625\,\text{kJ mol}^{-1}$
Step 6: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	The temperature has risen during the reaction, so it is an exothermic reaction. The ΔH value will have a negative sign in front, so it is $-36\,\text{kJ mol}^{-1}$. (1 mark)

Your turn

When a 1.00 g sample of magnesium chloride ($MgCl_2$) is added to 50.0 mL of water in a coffee cup calorimeter, a temperature rise of 8.0 K is measured. **Calculate** the heat of dissolution for 1 mol of $MgCl_2$. (6 marks)

How do you calculate heat of neutralisation?

Neutralisation reactions occur between an acid and a base to produce a neutral solution. You will learn more about acids, bases and neutralisation in Modules 14 and 15. Typically, these occur in water as an aqueous solution of acid is mixed with an aqueous solution of base. Thus, the same formulas and the specific heat capacity of water can be used to find the **heat of neutralisation**.

heat of neutralisation the enthalpy of a solution when an acid reacts with a base

The main differences to remember in these calculations are:

- the temperature of the two solutions should be measured to ensure they are the same before mixing. This gives a precise initial temperature.
- The final volume of solution after mixing should be used when calculating *Q*, as it is this total volume of water that is being heated.

Worked example 8.4C

Calculating the heat of neutralisation

Worked example 8.4C: Watch a video that shows how to solve this problem.

A 50.0 mL of a solution of 1.0 M NaOH was mixed with a 50.0 mL of a solution of 1.0 M CH_3COOH. The initial temperatures of both solutions were 22.2°C. Each 50.0 mL of 1.0 M solution contained 0.050 mol of the dissolved reactant. After mixing, the highest temperature measured was 28.7°C.

a **Calculate** the heat of neutralisation in kJ mol^{-1}. (2 marks)

b **Determine** the thermochemical equation for the reaction. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to find the heat of neutralisation and then write a thermochemical equation. Each question is worth a variety of marks, so we must gather the correct information, correctly apply the formulas, complete the calculations and express the reaction correctly.
Step 2: Select the correct formula and gather any data required. The *m* in the formula for *Q* is the mass of the water. The *m* in the formula for amount in moles is the mass of the sample.	$\Delta T = T_f - T_i$ $\Delta T = ?$, $T_f = 28.7°C$, $T_i = 22.2°C$ $Q = mc\Delta T$ $Q = ?$, $m_{water} = 50.0\,g + 50.0\,g$ (1 mL = 1 g), $c = 4.18\,J\,g^{-1}\,K^{-1}$ $\Delta H = \frac{Q}{n}$ $\Delta H = ?$, $n = 0.050\,mol$
Step 3: Substitute the known values into the formula and solve for the change in temperature.	$\Delta T = 28.7°C - 22.2°C$ $= 6.5°C$ or $6.5\,K$
Step 4: Substitute the known values into the formula and solve for the heat energy released into the water. The answer is in J.	$Q = (50.0 + 50.0)\,g \times 4.18\,J\,g^{-1}\,K^{-1} \times 6.5°C$ (1 mark) $= 2{,}717\,J$ (1 mark)
Step 5: Substitute the known values into the formula and solve for the heat of neutralisation. Remember that *Q* needs to be in kJ, so complete the required conversion.	$\Delta H = \frac{\frac{2{,}717\,J}{1{,}000}}{0.050\,mol}$ (1 mark) $= 54.34\,kJ\,mol^{-1}$
Step 6: Finalise your answer for part **a** and make sure you have expressed it using the correct units and number of significant figures.	The temperature has risen during the reaction, so it is an exothermic reaction. The ΔH value will have a negative sign in front, so it is $-54\,kJ\,mol^{-1}$. (1 mark)
Step 7: For part **b**, write a thermochemical equation for the combustion reaction, including states. Refer to Worked example 8.2C for a refresher. Remember that the ΔH must be expressed in kJ, so take into account any coefficients in the equation.	$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$ $\Delta H = -54\,kJ$ (1 mark for balanced equation; 1 mark for correct ΔH)

Your turn

A 20.0 mL of a solution of 2.0 M NaOH was mixed with a 20.00 mL of a solution of 2.0 M CH_3COOH. The initial temperatures of both solutions were 24.5°C. Each 2,050.0 mL solution contained 0.04 mol of the dissolved reactant. After mixing, the highest temperature measured was 38.2°C.

a **Calculate** the heat of neutralisation in kJ mol^{-1}. (4 marks)

b **Determine** the thermochemical equation for the reaction. (2 marks)

How do you calculate the heat of combustion?

The enthalpy change in a combustion reaction (or the heat of combustion) for a liquid fuel can be calculated in a similar way as for aqueous solutions. However, the methodology is quite different because a flame is used.

Figure 3 shows a spirit burner, which effectively burns a fuel while keeping the flame contained. The mass of the spirit burner is recorded before and after the experiment to determine the mass of the fuel used in the experiment to heat the water. A set volume of water is contained within the beaker or metal can, with a thermometer and stirrer. As the mass of the fuel and the temperature rise of the water are known, a ΔH value can be calculated.

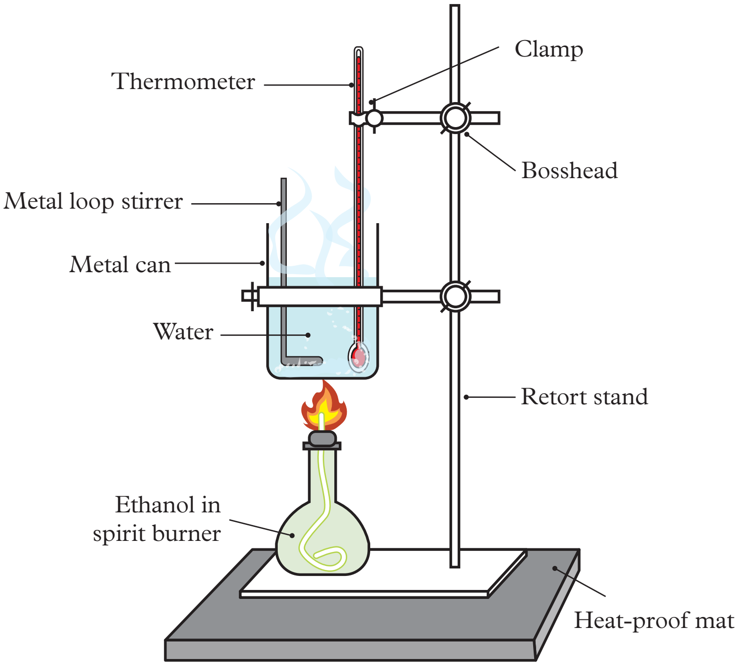


FIGURE 3 A spirit burner set up to record the enthalpy change in a combustion reaction

Study tip

Only alcohol-based fuels (e.g. methanol, ethanol) can be burned in a spirit burner.

Worked example 8.4D

Calculating the heat of combustion

Worked example 8.4D: Watch a video that shows how to solve this problem.

A 4.60 g sample of ethanol (C_2H_5OH) was burnt using a spirit burner. The energy produced by the flame was used to heat a beaker containing 150.0 mL of water. The starting temperature of the water was 18.0°C and the final temperature was 68.0°C. **Calculate** the heat of combustion for 1 mol of ethanol. (6 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to find the heat of combustion. The question is worth 6 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.

Think	Do
Step 2: Select the correct formula and gather any data required. The *m* in the formula for *Q* is the mass of the water. The *m* in the formula for amount in moles is the mass of the sample.	$\Delta T = T_f - T_i$ $\Delta T = ?,\ T_f = 68.0°C,\ T_i = 18.0°C$ $Q = mc\Delta T$ $Q = ?,\ m_{water} = 150.0\,g\ (1\,mL = 1\,g),\ c = 4.18\,J\,g^{-1}\,K^{-1}$ $n = \frac{m}{M}$ $n = ?,\ m_{ethanol} = 4.60\,g,\ M = 46.08\,g\,mol^{-1}$ $\Delta H = \frac{Q}{n}$ $\Delta H = ?$
Step 3: Substitute the known values into the formula and solve for the change in temperature.	$\Delta T = 68.0°C - 18.0°C$ $= 50°C$ or $50\,K$
Step 4: Substitute the known values into the formula and solve for the heat energy released into the water. The answer is in J.	$Q = 150.0\,g \times 4.18\,Jg–1\,K^{-1} \times 50\,K$ (1 mark) $= 31{,}350\,J$ (1 mark)
Step 5: Substitute the known values into the formula and solve for the amount of sample in moles.	$n = \frac{4.60\,J}{46.08\,J\,mol^{-1}}$ (1 mark) $= 0.09983\,mol$ (1 mark)
Step 6: Substitute the known values into the formula and solve for the heat of neutralisation. Remember that *Q* needs to be in kJ, so complete the required conversion.	$\Delta H = \frac{\frac{31{,}350\,J}{1{,}000}}{0.09983\,mol}$ (1 mark) $= 314.05\,kJ\,mol^{-1}$
Step 6: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	The temperature has risen during the reaction, so it is an exothermic reaction. The ΔH value will have a negative sign in front, so it is $-314\,kJ\,mol^{-1}$. (1 mark)

Your turn

A 1.68 g sample of 1-octanol was burnt using a spirit burner. The energy produced was used to heat a beaker filled with 250 g of water. The starting temperature of the water was 14.9°C and the final temperature was 36.4°C. **Calculate** the molar heat of combustion for 1-octanol. (6 marks)

Skill drill

Designing an experiment to determine heat of combustion

Science inquiry skill(s): Planning investigations (Lesson 1.4)

The experimental set-up for measuring the enthalpy change in a combustion reaction (or the heat of combustion) was shown in Figure 3. Consider an experiment where you have access to the following materials:

- Thermometer
- Heatproof mat
- Retort stand with boss head and clamp
- Beaker (or metal can)
- Measuring cylinder
- Distilled water
- Metal loop stirrer
- Electronic balance
- Spirit burner
- Fuels (methanol, ethanol, 1-propanol)

Practise your skills

1 **Design** an experiment to measure and calculate the heat of combustion for various fuels using the listed materials and the set-up shown in Figure 3. In your response:
   a **Define** an aim and hypothesis. (4 marks)
   b **Describe** your method. (2 marks)
   c **Conduct** a risk assessment. (3 marks)
   d **Create** a results table to record your results. (2 marks)

Real-world chemistry

The CORI cycle: energy when it's needed most

All living organisms, from a single-celled organism to a blue whale, require energy to function. The main source of energy for cells is a molecule called glucose. It undergoes a series of reactions in the mitochondria of cells. These reactions are summarised using the overall cellular respiration equation, which is exothermic:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \qquad \Delta H = -2{,}802.7\,kJ$$

This reaction requires 6 mol of oxygen for every 1 mol of glucose. Because it relies on oxygen, it is called an aerobic reaction.

Getting oxygen to the cells rapidly is vital, particularly during intense muscle activity. However, there are times, such as during a sprint, when oxygen cannot be supplied quickly enough but the muscles still require energy. Under these circumstances, anaerobic respiration occurs (in the absence of oxygen). The muscle cells convert glucose to lactic acid instead, which releases just 7% of the energy normally released in respiration.

Lactic acid build-up is not good for the muscles or the rest of the body, as it causes pain. To overcome this, it is quickly transported back to the liver where it is converted back to glucose, using a great deal of energy from cellular respiration. As a result, after intense exercise, the demand for oxygen is increased as the body experiences "oxygen debt".

The cycle in which glucose is transported to the cells and used in low oxygen conditions to make lactic acid, which is returned to the liver for conversion back into glucose is called the Cori cycle, after Austrian-American biochemist, Gerty Theresa Cori (1896–1957). Gerty contributed a great deal of research to how the body metabolises carbohydrates for energy.

For the discovery of the Cori cycle, Gerty and her husband Carl Cori, shared the 1947 Nobel Prize in Physiology or Medicine. Gerty is acknowledged as the first woman to ever receive this particular Nobel Prize. Her lifelong research led to knowledge that informed the treatment of diabetes, as well as the problems with a certain genetic enzyme disease.

FIGURE 4 Gerty Cori

Apply your understanding

1 The formula for lactic acid is $C_3H_6O_3$. **Infer** the number of molecules of lactic acid that could be made from one molecule of glucose. (1 mark)

2 **Calculate** the amount, in mol, of glucose that must form lactic acid to release the same amount of energy as 1 mol of glucose in aerobic respiration. (2 marks)

3 After intense exercise, athletes will breathe faster and have a raised heartbeat for some time, even though they are resting. **Explain** why. (3 marks)

Check your learning 8.4

Check your learning 8.4: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define** "specific heat capacity". (1 mark)

2 **Identify** what each symbol represents and the units used for each quantity, in the formula $Q = mc\Delta T$. (4 marks)

Analytical processes

3 **Compare** the two techniques described in Lesson 8.4 for measuring ΔH for aqueous solutions and combustion reactions. (2 marks)

4 **Calculate** the energy required to heat 72.0 mL of water, if the initial temperature of the water is 273 K and the final temperature is 17°C. Convert °C to K by adding 273. (2 marks)

5 **Calculate** the heat of dissolution of magnesium nitrate ($Mg(NO_3)_2$) if a temperature rise of 0.85°C is measured after a 0.72 g sample of the salt is added to 123 mL of water in a coffee cup calorimeter. (6 marks)

6 A 10.9 g sample of ethanol (C_2H_6O) is burnt to heat a beaker containing 320 mL of water. The initial temperature of the water was 273 K and the final temperature was 72°C. **Determine** the heat of combustion for 1 mol of ethanol. (6 marks)

Knowledge utilisation

7 **Investigate** and communicate the systematic errors and limitations of the spirit burner experiment when experimentally determining the ΔH of fuels. (3 marks)

8 **Discuss** the systematic errors involved with coffee cup calorimetry and spirit burner combustion measurements. **Evaluate** both methods and communicate improvements that could be made to each method, based on the errors identified. (5 marks)

Practical

Lesson 8.5

Measuring the enthalpy of a reaction using calorimetry

Learning intentions and success criteria

Video demonstration

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Practical

Lesson 8.6

Measuring the heat of combustion

Learning intentions and success criteria

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

MODULE 8

Lesson 8.7

Review: Chemical reactions

Summary

8.1
- Phase changes are physical changes that involve changes in energy.
- Chemical reactions also involve energy changes, and include single displacement, double displacement, acid–base, combustion, combination, decomposition and simple redox reactions.
- Chemical reactions can be represented using balanced chemical equations, including states.

8.2
- Endothermic reactions involve the absorption of heat from the environment and a decrease in temperature. Exothermic reactions involve the release of heat into the environment and an increase in temperature.
- Enthalpy, H, is the total heat content of a substance.
- Enthalpy level diagrams are used to show the change in heat energy before, during and after a reaction.
- Thermochemical equations are balanced chemical equations that also show enthalpy.

8.3
- Practical: Applying Hess's law

8.4
- Specific heat capacity is a measure of the amount of heat energy it takes to increase the temperature of a specific mass of a substance by 1 K or 1°C.
- Solution calorimetry is the process of measuring heat absorbed or released during a reaction. It can be used to experimentally determine the specific heat capacity of a substance.

8.5
- Practical: Measuring the enthalpy of a reaction using calorimetry

8.6
- Practical: Measuring the heat of combustion

Key formulas

Enthalpy change	$\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma H_{\text{bonds formed}}$
Specific heat capacity	$Q = mc\Delta T$

Review questions 8.7A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 The enthalpy for the formation of NO_2 from elements is given by:

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g) \quad \Delta H = +132\,kJ$

What is the enthalpy change for the following reaction?

$$4NO_2(g) \rightarrow 2N_2(g) + 4O_2(g)$$

A +66 kJ
B +264 kJ
C –264 kJ
D –132 kJ

2 How much energy is released when 1.01101 g of hydrogen reacts with excess oxygen?

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad \Delta H = -571.6\,kJ$

A 142.9 kJ
B 571.6 kJ
C 14,290 kJ
D 28,500 kJ

3 The molar enthalpy of combustion of glucose, $C_6H_{12}O_6$ is –2,840 kJ mol^{-1}. The energy released when 1 g of glucose undergoes complete combustion is

A 511,700 kJ.
B 15.76 kJ.
C 94.57 kJ.
D 85,300 kJ.

4 Which of the following lists contains only endothermic processes?

A Boiling, evaporation, condensation, deposition
B Sublimation, deposition, melting, boiling
C Freezing, deposition, condensation, melting
D Melting, boiling, evaporation, sublimation

5 A student carried out a calorimetry experiment and found that the experimental heat of combustion was 55% of the value from an online reference source. The most significant reason for the error would be

A the values calculated using bond energies are not precise.
B it is difficult to measure the temperatures precisely using a thermometer.
C it is difficult to measure the mass of fuel used precisely as it is in a spirit burner.
D some of the heat of combustion heats surroundings rather than heating the water.

6 The reaction

$Pb(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$

can be classified as both

A a single displacement reaction and a combination reaction.
B a combination reaction and a decomposition reaction.
C a single displacement reaction and a redox reaction.
D a combination reaction and a redox reaction.

Review questions 8.7B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

7 **Identify** the difference between heat and energy. (2 marks)

8 **Explain** how endothermic and exothermic reactions relate to the law of conservation of energy and the breaking and reforming of bonds. (4 marks)

9 **Sketch** enthalpy level diagrams for exothermic and endothermic reactions, and include labels for enthalpy of reactants, enthalpy of products and ΔH. (6 marks)

10 **Explain**, in terms of average bond energies, why:

a the reactants and products in chemical reactions have different enthalpies (3 marks)
b some reactions are exothermic and others are endothermic (3 marks)
c exothermic and endothermic reactions do not break the law of conservation of energy. (3 marks)

11 **Identify** three limitations of using bond energies to calculate enthalpies of reaction. (3 marks)

12 A 1.2 g sample of octane (C_8H_{18}) is burnt to raise the temperature of 50.0 mL of water by 30.0°C. **Calculate** the molar enthalpy change of the combustion reaction for octane. (6 marks)

Analytical processes

13 **Discriminate** between exothermic and endothermic reactions. (1 mark)

14 **Consider** each of these reactions. **Determine** balanced equations for each, including state symbols and **classify** them by their type of reaction.

a Pentane burns in excess oxygen, forming carbon dioxide and water vapour. (2 marks)
b Silver nitrate solution reacts with magnesium chloride solution to form insoluble silver chloride and soluble magnesium nitrate solution. (2 marks)
c Nitric acid solution reacts with calcium hydroxide solution to form calcium nitrate solution and liquid water. (2 marks)

d An aqueous solution of chlorine reacts with a solution of potassium bromide forming a solution containing bromine and potassium chloride. (2 marks)

15 Sodium hydroxide (NaOH) and hydrochloric acid (HCl) form sodium chloride (NaCl) and water (H_2O), where the relative enthalpy values of the reactants is given as 703 kJ mol^{-1} and that of the products is given as 646.7 kJ mol^{-1}. **Construct** an enthalpy level diagram for the following chemical reaction. **Determine** the change in enthalpy and **identify** whether the reaction is endothermic or exothermic. (3 marks)

16 **Determine** the energy required to break 1 mol of propane, $CH_3CH_2CH_3$, into its constituent atoms, from its bond enthalpies. (1 mark)

17 **Determine** the amount of energy required to break 1 mol of ethanol into its constituent atoms. (1 mark)

18 Ethanol (CH_3CH_2OH) undergoes combustion in the presence of oxygen to form carbon dioxide and water vapour.

a Use bond enthalpies to **determine** the enthalpy of combustion of ethanol. (5 marks)

b **Compare** this with the value given in Table 2 from Lesson 8.2. **Determine** which of the two values is likely to be more valid and **explain** why. (3 marks)

19 **Use** the data in Table 2 from Lesson 8.2 to **determine** the thermochemical equations for the combustion reactions of:

a ethane (2 marks) **b** propane. (2 marks)

20 A 0.55 g sample of methanol (CH_3OH) is burnt using a spirit burner. The energy produced by the flame was used to heat a conical flask containing 50.0 mL of water. The starting temperature of the water was 25.3°C and the final temperature was 35.8°C. **Determine** the enthalpy change for combustion of 1 mol of methanol. (6 marks)

Data drill

Molar heats of combustion of alkanes

Alkanes are molecules that have the general formula C_nH_{2n+2}. Primary alcohols consist of single-bonded carbon and hydrogen atoms, with an –OH group bonded to the first carbon atom in place of one of the hydrogen atoms.

Both can react with oxygen in combustion reactions. Figure 1 shows the molar enthalpies of combustion of some alkanes and alcohols.

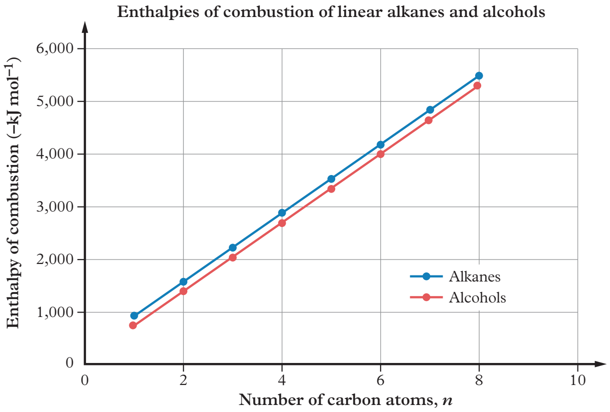


FIGURE 1 Molar heats of combustion (kJ mol^{-1}) for various alkanes and primary alcohols

Apply understanding

1 **Identify** the molar heat of combustion of methane, the alkane with one carbon atom. (1 mark)

2 **Calculate** the difference in the molar heats of combustion of ethane, with two carbon atoms and methane. (1 mark)

Analyse evidence

3 **Identify** the trend between the number of carbon atoms in the alkanes and their molar heats of combustion. Provide supporting evidence. (2 marks)

4 **Compare** the trend in the heats of combustion of alcohols with that of alkanes. Provide supporting evidence. (2 marks)

Interpret evidence

5 **Predict** the molar heat of combustion for decanol, an alcohol with 10 carbon atoms. **Justify** your prediction. (2 marks)

Module 8 checklist: Chemical reactions

Quizlet: Revise key terms online to test your understanding

UNIT 1

Topic 3 review

Multiple choice

(1 mark each)

1 Atropine is used in eyedrops to dilate pupils. It has the molecular formula $C_{17}H_{23}NO_3$. What is the percentage composition of oxygen in atropine?

A 14%

B 17%

C 20%

D 6.8%

2 Which of the following represents a double displacement reaction between iron(III) chloride and silver nitrate?

A $FeCl_3(aq) + AgNO_3(aq) \rightarrow FeNO_3(aq) + AgCl(s)$

B $FeCl_3(aq) + 3AgNO_3(aq) \rightarrow Fe(NO_3)_3(aq) + 3AgCl(s)$

C $FeCl_3(aq) + 2AgNO_3(aq) \rightarrow FeCl_2NO_3(aq) + Ag_2NO_3Cl(aq)$

D $2FeCl_3(aq) + 6AgNO_3(aq) \rightarrow 6Ag(s) + 3Cl_2(g) + 2Fe(NO_3)_3(aq)$

3 Thionyl chloride ($SOCl_2$) can be prepared using the following reaction in Figure 1:

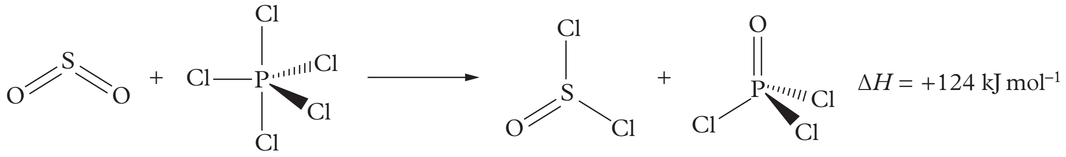


FIGURE 1 Thionyl chloride

Based on the data from the table, what is the average bond enthalpy for the S–Cl bond?

TABLE 1 Bond enthalpy

Bond	S–Cl	P–Cl	S=O	P=O
Bond enthalpy (kJ mol⁻¹)	to be determined	326	522	544

A 138 kJ mol⁻¹

B 253 kJ mol⁻¹

C 304 kJ mol⁻¹

D 506 kJ mol⁻¹

4 Which of the following about 1 mol of Ag and 1 mol of Au is true?

A They are equal in mass.

B Their molar masses are equal.

C They have the same atomic mass.

D They contain the same number of atoms.

5 The molecular formula for vitamin C is $C_6H_8O_6$. The empirical formula is

A CHO.

B CH_2O.

C $C_3H_4O_3$.

D $C_2H_4O_2$.

6 A 10.0 mol sample of a compound was found to have a mass of 510.0 g. The compound has a molar mass of

A 10.0 $g\,mol^{-1}$.
B 510.0 $g\,mol^{-1}$.
C $\frac{510.0}{10.0}\,g\,mol^{-1}$.
D $(10.0 \times 510.0)\,g\,mol^{-1}$.

7 Which of the following is an endothermic chemical reaction?

A A candle burning
B Cellular respiration
C Melting of ice blocks in a glass of water
D Condensation of water vapour on a cold surface

8 The enthalpy diagram for the formation of 1 mol of ammonia from its elements is shown in Figure 2.

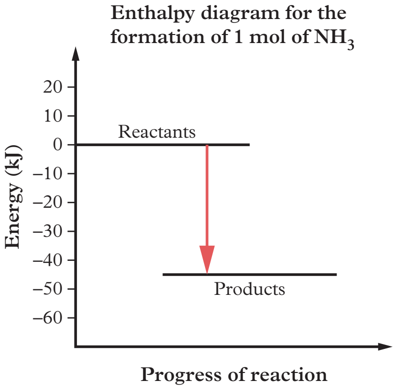


FIGURE 2 Enthalpy diagram

The thermochemical equation for this reaction is

A $N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g) \quad \Delta H = 46\,kJ$
B $\frac{1}{2}N_2\,(g) + \frac{3}{2}H_2\,(g) \rightarrow NH_3\,(g) \quad \Delta H = 46\,kJ$
C $N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g) \quad \Delta H = -46\,kJ$
D $\frac{1}{2}N_2\,(g) + \frac{3}{2}H_2\,(g) \rightarrow NH_3\,(g) \quad \Delta H = -46\,kJ$

9 What is the empirical formula for the compound that is 63.7% N and 36.3% O?

A NO
B N_2O
C NO_2
D N_4O_2

10 A compound has the empirical formula of C_3H_2Cl and a molar mass of 147.0 $g\,mol^{-1}$. The molecular formula is

A CHCl.
B C_3H_2Cl.
C $C_{12}H_8Cl_4$.
D $C_6H_4Cl_2$.

11 The molar mass of sodium chloride is 58.44 $g\,mol^{-1}$. This means that

A a crystal of sodium chloride has a mass of 58.44 g.
B 58.44 g of NaCl contains 6.02×10^{23} atoms altogether.
C a formula unit of sodium chloride has a mass of 58.44 u.
D a molecule of sodium chloride has a molecular mass of 58.44 u.

12 Aluminium has a specific heat capacity of 0.9 $J\,g^{-1}\,K^{-1}$. A small aluminium block of mass 21.6 g was heated from 25.0°C to 32.1°C. The quantity of heat involved was

A 138.0 J released by the block.
B 19.44 J released by the block.
C 19.44 J absorbed by the block.
D 138.0 J absorbed by the block.

13 Which of the following statements about temperature and heat is correct?

A Temperature is a measure of the average kinetic energy of the particles and heat is a form of energy.
B Temperature measures the bond enthalpy in a substance and heat measures the change in enthalpy during a reaction.
C Temperature is a measure of the average kinetic energy of the particles and heat is the change in enthalpy during a reaction.
D Temperature is a measure of the total amount of energy in a substance and heat is the total potential and kinetic energy of the particles.

14 The law of conservation of mass states that

A mass is neither created nor destroyed in a physical process.
B mass is neither created nor destroyed in a chemical reaction.
C mass and energy are neither created nor destroyed in a physical process.
D mass and energy are neither created nor destroyed in a chemical reaction.

15 The molar mass of copper sulfate pentahydrate, $CuSO_4{\bullet}5H_2O$ is

A 249.71 u.
B 159.61 u.
C 249.71 $g\,mol^{-1}$.
D 159.61 $g\,mol^{-1}$.

Short response

16 **Calculate**:

a the amount, in mol, of sucrose commonly known as granulated sugar, $C_{12}H_{22}O_{11}$, in a 1.00 kg packet of sugar (2 marks)

b the number of molecules of sucrose that this represents. (2 marks)

17 **Calculate** the number of Cu^{2+} ions and Cl^{-} ions in 1.00 g of $CuCl_2$. (4 marks)

18 The thermochemical equation for the neutralisation of a solution of NaOH by a solution of HCl is:

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) \quad \Delta H = -55.89\,kJ$$

100.0 mL of solution containing 0.2 mol of NaOH and 100.0 mL of solution containing 0.2 mol of HCl were mixed. It can be assumed that the total volume is the sum of the two volumes. Both solutions had a temperature of 21.5°C before mixing.

a **Identify** the heat released by the neutralisation reaction. (1 mark)

b **Calculate** the final temperature of the combined solutions. (3 marks)

19 Aspirin can be made in the laboratory by reacting salicylic acid with an excess of ethanoic anhydride. The reaction is shown in Figure 3.

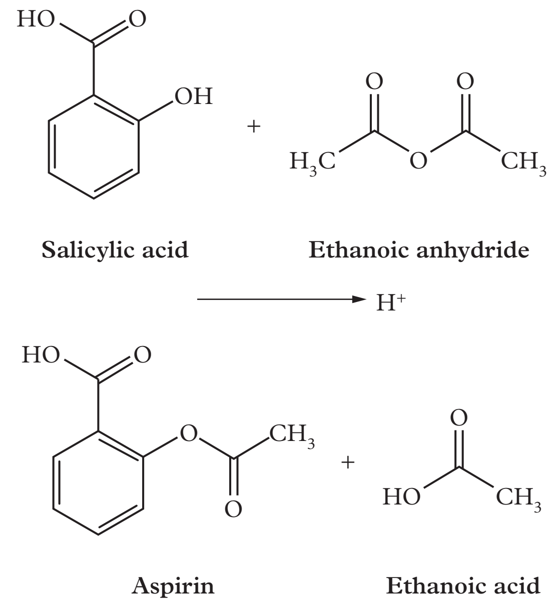


FIGURE 3 Aspirin

The molar masses of salicylic acid and aspirin, respectively, are 138.12 g mol^{-1} and 180.16 g mol^{-1}. A chemist uses 2.00 g of salicylic acid in a reaction mixture and 1.85 g of aspirin was formed.

a **Calculate** the theoretical yield of aspirin. (4 marks)

b **Calculate** the percentage yield of aspirin. (2 marks)

20 A spirit burner containing propan-1-ol, C_3H_8O, is used to heat 80.0 g of water in a test tube. The mass of the spirit burner is 0.25 g less at the end of the reaction and the temperature of the water rose by 15.1°C.

a **Calculate** the quantity of heat absorbed by the water. (2 marks)

b **Calculate** the experimental value for the molar heat of combustion of propan-1-ol. (4 marks)

21 Lead(II) nitrate and potassium iodide solutions react to form a precipitate of lead(II) iodide. An investigation was carried out to determine how the mass of lead(II) iodide precipitated depended on the mass of lead(II) nitrate that reacted with 2.00 g of potassium iodide, when both masses were dissolved and mixed. The results were plotted on a graph in Figure 4.

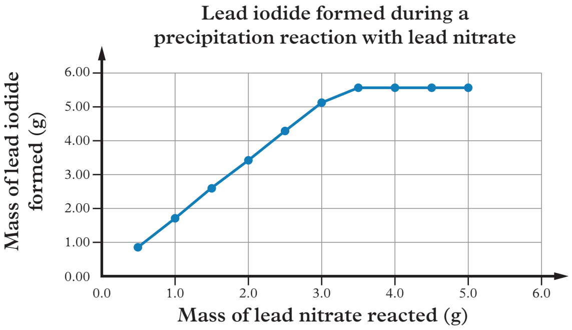


FIGURE 4 Lead iodide formed during a precipitation reaction

a **Determine** the balanced equation for the reaction. (1 mark)

b **Describe** the trends shown in the results. (2 marks)

c **Propose** a reason why increasing the mass of lead nitrate did not lead to a continued increase in the mass of precipitate. (1 mark)

22 Copper is obtained from mineral-bearing ores. One mineral commonly found is covellite, consisting of copper(II) sulfide. 1.00 tonne of a particular ore body contains 1.1% CuS. (1 tonne = 1,000 kg)

After separation of the mineral from large amounts of waste rock in the ore, the copper(II) sulfide is roasted in air and the copper produced is solidified. It reacts with oxygen as shown:

$$CuS(s) + O_2(g) \rightarrow Cu(s) + SO_2(g)$$

a **Calculate** the mass of CuS present in 1.00 tonne of ore. Express your answer in grams. (2 marks)

b **Calculate** the mass of copper that could be obtained by roasting the copper(II) sulfide. (4 marks)

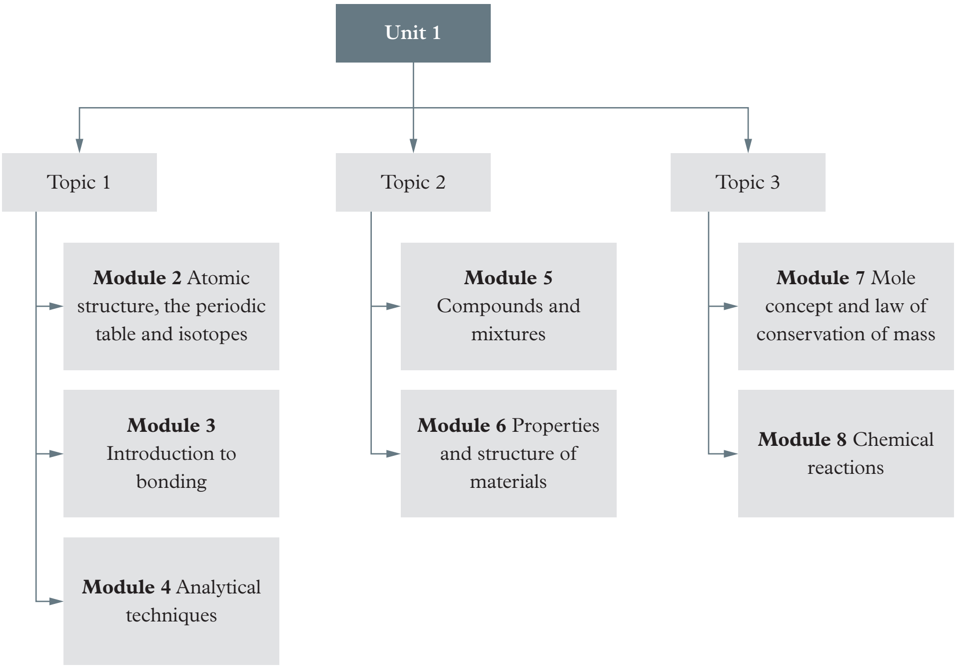
TOTAL MARKS

/49

UNIT 1 Review

Part A – Revisit and revise

Part A of the Unit review asks you to reflect on your learning and identify areas in which you need more work.



Part B – Exam essentials

Now that you've completed your revision for Unit 1, it's time to learn and practise some of the skills you'll need to answer exam questions like a pro! To help you, our expert authors have created the following advice and tips to help you maximise your results on the end-of-year examination.

Exam tip 1: Look at mark allocation and space given

Use the number of marks and the space provided to guide the detail expected in your response.

- For each mark, provide a separate point and elaborate as required.
- For average handwriting, the number of answer lines provided is a guide to how much to write.

See it in action

Read the exam-style question below and see how the tip has made a difference between a response that has scored full marks and a response where marks have been lost.

QUESTION 1 (5 marks)

Silicon and carbon appear in the same column of the periodic table. Despite this, the compounds carbon dioxide and silicon dioxide are very different. Carbon dioxide is a colourless gas, whereas silicon dioxide is a hard solid with a melting point of 1,670°C. Explain why these compounds are so different.

Complete response

Clearly identifies/classifies the type of substance for both compounds [1 mark]

Describes the strong covalent bonding and the weak intermolecular forces in carbon dioxide [1 mark]

Carbon dioxide is a covalent molecular substance and silicon dioxide is a covalent network substance. Carbon dioxide has strong covalent bonds within each molecule but only weak intermolecular forces between the molecules. It takes very little energy to overcome the weak intermolecular forces, so carbon dioxide is a gas at room temperature. Silicon dioxide has strong covalent bonds extending throughout the solid crystal lattice. It takes a great deal of energy to overcome the strong covalent bonds, to scratch or break so it is hard; and to melt so very high melting point.

Relates the hardness and high melting point to the strong covalent bonding [1 mark]

Describes the strong covalent bonding throughout the crystal lattice in silicon dioxide [1 mark]

Relates the low boiling point to how it is easy to overcome weak intermolecular forces [1 mark]

Incomplete response

Correctly identifies the relationship but does not explain why; does not describe the nature of the bonding/ intermolecular forces [0 marks]

Clearly identifies/classifies the type of substance for both compounds [1 mark]

Carbon dioxide is a covalent molecular substance. Silicon dioxide is a covalent network substance. It is known that covalent molecular substances have low melting points. It is known that covalent network substances are generally hard as it is difficult to break the strong covalent bonds. It is also known that covalent network substances have very high melting points, due to the energy required to break covalent bonds.

Provides some explanation of the properties, by linking to the strong covalent bonds but provides no description of the bonding in covalent network substances [1 mark]

Think like an assessor

To maximise your marks on an exam, it can help to think like a QCAA assessor. Consider how many marks each question is worth and what information the assessor is looking for.

A student has given the following response in a practice exam. Imagine you are a QCAA assessor and use the marking guide below to mark the response.

QUESTION 2 (5 marks)

The table shows the first, second and third ionisation energies for selected elements in the third period of the periodic table. Sodium shows a large increase from the first to the second ionisation energy, whereas magnesium shows a large increase from the second to the third ionisation energy. Silicon does not show a large increase.

	Element		
	Na	**Mg**	**Si**
First ionisation energy (kJ mol^{-1})	496	738	787
Second ionisation energy (kJ mol^{-1})	4,653	1,450	1,577
Third ionisation energy (kJ mol^{-1})	6,910	7,730	3,232

Explain the large increases for sodium and magnesium, and why silicon does not show this.

It is easier to remove the first electron from sodium, and then there is a large increase to remove the second and third electrons. Sodium has 1 outer electron in the 3s orbital, which uses less energy to remove, and then electrons are being removed from orbitals closer to the nucleus which requires more energy. The same idea can be used for Mg. Silicon is different, because it is not a metal and it has four valence electrons.

Marking guide

Question 2

- identifies that the electron configuration of Na has 1 electron in its outermost energy level/set of orbitals (can write full electron configuration, or refer to 1 valence electron). [1 mark]
- identifies the electron configuration of Mg [1 mark]
- identifies the electron configuration of Si [1 mark]
- explanation includes:
  - takes less energy to remove any electrons from the higher energy levels/outermost orbitals
  - much more energy to remove the electrons in orbitals closer to the nucleus
  - so the big jump in ionisation energy occurs after all the valence electrons are removed [1 mark]
- identifies that for Na and Mg, the big increase in ionisation energy occurs after 1 or 2 electrons have been removed, respectively, whereas for Si, it would be after 4 electrons were removed, which is not shown in the table. [1 mark]

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved? Write your own response to the same question to receive full marks from a QCAA assessor.

Exam tip 2: Show all your working for calculation questions

The marks are in the details. When calculating numerical answers, show the substitution of numbers into formulas, followed by the final answer. This working is marked.

- Writing the rule is not awarded a mark but is good practice and helps prevent errors.
- Mentally check that the magnitude (size) of the answer makes sense.
- Include units if appropriate for the answer. Some marking schemes require correct units.

See it in action

Read the exam-style question below and see how the tip has made a difference between a response that has scored full marks and a response where marks have been lost.

QUESTION 3 (5 marks)

A 100.0 g sample of propane, $C_3H_8(g)$, was burnt in sufficient oxygen to ensure complete combustion. The thermochemical equation for the reaction is shown.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H = -2{,}220 \text{ kJ}$$

a) Calculate the amount, in mol, of propane burnt. *[3 marks]*

b) Calculate the heat released by the combustion of 100.0 g of propane. *[2 marks]*

Complete response

a) Calculate the amount, in mol, of propane burnt. *[3 marks]*

$M(C_3H_8) = 3 \times 12.01 + 8 \times 1.01 = 44.11 \text{ g mol}^{-1}$ Calculates correct molar mass [1 mark]

$n(C_3H_8) = \frac{m}{M}$

$= \frac{100.0 \text{ g}}{44.11 \text{ g mol}^{-1}}$ Completes correct substitution [1 mark]

$= 2.267 \text{ mol}$ of propane

Calculates correct amount, in mol [1 mark]

b) Calculate the heat released by the combustion of 100.0 g of propane. [*2 marks*]

heat released = n × heat of reaction (without negative sign)

Identifies that they need to multiply the amount, in mol, by the absolute (positive) value of the heat of reaction, since the question asks for heat released [1 mark]

$= 2.267 \text{ mol} \times 2{,}220 \text{ kJ mol}^{-1}$

$= 5{,}033 \text{ kJ}$

Calculates correct heat released [1 mark]

Incomplete response

a) Calculate the amount, in mol, of propane burnt. [*3 marks*]

$M(C_3H_8) = 3 \times 12.01 + 8 \times 1.01 = 44.11 \text{ g mol}^{-1}$

Calculates correct molar mass [1 mark]

$n(C_3H_8) = \frac{m}{M}$

$= 2.267$ mol of propane

Calculates correct amount, in mol, [1 mark] but does not show working [0 marks]

b) Calculate the heat released by the combustion of 100.0 g of propane. [*2 marks*]

heat released = n × heat of reaction

$= 2.267 \text{ mol} \times -2{,}220 \text{ kJ mol}^{-1}$

Multiplies the amount, in mol, but includes the negative sign for the heat of reaction [0 marks]

$= -5{,}033 \text{ kJ}$

Arrives at a consequentially incorrect answer [1 mark]

Think like an assessor

To maximise your marks on an exam, it can help to think like a QCAA assessor. Consider how many marks each question is worth and what information the assessor is looking for.

A student has given the following response in a practice exam. Imagine you are a QCAA assessor and use the marking guide below to mark the response.

QUESTION 4 (4 marks)

A 0.100 mol sample of solid ammonium chloride, NH_4Cl, is dissolved in 50.0 mL (50.0 g) of water. The temperature of the water drops by 7.2°C.

a) Calculate the quantity of heat absorbed by the solution, in kJ. [*2 marks*]

$Q = mc\Delta T$

$= 50.0 \text{ mL} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 7.2 \text{ K}$

$= 1{,}504.8 \text{ kJ}$

b) Calculate the molar enthalpy of solution. [*2 marks*]

Molar enthalpy of reaction $= Q \times n$

$= 1{,}504.8 \text{ kJ} \times 0.100 \text{ mol}$

$= 150.48 \text{ kJ}$

Marking guide

Question 4a	- correctly substitutes m, c and ΔT [1 mark] - correctly calculates Q and expresses in kJ; accept 1.5 kJ to 1.5048 [1 mark]
Question 4b	- correctly substitutes Q (from **4a**) and n [1 mark] - arrives at a consequentially correct value with units of kJ or kJ mol^{-1} [1 mark]

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved? Write your own response to the same question to receive full marks from a QCAA assessor.

Exam tip 3: Follow convention for writing chemical reaction equations

Use correct chemical conventions in all equations for reactions.

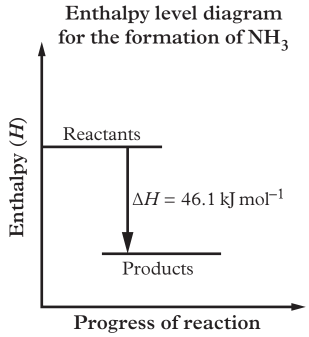
- Correct symbols for elements use upper case for the first letter and lower case for the second letter.
- Show correct state symbols (s), (l), (aq) and (g).
- Thermochemical equations should use the unit kJ, rather than kJ mol^{-1} as the energy is for the amount, in mol, shown in the equation.
- Use the correct signs for ΔH: positive (+) for endothermic reactions and negative (–) for exothermic reactions.

See it in action

Read the exam-style question below and see how the tip has made a difference between a response that has scored full marks and a response where marks have been lost.

QUESTION 5 (2 marks)

The enthalpy level diagram for the formation of 1 mol of gaseous ammonia, NH_3, from nitrogen gas and hydrogen gas is shown.



Write a balanced thermochemical equation for this reaction.

Complete response

Correctly balances the equation using fractional coefficients since $\frac{1}{2}$ mol of nitrogen molecules and $\frac{3}{2}$ mol of hydrogen produces 1 mol of ammonia [1 mark]

$$\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightarrow NH_3(g) \qquad \Delta H = -46.1\ kJ$$

Recognises that the reaction is exothermic and has used a negative sign for ΔH and correctly uses kJ (rather than kJ mol^{-1}) since the reaction is for 1 mol of NH_3 [1 mark]

Incomplete response

Appears to have correctly balanced the equation, but has provided the wrong formula for nitrogen (N instead of N_2); has also missed the state symbol for NH_3 [0 marks]

$N(g) + \frac{3}{2}H_2(g) \rightarrow NH_3$ $\Delta H = 46.1\ kJ\ mol^{-1}$

Has not recognised that the reaction is exothermic and has not included a negative sign for ΔH; incorrectly uses kJ mol^{-1} (rather than kJ) [0 marks]

Think like an assessor

To maximise your marks on an exam, it can help to think like a QCAA assessor. Consider how many marks each question is worth and what information the assessor is looking for.

A student has given the following response in a practice exam. Imagine you are a QCAA assessor and use the marking guide below to mark the response.

QUESTION 6 (4 marks)

When 0.10 mol of liquid ethanol, C_2H_5OH, is burnt completely in excess oxygen, 136.7 kJ of energy is released.

a) Calculate the molar enthalpy of the combustion reaction. [*2 marks*]

$\text{molar enthalpy of the reaction} = \frac{\text{energy}}{n} = \frac{136.7}{0.10} = 1{,}367\ kJ$

b) Write the thermochemical equation for the reaction. [*2 marks*]

$C_2H_5OH + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ $\Delta H = 1{,}367\ kJ$

Marking guide

Question 6a	• correctly substitutes *Q* and *n* [1 mark] • correctly calculates molar enthalpy; accept –1,367 kJ [1 mark]
Question 6b	• shows the balanced chemical equation with state symbols; accept $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$ [1 mark] • shows the molar enthalpy at the end of the thermochemical equation; $\Delta H = -1{,}367\ kJ$ [1 mark]

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved? Write your own response to the same question to receive full marks from a QCAA assessor.

Part C – Practice exam questions

Now it's time to put the tips and advice you've learned into practice while you complete these exam-style questions!

Multiple choice

(1 mark each)

1 A monatomic species has 17 protons, 20 neutrons and 18 electrons. The correct nuclear symbol for this species is

A ${}^{37}_{20}\text{Cl}^{+}$.

B ${}^{37}_{20}\text{Cl}^{-}$.

C ${}^{37}_{17}\text{Cl}^{-}$.

D ${}^{37}_{17}\text{Ar}^{-}$.

2 Consider the following substances and their properties in Table 1.

TABLE 1 Properties of substances

Colour	Melting point (°C)	Boiling point (°C)	Conducts electricity as a solid	Conducts electricity as a liquid
Silver	1,668	3,287	Yes	Yes
Black	114	184	No	No
Yellow	402	953	No	Yes
Colourless	1,713	2,950	No	No

Which substance has covalent molecular bonding?

A Black substance

B Silver substance

C Yellow substance

D Colourless substance

3 Which of the following is correct about atomic absorption spectroscopy (AAS)?

A AAS is both qualitative and quantitative.

B AAS measures the absorbance of a sample across a range of wavelengths to create a spectrum.

C AAS is a qualitative analytical technique that can be used to determine the identity of all of the elements in a sample.

D Concentrations of unknown samples are determined with reference to measurements using solutions of known concentration.

4 Which of the following would have the smallest first ionisation energy?

A Calcium, Ca

B Bromine, Br

C Krypton, Kr

D Potassium, K

5 Consider hydrocarbons with the formulas C_5H_{12} and C_5H_{10}. Which of the following is correct?

A C_5H_{12} is an alkane.

B Both C_5H_{12} and C_5H_{10} are alkanes.

C C_5H_{12} is unsaturated, whereas C_5H_{10} is saturated.

D All compounds with the formula C_5H_{10} are unsaturated.

6 Which of the following elements will have a condensed electron configuration that uses [Ne]?

A Potassium

B Carbon

C Calcium

D Phosphorus

7 Which one of the following compounds is an ionic compound?

A BN

B SiO_2

C BrCl

D K_2SO_4

8 Which of the following statements is always true for ionic compounds?

A Ionic compounds consist of a lattice of metal cations in a fixed ratio with anions.

B Ionic compounds have high melting points, and form liquids that conduct electricity.

C Ionic compounds dissolve well in water to produce aqueous solutions that conduct electricity.

D Ionic compounds are brittle and shatter with irregular breaks when struck with a hammer.

9 Which would you need to know to determine the identity of an isotope of a given element?

A The atomic number

B The number of protons

C The number of neutrons

D The relative atomic mass

10 Which of the following statements is true for endothermic reactions?

A The products have a higher enthalpy than the reactants and the surroundings feel colder.
B The products have a higher enthalpy than the reactants and the surroundings feel hotter.
C The products have a lower enthalpy than the reactants and the surroundings feel colder.
D The products have a lower enthalpy than the reactants and the surroundings feel hotter.

11 Which of the following groups lists only covalent molecular compounds?

A CO, CO_2, SiO_2
B Al_2O_3, P_4O_6, SO_3
C N_2O_4, NH_3, NH_4NO_3
D C_6H_6, H_2O_2, CH_3OH

12 Which of the following provides the electron configurations for Cr and Cu, respectively?

A $[Ar]4s^2 3d^4$, $[Ar]4s^2 3d^9$
B $[Ar]4s^2 4d^4$, $[Ar]4s^2 4d^9$
C $[Ar]4s^1 3d^5$, $[Ar]4s^1 3d^{10}$
D $[Ar]4s^1 4d^5$, $[Ar]4s^1 4d^{10}$

13 When a piece of metal is hit with a hammer, it changes shape but does not break. The best explanation for this is that

A metals are both malleable and ductile.
B metallic bonding is the strongest type of bonding.
C ions in the lattice slide past each other and metallic bonds reform.
D metals are made of a lattice of metal ions with a sea of delocalised electrons.

14 Separation techniques of mixtures rely on differences in physical properties.

Decantation is able to separate

A all solids from liquids, after the solid settles.
B solutes from solvents, after the solute settles.
C larger solid particles from liquids, after the solid settles.
D smaller solid particles from liquids, after the solid settles.

15 The electron configuration for the oxide ion is

A $[He]2s^2 2p^2$.
B $[He]2s^2 2p^4$.
C $[He]2s^2 2p^5$.
D $[He]2s^2 2p^6$.

16 The successive ionisation energies for an element are shown in Figure 1.

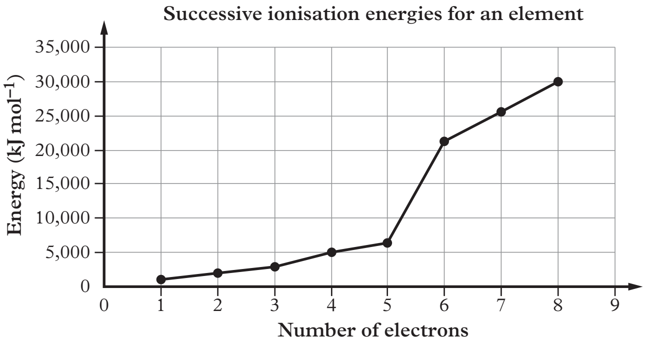


FIGURE 1 Successive ionisation energies for an element

The element could be

A sodium.
B chlorine.
C nitrogen.
D phosphorus.

17 The emission spectrum of an element

A is formed when electrons transition to higher energy levels.
B always appears in the visible range of the electromagnetic spectrum.
C has wavelengths corresponding to energy differences between electron orbitals.
D consists of dark lines superimposed on the brightly coloured electromagnetic spectrum.

18 The compound ethanoic acid is a component of vinegar.

Which of the following statements is true?

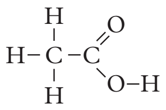


FIGURE 2 Ethanoic acid

A The formula unit is CH_2O and the empirical formula is $C_2H_4O_2$.
B The formula unit is CH_2O and the molecular formula is $C_2H_4O_2$.
C The empirical formula is CH_2O and the formula unit is $C_2H_4O_2$.
D The empirical formula is CH_2O and the molecular formula is $C_2H_4O_2$.

19 The formula for the carbonate ion is CO_3^{2-}. The polyatomic ion is made of

A one carbon atom bonded to three oxygen atoms, with two electrons added.

B three carbon atoms bonded to three oxygen atoms, with two electrons added.

C one carbon atom bonded to three oxygen atoms, with two electrons removed.

D three carbon atoms bonded to three oxygen atoms, with two electrons removed.

Short response

20 Calculate the amount, in mol, in 2.922 g of NaCl. (2 marks)

21 Discriminate between homogeneous mixtures and heterogeneous mixtures, using examples. (2 marks)

22 Determine the Lewis structures of the following:

a bromine, Br_2 (1 mark)

b methanol, CH_3OH (1 mark)

c ammonium ion, NH_4^+. (2 marks)

23 Potassium exists as three isotopes, with the following masses and abundances:

Potassium-39, 38.964, 93.258%

Potassium 40, 39.964, 0.012%

Potassium-41, 40.962, 6.730%

Determine the relative atomic mass of potassium. (2 marks)

24 Analyse Figure 3 to determine the periodic trend and the group trend in electronegativities. (4 marks)

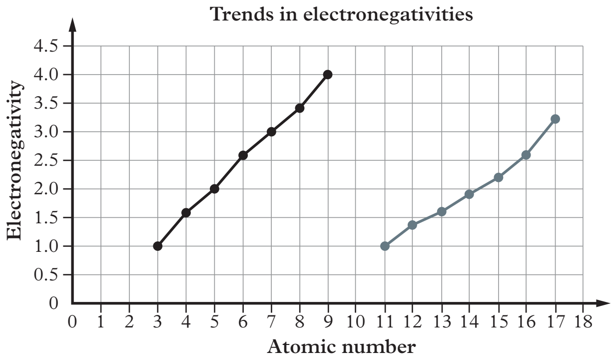


FIGURE 3 Trends in electronegatives

25 Determine the condensed electron configurations for

a Fe (1 mark)

b S^{2-}. (1 mark)

26 Determine the

a chemical formula for the following compounds.

i dinitrogen tetroxide (1 mark)

ii calcium phosphate (1 mark)

b IUPAC name for the following compounds.

i SO_3 (1 mark) **ii** Na_2CO_3 (1 mark)

27 Compare the structures of the following allotropes of carbon: graphite, diamond, graphene, and the fullerenes. (2 marks)

28 Identify the number of protons, neutrons and electrons in $^{56}_{26}Fe^{3+}$. (3 marks)

29 A compound consists of 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. The molar mass of the compound is 180.18 $g\,mol^{-1}$. Determine the molecular formula of the compound. (6 marks)

30 Explain the following, using a diagram where appropriate.

a Copper is a very good conductor of electricity as a solid, while copper sulfate does not conduct electricity as a solid but will conduct electricity when in aqueous solution. (5 marks)

b Aqueous solutions of oxygen do not conduct electricity, but aqueous solutions of carbon dioxide are able to conduct electricity to a small extent. (3 marks)

31 The electronegativities of the three elements are shown in Table 2.

TABLE 2 Electronegatives of three elements

C	H	O
2.6	2.2	3.4

Consider the first six saturated hydrocarbons (alkanes) and their counterpart primary alcohols. Primary alcohols contain an oxygen atom bonded to a carbon atom at the end of the carbon chain. A hydrogen atom is then bonded to this oxygen atom, forming an OH group.

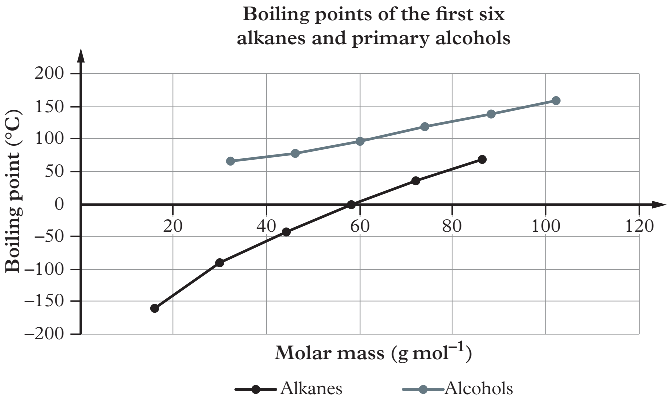


FIGURE 4 Boiling points of the first six alkanes and primary alcohols

a Analyse the data in Figure 4 to determine the relationship between boiling point and the size of the molecule. Consider anomalies. (3 marks)

b Alkanes contain C–C and C–H bonds, while alcohols contain C–C and C–H bonds, as well as one C–O bond and one O–H bond. Calculate the difference in electronegativity between the atoms in all four of these types of bonds. (2 marks)

c Use your answers from part **b** to infer the cause of differences in intermolecular forces of the alkanes and the alcohols. (3 marks)

32 A 0.25 g mass of hexane, C_6H_{14}, is combusted in a spirit burner to heat 100.0 g of water. The temperature increases by 20°C.

a Calculate the quantity of heat supplied to the water. (2 marks)

b Calculate the experimental value for the molar enthalpy of combustion of hexane. (4 marks)

c The accepted value for the molar enthalpy of combustion of hexane is 4,160 kJ mol^{-1}. Calculate the percentage error. (2 marks)

33 Table 3 shows the results of some physical tests on pure substances A–D. Not all tests were conducted on each substance. The identities of the substances, not necessarily in order, are silicon carbide, sodium chloride, lead and decanoic acid.

TABLE 3 Results of physical tests on four pure substances

Substance	Melting point (°C)	Electrical conductivity (solid)	Brittle or malleable
A	32	No	Not tested
B	2,730	No	Brittle, irregular fracture
C	327	Yes	Not tested
D	801	Not tested	Brittle, regular fracture

a Deduce the names of each of the four substances A to D. Justify your responses. (4 marks)

b Substance D was found to dissolve in water. Predict the electrical conductivity of substance D in solid form and when it is dissolved in water. (2 marks)

c Select a substance that is likely to be useful as an industrial abrasive, and explain, with reference to its structure, why it would be suitable. (3 marks)

34 A 2.00 g sample of potassium iodide and 2.00 g of lead(II) nitrate are each weighed out and dissolved in water. Then, the two solutions are mixed. They react to form a precipitate of lead(II) iodide.

a Determine a balanced equation for the reaction. (1 mark)

b Determine the limiting reagent. (5 marks)

c Calculate the mass of solid lead(II) iodide that forms. (2 marks)

35 A calibration curve is constructed to help determine the concentration of Fe in wastewater obtained from an abandoned mine (Figure 5). The concentration of Fe in the collected wastewater sample is found to be too high to use AAS, so 1.00 mL of sample is diluted to 100.0 mL. Two separate samples of the diluted wastewater are tested, and their absorbances are 0.170 and 0.320, respectively.

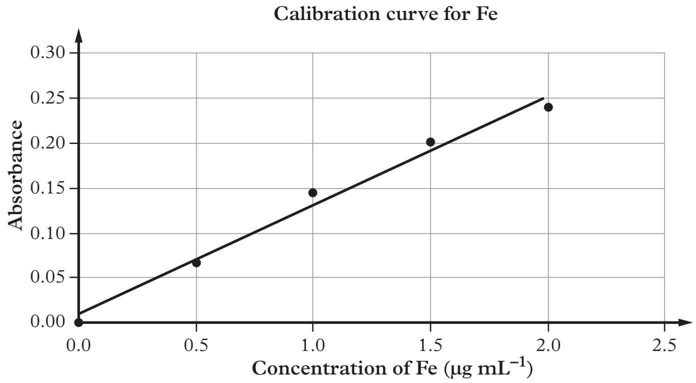


FIGURE 5 Calibration curve of Fe

a Determine the undiluted concentration of Fe in the first sample. (2 marks)

b Explain why it is not feasible to determine the concentration of Fe in the second sample. (1 mark)

c Describe a procedure that you could use to determine the concentration of Fe in the second sample. (2 marks)

TOTAL MARKS

/96

UNIT 2

Molecular interactions and reactions

Unit 2 overview

Over 70% of Earth's surface is covered in liquid water. The chemical reactions of life depend on the solution properties of water. The nature and strength of the intermolecular forces between the water molecules explain its melting point, boiling point and vapour pressure, as well as its ability to dissolve some substances, but not others. Most of the water on Earth is in the oceans, a mixture of water containing soluble ions from dissolved minerals, dissolved organic molecules, as well as dissolved gases such as oxygen and carbon dioxide.

Water on Earth is present as a solid, liquid and gas. Water as a vapour is subject to the same laws of temperature, volume and pressure as other atmospheric gases. The constantly moving molecules of gases display high levels of kinetic energy. Frequent collisions between high-energy gas molecules increase the rate of chemical reactions, such as in the atmosphere. Another way to increase the rate of chemical reactions in solution is through biological catalysts, such as enzymes in blood, which is around 51% water by mass.

Unit objectives

1. Describe ideas and findings about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
2. Apply understanding of intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
3. Analyse data about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
4. Interpret evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
5. Evaluate processes, claims and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.
6. Investigate phenomena associated with intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

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Unit 2 Topics

Topic	Module	
Topic 1 Intermolecular forces and gases	**Module 9**	Intermolecular forces
	Module 10	Chromatography techniques
	Module 11	Gases
Topic 2 Aqueous solutions and acidity	**Module 12**	Aqueous solutions and molarity
	Module 13	Solubility and identifying ions in solution
	Module 14	Acids and bases
	Module 15	Reactions of acids
Topic 3 Rates of chemical reactions	**Module 16**	Rates of chemical reactions

MODULE

9 Intermolecular forces

Introduction

Intermolecular forces are a key concept to understanding most areas of chemistry. No other concept is used as extensively as intermolecular forces, the interactions between different molecules and ions.

In organic chemistry, how electrons are shared between atoms in molecules (intramolecular covalent bonds) explains the characteristic reactions of certain chemicals and allows chemists to predict what product will form from two reactants. In analytical chemistry (and forensic analysis), intermolecular forces (i.e. the attractive forces that exist between different molecules and ions) are used to separate substances, based on their properties and attraction to other substances. Biochemists use intermolecular forces to explain the interaction of biomolecules within the body, their functions, their properties and their various health risks.

Inorganic chemists use intermolecular forces to explain how ionic compounds can dissolve in water and certain other organic solvents. Materials chemists use intermolecular forces to determine the properties of plastics, clothing and other various fabrics.

A strong grasp of this topic is essential for future studies in Units 1–4 Chemistry.

Prior knowledge

Prior knowledge quiz

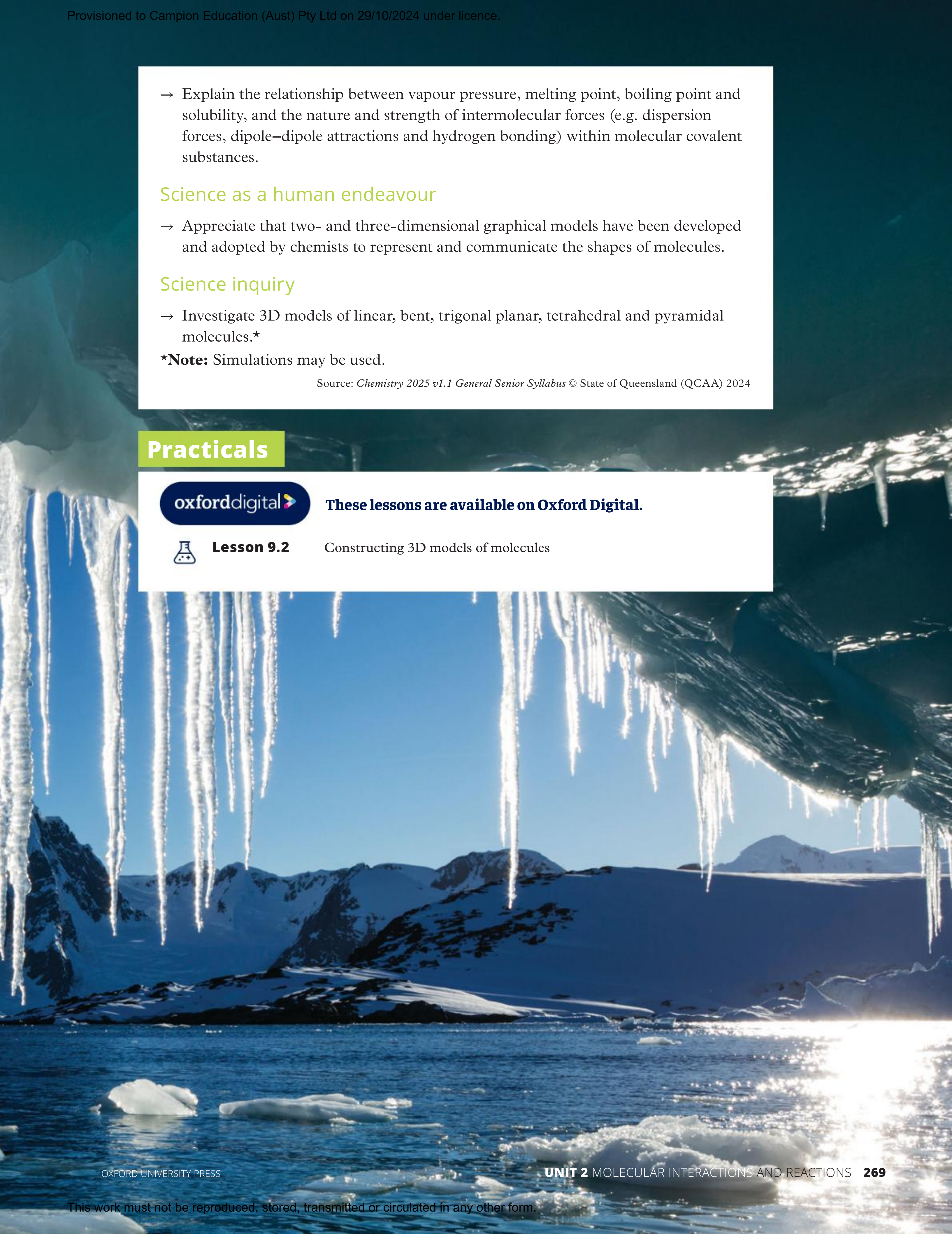
Check your understanding of bonding, Lewis structures and physical properties before you start.

Subject matter

Science understanding

→ Apply the valence shell electron pair repulsion (VSEPR) theory to determine the shape and bond angles of linear, bent, trigonal planar, tetrahedral and pyramidal molecules. (Hybridization involving d-orbitals are not required.)

→ Determine the polarity of molecules using molecular shape, understanding of symmetry, and comparison of the electronegativity of elements.



→ Explain the relationship between vapour pressure, melting point, boiling point and solubility, and the nature and strength of intermolecular forces (e.g. dispersion forces, dipole–dipole attractions and hydrogen bonding) within molecular covalent substances.

Science as a human endeavour

→ Appreciate that two- and three-dimensional graphical models have been developed and adopted by chemists to represent and communicate the shapes of molecules.

Science inquiry

→ Investigate 3D models of linear, bent, trigonal planar, tetrahedral and pyramidal molecules.*

***Note:** Simulations may be used.

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Practicals

oxforddigital

These lessons are available on Oxford Digital.

Lesson 9.2 Constructing 3D models of molecules

Lesson 9.1

Shapes of covalent molecules

Learning intentions and success criteria

Key ideas

- → The shape and bond angles of a covalent molecule is determined by the valence shell electron pair repulsion (VSEPR) theory.
- → Covalent compounds can have a linear, bent, trigonal planar, tetrahedral or pyramidal shape.

molecular shape the shape formed between a central atom and bonded atoms when covalent bonds form

valence shell electron pair repulsion (VSEPR) theory states that electron pairs in bonds and lone electron pairs around a central atom repel each other, forming a 3D shape

bond angle the angle between covalent bonds attached to the same central atom

What do you remember about covalent molecules?

Module 3 covered the covalent bonding that occurs between non-metallic elements. The group of molecules formed from these elements is biologically important because it includes DNA, proteins, fats, sugars and most medicines.

You learnt that covalent bonds occur between two non-metals when they share their electrons. They can also occur between a non-metal and a metalloid. Non-metals and metalloids are found on the periodic table in the top right-hand side (Figure 1) and include hydrogen. The noble gases are largely unreactive and only form molecules with covalent bonds with great difficulty. No compounds are known for helium and neon.

Recall that covalent molecules share valence electrons to fill their outer shell. These molecules can have bonding electrons, which link two atoms together, and non-bonding electrons (lone pairs). The arrangement of these bonding and non-bonding electrons leads to a geometric shape.

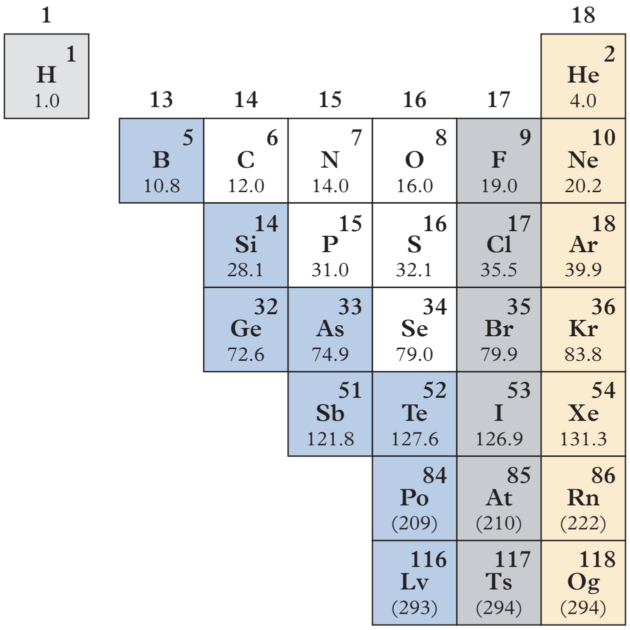


FIGURE 1 The non-metals (yellow, white and grey) and the metalloids (blue) on the periodic table form covalent bonds.

How does VSEPR theory affect molecular shape and bond angles?

Covalent molecules form three-dimensional (3D) shapes depending on the number of bonds, the type of bond (single, double or triple) and the number of non-bonding pairs of electrons. The **molecular shape** that is formed can be predicted by the **valence shell electron pair repulsion (VSEPR) theory**, which states that electron pairs, whether they are bonding or non-bonding electrons, repel each other.

To determine the shape of a simple, small molecule, the central atom must first be identified. All bonds attach to this atom, making it easily identifiable. Each pair of electrons exerts a repulsive force on the other pairs of electrons. These forces push the electron pairs away from each other, so they move as far apart as possible around the central atom.

Although lone pairs of electrons influence **bond angles**, they are not considered when describing the shape of the molecule. Only the position of bonded atoms are considered when describing the molecule's shape.

Lewis structures are important in determining the number of lone and bonding pairs of electrons so that VSEPR can be applied, but do not always depict the shape accurately. Ball-and-stick models are the best ways to visualise the shape of a molecule. Let's look at these with each molecule shape.

How are electrons arranged in linear molecules?

A **linear** molecule shape results when the central atom forms two covalent bonds and there are no non-bonding pairs of electrons on the central atom. The electrons in the covalent bonds push away from each other, as far apart from each other as possible. This results in a central bond angle of 180°, placing the terminal atoms on opposite sides of the central atom (Figure 2).

The carbon dioxide molecule is an example of a linear molecule (Figure 3). Carbon is the central atom and forms two double bonds with oxygen atoms with a bond angle of 180°. All molecules composed of two atoms are also linear.

linear
the molecular shape in which the central atom has two covalent bonds and no lone pairs of electrons, forming a straight line

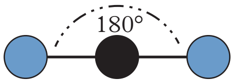


FIGURE 2 Ball-and-stick model of a linear covalent molecule; the black ball is the central atom and the bonds are at an angle of 180° to one another.

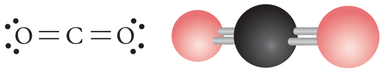


FIGURE 3 Lewis structure and ball-and-stick model of carbon dioxide, demonstrating a bond angle of 180° and a linear shape

How are electrons arranged in trigonal planar molecules?

A **trigonal planar** molecular shape results when the central atom forms three covalent bonds and there are no non-bonding pairs of electrons on the central atom. The Lewis structure looks like a T-shape (Figure 4); however, the bonding electrons push away from each other and spread as far apart as possible, resulting in a bond angle of 120° (Figure 5).

The methanal or formaldehyde molecule is an example of a trigonal planar molecule. Carbon is the central atom and forms a double bond with an oxygen atom and two single bonds with hydrogen atoms. The bond angles can be experimentally measured and are close to 120°.

trigonal planar
the molecular shape in which the central atom has three covalent bonds and no lone pairs of electrons; sometimes called triangular planar

tetrahedral
the molecular shape in which the central atom has four covalent bonds and no lone pairs of electrons, forming a tetrahedral structure

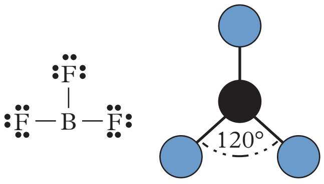


FIGURE 4 Lewis structure and ball-and-stick model of a trigonal planar covalent molecule; the black ball is the central atom and the bonds are at a bond angle of 120° to one another.

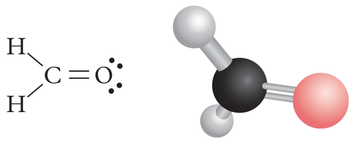


FIGURE 5 Lewis structure and ball-and-stick model of formaldehyde, demonstrating a bond angle of 120° and a trigonal planar shape

How are electron arranged in tetrahedral molecules?

A **tetrahedral** molecular shape results when the central atom forms four covalent bonds and there are no non-bonding pairs of electrons on the central atom. The electrons in the covalent bonds push away from each other so that they are as far away from each other as possible, resulting in a bond angle of about 109° (Figure 6).

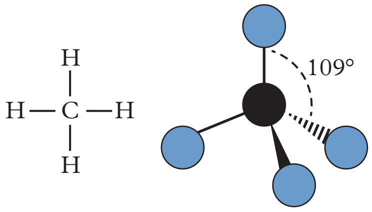


FIGURE 6 Lewis structure and ball-and-stick model of a tetrahedral covalent molecule; the black ball in the ball-and-stick model is the central atom and the bonds are at a bond angle of 109° to one another.

Study tip

Review your knowledge of Lewis structures. Use them to determine the bonding and non-bonding (lone) pairs of electrons before you determine the shape of the molecule.

As the bonds in a tetrahedral molecule are not in line with one another (or do not exist in a two-dimensional (2D) plane), a solid wedge and dashed wedge are used to represent bond direction. A dashed wedge represents a bond that is oriented away from the viewer (into the paper or screen) and a solid wedge represents a bond that is oriented towards the viewer (coming out of the paper or screen). This is shown in the methane molecule in Figure 7. Carbon is the central atom and forms four single bonds with hydrogen atoms. The bond angle is close to 109°.

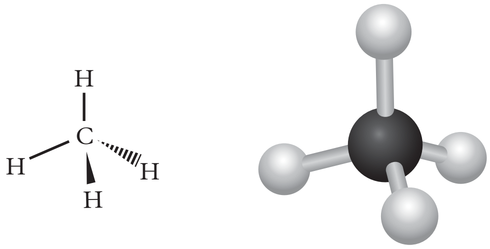


FIGURE 7 Lewis structure and ball-and-stick model of methane, demonstrating a bond angle close to 109° and a tetrahedral shape

How are electrons arranged in pyramidal molecules?

pyramidal
the molecular shape in which the central atom has three covalent bonds and one lone pair of electrons, forming a pyramid structure

A **pyramidal** molecular shape results when the central atom forms three covalent bonds, with one non-bonding pair of electrons. The lone pair and the bonding electrons push the bonds to create as much space between them as possible. However, the lone pair of electrons occupies slightly more space than a bonding pair of electrons. They push the bonds further together, reducing the angles between them. This means that the bonds are further away from the lone pair of electrons, but closer to each other, resulting in a bond angle of 107° (Figure 8), which is slightly less than the 109° bond angle in a tetrahedral shape.

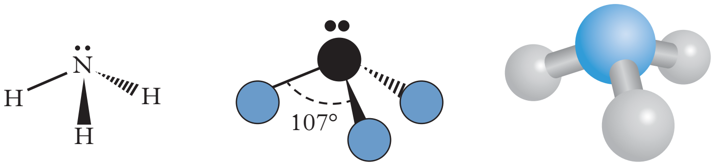


FIGURE 8 Lewis structure and ball-and-stick models of a pyramidal covalent molecule; the black ball in the ball-and-stick model is the central atom and the bonds are at a bond angle of 107° to one another, resulting in a pyramid shape.

The triangular solid and dashed wedges are also used in this model to represent the bonds that are not in line but are oriented towards or away from the viewer. The ammonia molecule is an other example of a pyramidal molecule. Nitrogen is the central atom and forms three single bonds with hydrogen atoms, with bond angle 107°.

How are electrons arranged in bent molecules?

bent
the molecular shape in which the central atom has two covalent bonds and two lone pairs of electrons, forming a V-shape

A **bent** molecular shape results when the central atom forms two covalent bonds and has two non-bonding pairs of electrons. The two non-bonding (lone) pairs of electrons push the bonds even closer, reducing the angle between the bonds. This results in a bond angle of 104.5°.

One example of this molecular shape is water (Figure 9). Oxygen is the central atom and forms two single bonds with hydrogen atoms and bond angle 104.5°.

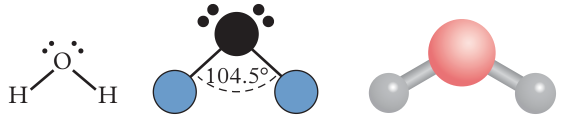


FIGURE 9 Lewis structure and ball-and-stick models of water, demonstrating a bond angle of 104.5° and a bent shape

Challenge

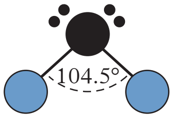
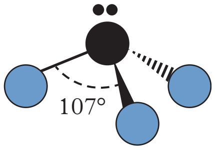
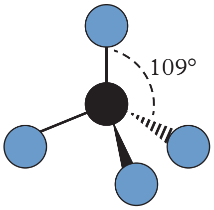
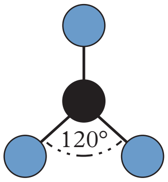
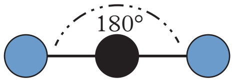
Bent molecules

Discuss whether all molecules that share the same shape have the same bond angle. **Propose** some factors that may change bond angles. (5 marks)

How do you predict molecular shape?

A summary of covalent molecular shapes is provided in Table 1. This will be useful to help you predict molecular shape for simple molecules.

TABLE 1 Bond shapes and angles of covalent molecules

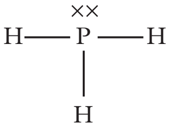
Name	Shape	Number of bonds	Number of non-bonded pairs of electrons	Bond angle	Example
Bent		2	2	104.5°	H_2O
Pyramidal		3	1	107°	NH_3
Tetrahedral		4	0	109°	CH_4
Trigonal planar		3	0	120°	CH_2O
Linear		2	0	180°	CO_2

Worked example 9.1A

Determining the shape of phosphine

Worked example 9.1A: Watch a video that shows how to solve this problem.

Apply the VSEPR model to **determine** the shape of a phosphine molecule (PH_3) and its bond angles. (2 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the question is asking you to do.	"Apply" means to use knowledge and understanding in response to a given situation or circumstance. "Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to apply our understanding of VSEPR theory to find the shape of the molecule. The question is worth 2 marks, so we must recall the theory and apply it to the molecule.
Step 2: Draw the Lewis structure for the molecule. If you need a refresher, review Worked example 3.2A. This is not a necessary step but will help you while you are learning.	
Step 3: Identify the number of bonds and number of non-bonding electron pairs.	There are three bonds and one pair of non-bonding electrons (one lone pair).
Step 4: Refer to Table 1 to determine the shape of the molecule.	Pyramidal (1 mark)
Step 5: Refer to Table 1 to determine the bond angles in this shape.	107° (1 mark)

Your turn

Apply the VSEPR model to **determine** the shape of a hydrogen sulfide molecule (H_2S) and its bond angles. (2 marks)

Challenge

Molecular shape of more complex molecules

Consider the structure of ethanoic acid (Figure 10). **Identify** how many atoms have more than one atom connected to them. **Determine** how many shapes it includes and what these shapes are. **Evaluate** whether it is fair to say that ethanoic acid has a single shape. (4 marks)

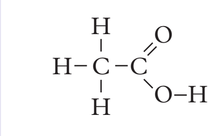


FIGURE 10 The structure of ethanoic acid

Skill drill

Experimentally determining the bond angles and shape of molecules with carbon as the central atom

Science inquiry skill(s): Communicating scientifically (Lesson 1.9)

A student sets out to determine the bond angles in a set of tetrahedral molecules. They perform an analysis and obtain the data shown in Table 2.

TABLE 2 Results of bond angle experiment

Molecule	Angle of bonds (hydrogen to carbon to hydrogen)	Shape
CH_4	109.47	tetrahedral
CH_3F	109.99	tetrahedral
CH_3Cl	110.45	tetrahedral

Practise your skills

1 **Identify** the independent and dependent variables in the experiment. (2 marks)
2 **Construct** a graph of the data. (2 marks)
3 **Describe** the trend in the data. (1 mark)
4 If all of the molecules are tetrahedral, **propose** a reason as to why their bond angles are not the same. (2 marks)

Real-world chemistry

Discovery of a double helix

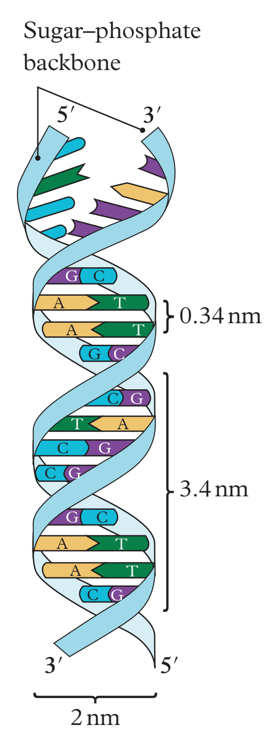


FIGURE 11 The structure of the DNA molecule, showing the pairing of the nitrogen-containing: bases guanine (G) with cytosine (C) and adenine (A) with thymine (T)

In 1953, the journal *Nature* published a ground-breaking paper by James Watson and Francis Crick, with the assistance of Maurice Wilkins, outlining the structure of the DNA molecule and its double helix geometry. The structure consisted of twin strands with pairs of nitrogen-containing units called bases linked by hydrogen bonds. These strands spiralled around each other and could be easily pulled apart for replication (Figure 11).

Biologist Watson and physicist Crick were working at the University of Cambridge in the UK. They already had key pieces of information about the structure and were in a race against other scientists to develop a model for the DNA molecule. They already knew that the structure contained four nitrogen-containing bases (guanine, cytosine, adenine and thymine) and that these bases paired up within the structure by hydrogen bonding.

Watson and Crick used X-ray crystallography, in which an X-ray beam is directed through a crystallised DNA molecule and the resultant image is captured on a photographic glass plate. This image provided information and data for the internal structure of the DNA molecule. But they could not find a geometric shape that matched what they already knew of DNA, and no image that they obtained from X-ray crystallography was clear enough to provide the missing information.

They discussed their lack of data with colleague Maurice Wilkins, who was working at Kings College in London. Wilkins provided them with an X-ray crystallography image taken by his colleague Rosalind Franklin (Figure 12), and this provided the last piece of the puzzle for them to complete their model. Watson and Crick used Franklin's data and image without her knowledge and consent.

▶

FIGURE 12 Rosalind Franklin made large contributions to the science of X-rays.

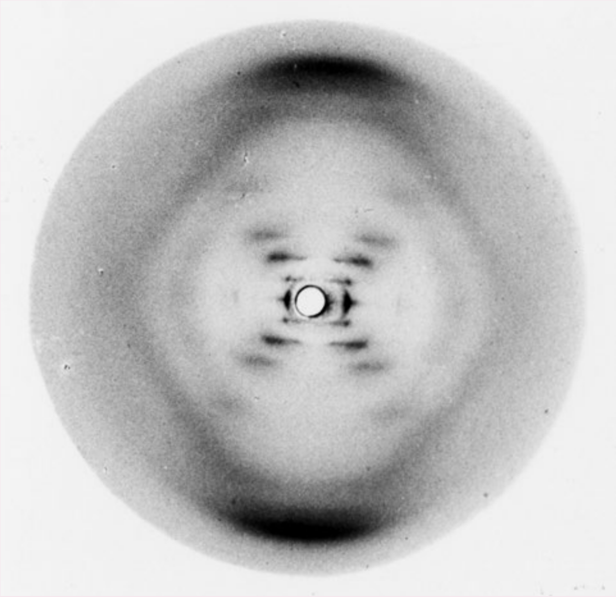


FIGURE 13 The X-ray image taken by Franklin

In 1958, Rosalind Franklin died of ovarian cancer, which she most likely developed as a result of her exposure to X-rays. In 1962, Watson, Crick and Wilkinson were awarded the Nobel Prize in Physiology or Medicine. The Nobel Prize rules state that the award cannot be received posthumously (after death), so Rosalind Franklin went largely unrecognised for her contribution during her lifetime. The work of these four scientists gave birth to molecular biology, combining the fields of chemistry and biology.

Apply your understanding

1 **Suggest** what the darker regions on the photographic glass taken by X-ray crystallography demonstrate. (1 mark)

2 **Propose** a reason why the DNA must be crystalised before conducting the X-ray. (1 mark)

Check your learning 9.1

Check your learning 9.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** the difference between tetrahedral, pyramidal and bent molecular shapes. (3 marks)

2 **Describe** the difference between single, double and triple bonds by explaining the behaviour of electrons. (3 marks)

Knowledge utilisation

3 **Determine** the shapes the following molecules would form. **Use** a Lewis structure and a ball-and-stick model to **justify** your answer.

- **a** Hydrogen iodide (HI) (3 marks)
- **b** Hydrogen cyanide (HCN) (3 marks)
- **c** Nitrogen triiodide (NI_3) (3 marks)
- **d** Borane (BH_3) (3 marks)
- **e** Dichlorine monoxide (OCl_2) (3 marks)
- **f** Fluoromethane (CH_3F) (3 marks)
- **g** Ethene (C_2H_4) (3 marks)
- **h** Ethyne (C_2H_2) (3 marks)
- **i** Hydrogen sulfide (H_2S). (3 marks)

Practical

Lesson 9.2

Constructing 3D models of molecules

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Learning intentions and success criteria

Lesson 9.3

Polarity and intermolecular forces

Key ideas

- → The polarity of a covalent bond depends on the electronegativity of the bonded atoms.
- → The polarity of a covalent molecule depends on its molecular shape and the sum of the polarity of its bonds.
- → Intermolecular forces between molecules include dispersion forces, dipole–dipole attractions and hydrogen bonding.

Learning intentions and success criteria

How does electronegativity affect bond polarity?

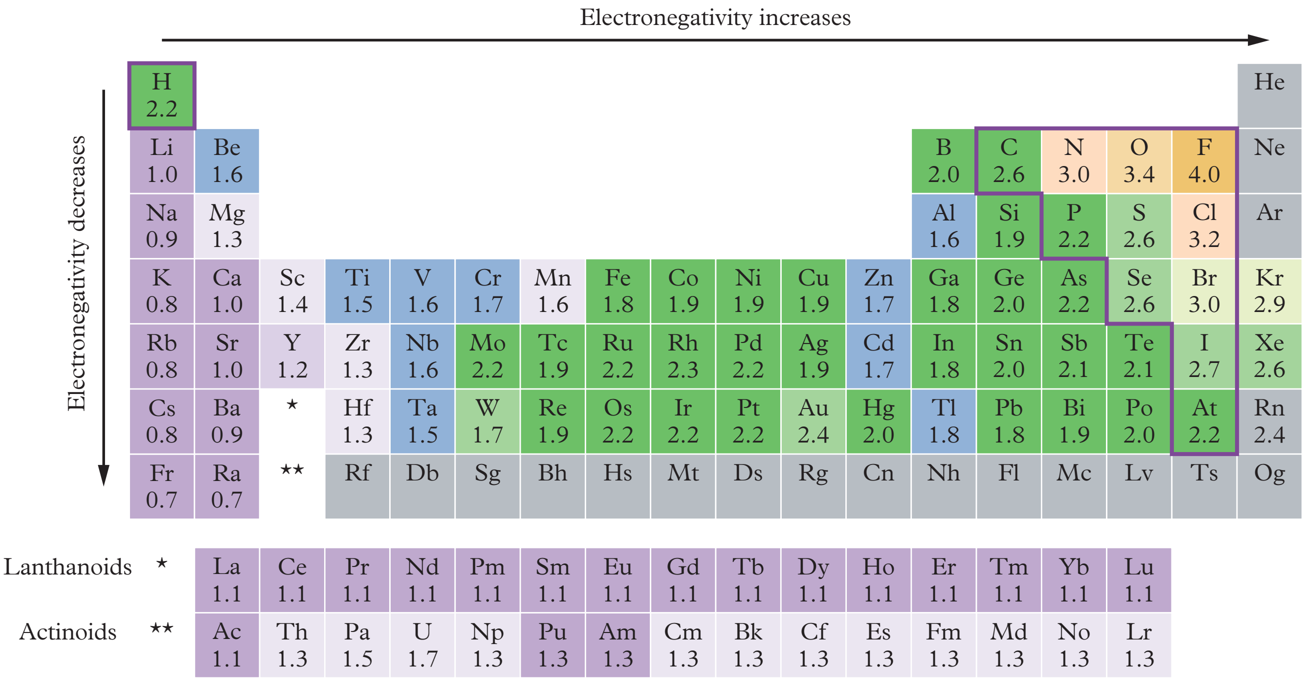
The shape of a molecule and how electrons are shared in covalent bonds are key to the understanding of the concept of **polarity**. Polarity helps to explain the properties of covalent compounds and their behaviour in different situations. Covalent bonds are classified as **polar** or non-polar depending on the electronegativity of the atoms involved in the bond.

The polarity of a bond depends on the electronegativity difference of the two bonded atoms. Remember that electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons within a covalent bond (Module 2). Electronegativity increases across the periodic table from group 1 to group 18 and decreases down the periodic table from period 1 to period 7. Fluorine is the most electronegative element of the periodic table with an electronegativity value of 4.0.

polarity
gives the molecule partially positive and partially negative ends; determined from the electronegativities of the atoms and the shape of a molecule

polar
having a dipole or permanent separation of charge; for example, a bond or molecule

FIGURE 1 Water is a polar molecule… but what does this mean?



H 2.2																	He
Li 1.0	Be 1.6											B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0	Kr 2.9
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	*	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 2.0	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

Lanthanoids	*	La 1.1	Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.1
Actinoids	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3

FIGURE 2 Electronegativity increases across periods and down groups on the periodic table.

The electronegativity difference between two bonded atoms is found by subtracting the two electronegativities. Table 1 provides a guide for how polar a covalent bond is based on this electronegativity difference. Larger differences are associated with ionic bonds.

TABLE 1 Electronegativity differences and the polarity of covalent bonds

Electronegativity difference	Polarity
0	Pure covalent
0.1 to 0.4	Non-polar covalent
0.5 to 0.9	Moderately polar covalent
1.0 to 1.8	Very polar covalent

Challenge

Noble gases

Explain why the noble gases helium, neon and argon do not have electronegativity values reported in Figure 2. (2 marks)

Non-polar bonds

When two chlorine atoms covalently bond to form a Cl_2 molecule, the atoms have the same electronegativity of 3.0, so the positively charged atomic nuclei pull on the electrons within the bond equally. The electrons do not move to either side of the molecule and there is no partial positive or negative charge assigned to the molecule (Figure 3). This is a non-polar bond.

FIGURE 3 The fluorine (F_2) molecule has no overall positive or negative charges.

Polar bonds

When fluorine (electronegativity of 4.0) is covalently bonded to hydrogen (electronegativity of 2.2), the fluorine atom pulls on the electrons within the bond (Figure 4). The electronegativity difference = 4.0 – 2.2 = 1.8.

As the electrons are attracted to the fluorine atom, they give the fluorine a partial negative charge. It is important to note that this is not a full charge (as in ionic bonding). Fluorine is slightly more negative, not because it has gained an electron but because the electrons are closer to the fluorine; the electron cloud is more dense near the fluorine atom. Similarly, as the electrons are pulled away from the hydrogen atom, hydrogen gains a partial positive charge. These partial positive and negative charges are represented by δ^+ (delta plus) or δ^- (delta minus) (Figure 5). The H–F bond is a polar bond.

As one side of the hydrogen fluoride molecule is more positive, and the opposite side is more negative, it creates a **dipole**. A permanent dipole occurs when the opposite sides of covalent bonds or molecules have opposite charges, so the molecule is said to be polar.

The hydrogen fluoride molecule contains a polar bond, whereas fluorine molecules do not.

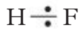


FIGURE 4 Electrons in the H–F bond are not in the centre of the bond but are closer to the fluorine.

$H^{\delta+} — F^{\delta-}$

FIGURE 5 Representation of partial positive and negative charges on a H–F bond: fluorine is partially negative (δ^-) and hydrogen is partially positive (δ^+).

dipole
the separation of opposite charges formed in a bond or molecule so that it has a partial positive charge at one end and a partial negative charge at the opposite end

How does molecular shape affect molecular polarity?

The shape of a molecule and the bond polarities influence its overall polarity.

Non-polar molecules

Consider the tetrahedral molecule carbon tetrachloride, which has carbon as the central atom bonded to four chlorine atoms (Figure 6). Carbon has an electronegativity value of 2.6 and chlorine has a value of 3.2, giving an electronegativity difference of 0.6; therefore, the individual C–Cl bonds in the molecules are slightly polar. However, the carbon tetrachloride molecule is symmetrical, so the partial charges created by the polar bonds cancel each other out and the overall effect is that the carbon tetrachloride molecule is non-polar. The spread of electrons is even – that is, electron density is evenly distributed.

Polar molecules

The ammonia molecule has a pyramidal shape (Figure 7). The central nitrogen atom has a larger electronegativity value than hydrogen. The pull of electrons (blue arrows) occurs in one direction in the molecule. This makes the hydrogen end of the molecule electropositive and the nitrogen end of the molecule electronegative, resulting in a polar molecule. There is a greater electron density towards the nitrogen end than the hydrogen end.

Similarly, if there is a difference in electronegativity between atoms in a linear molecule, it can be polar. Tetrahedral molecules can also be polar if one of the atoms attached to the central atom is more electronegative than another.

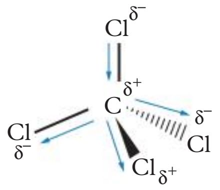


FIGURE 6 Tetrahedral structure of the carbon tetrachloride molecule. The blue arrows show the direction that the electrons are being pulled within the molecule.

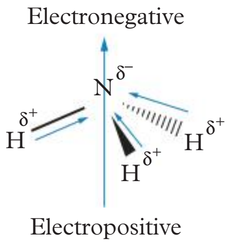


FIGURE 7 The pyramidal structure of the ammonia molecule results in it being polar.

Worked example 9.3A

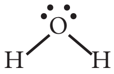
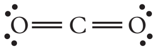
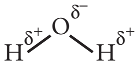
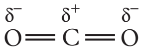
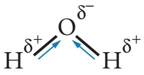
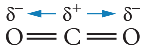
Determining the polarity of covalent molecules

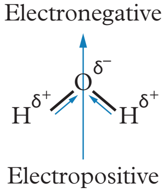
Worked example 9.3A: Watch a video that shows how to solve this problem.

For each of the following covalent molecules, **construct** Lewis structures, indicating which atoms are electropositive and electronegative, the direction of electron pull and the overall effect. Then, **determine** their polarity.

a water (H_2O) (5 marks)

b carbon dioxide (CO_2) (5 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Construct" means to display information in a diagrammatical or logical form. "Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to construct Lewis structure and use the bond polarities, together with the molecular shape, to make a decision about the overall polarity of the molecules. The questions are worth 3 marks, so we must analyse the molecules, represent them correctly and correctly judge their polarity.
Step 2: Draw the Lewis structures for the molecules. If you need a refresher, review Worked example 3.2A.	**a**  (1 mark) **b**  (1 mark)
Step 3: Determine the shape of the molecule. If you need a refresher, review Worked example 9.1A.	**a** Water has a bent shape. (1 mark) **b** Carbon dioxide has a linear shape. (1 mark)
Step 4: Identify the most electronegative elements and label them with $\delta-$. Label the atom attached to this element with $\delta+$. This represents a polar bond.	**a** Oxygen has an electronegativity value of 3.4 and hydrogen has an electronegativity value of 2.2, giving a difference of 3.4 – 2.2 = 1.2. Therefore, oxygen is electronegative ($\delta-$) and hydrogen is electropositive ($\delta+$).  **b** Oxygen has an electronegativity value of 3.4 and carbon has an electronegativity value of 2.6, giving a difference of 3.4 – 2.6 = 0.8. Therefore, oxygen is electronegative ($\delta-$) and carbon is electropositive ($\delta+$).  $\overset{\delta^-}{O}=\overset{\delta^+}{C}=\overset{\delta^-}{O}$
Step 5: Label the pull of electrons toward the electronegative element.	**a** As the oxygen atom is electronegative, it pulls the electrons within the bonds.  (1 mark) **b** As the oxygen atom is electronegative, it pulls the electrons within the bonds.  (1 mark)

Think	Do
Step 6: Based on the shape of the molecule, determine whether one side of the molecule has a higher electron density and is therefore more electronegative than another side. Use an arrow to represent the overall dipole of the molecule if one exists.	**a** Because oxygen pulls the electrons in the same direction, the net effect is that the oxygen end of the molecule is more negative and the hydrogen end is more positive, creating a polar molecule.  (1 mark) **b** Because oxygen pulls the electrons in opposite directions, the net effect is that the electron pull cancels out and the molecule has no permanent dipole. (1 mark)
Step 7: Determine whether the molecule is polar or non-polar.	**a** Water is polar. (1 mark) **b** Carbon dioxide is non-polar. (1 mark)

Your turn

For each of the following covalent molecules, **construct** Lewis structures, indicating which atoms are electropositive and electronegative, the direction of electron pull and the overall effect. Then, **determine** their polarity.

a chloromethane (CH_3Cl) (5 marks)

b carbon tetrachloride (CCl_4) (5 marks)

What are intermolecular forces?

The attraction between molecules is called an intermolecular force. These attractions are not classified as bonds, even though some are called bonding. There are three main types of intermolecular forces: dispersion forces, dipole–dipole attractions, and a type of dipole–dipole attraction called hydrogen bonding.

What are dispersion forces?

Dispersion forces occur between all covalent molecules, including polar and non-polar molecules. They contain valence electrons that constantly move throughout the molecule. At any given time, there may be more electrons on one side of the molecule than on the other, randomly creating a dipole that only exists in this moment, and is not permanent. This "**dipole moment**" is temporary and so any electrostatic interaction that it creates is fleeting, disappearing until the next dipole moment occurs. You can see how this occurs between molecules of hydrogen gas H_2 in Figure 8.

Therefore, dispersion forces are weak because they only occur for a short moment. They require another molecule to also be experiencing a dipole moment before they are attracted to one another. Of the three types of intermolecular forces you will study, dispersion forces are the only ones that non-polar molecules will experience.

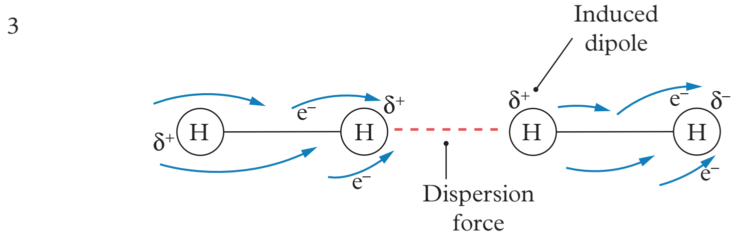
Dispersion forces also occur between polar molecules but are weaker than the other two types of intermolecular forces. They increase as the number of electrons in a molecule increases, and as the surface area available to interact with neighbouring molecules increases.

dispersion force
the weak interaction between all molecules that occurs due to the formation of temporary dipoles or dipole moments

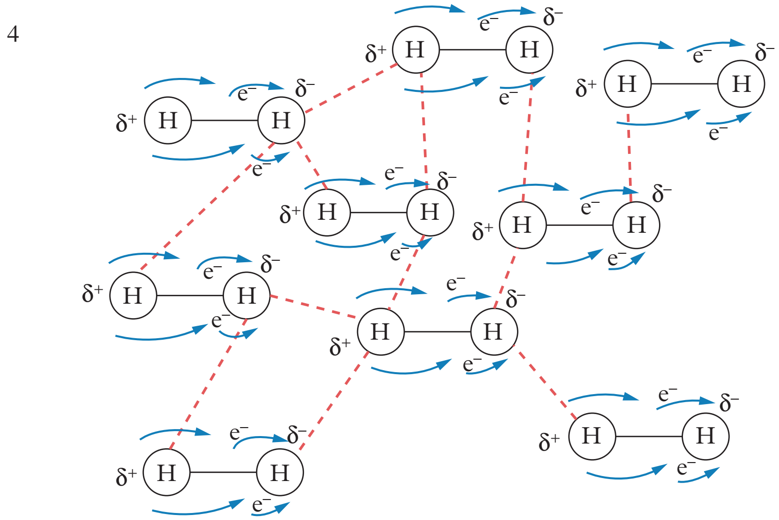
dipole moment
when a bond gains a temporary negative charge at one end and a temporary positive charge at the other end; also known as a temporary dipole

The shared electrons in H_2 move randomly around the bond. Overall, they spend the same amount of time at each of the hydrogen atoms.

Electrons can gather closer at one end to create a temporary negative charge at one end and temporary positive charge at the opposite end. This is called a temporary dipole (dipole moment).



When a H_2 molecule with a dipole moment gets close to another H_2 molecule, it creates a dipole in the second H_2. Electrons in the first H_2 molecule repel the electrons in the second H_2 and cause an induced dipole.



As a consequence, H_2 molecules with dipoles induce dipoles in many other nearby H_2 molecules. The attraction between the temporary dipoles is called a dispersion force.

FIGURE 8 Dispersion forces form between the temporary dipoles in H_2 molecules.

Study tip

Remember that dispersion forces are present between all molecules and should be listed with any other intermolecular forces acting.

What are dipole–dipole attractions?

Dipole–dipole attractions occur in molecules that have a permanent dipole, in other words, polar molecules. The positive end of the molecule is attracted to the negative end of an adjacent molecule. Repulsion also occurs when the negative ends or positive ends of molecules align.

Figure 9 represents the attractions between adjacent hydrogen chloride HCl molecules. The molecules are polar, so they have a permanent dipole. Attraction occurs between the partially negative chlorine of one molecule and the partially positive hydrogen of another molecule. As the dipole within these molecules is permanent, the attraction is stronger than the attraction created by dispersion forces.

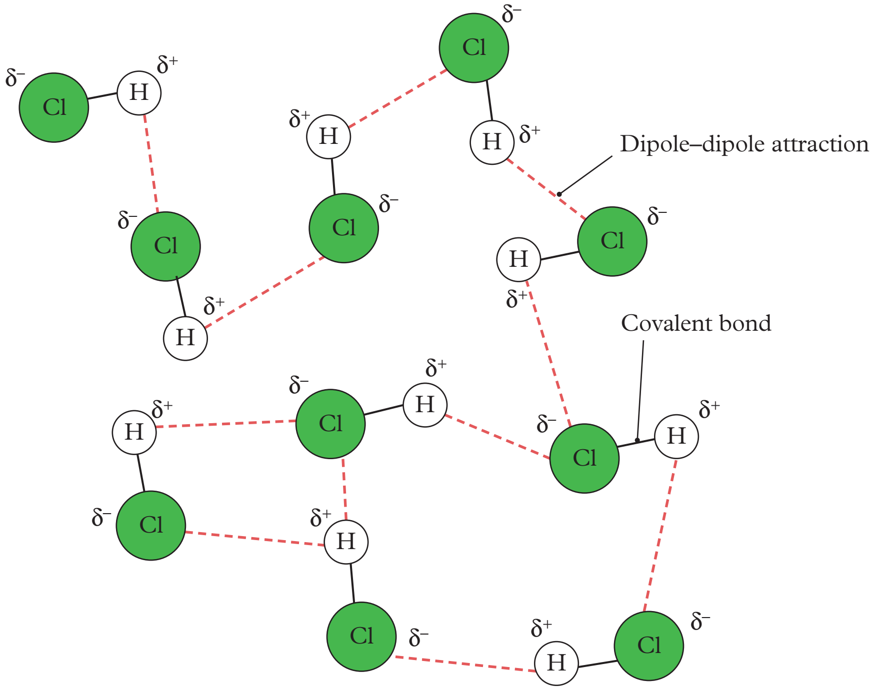


FIGURE 9 The dipole–dipole attractions between neighbouring HCl molecules

Study tip

Electrostatic attractions are always between a positive and negative. This can be a full positive and negative charge, as with ionic substances, or a partial positive and negative charge, as with polar covalent molecules.

Polar molecules can align in various arrangements when they are attracted to one another. This means that they will pack quite closely together when they interact.

Depending on the electronegativities, some polar bonds have a larger dipole than others and result in stronger dipole–dipole attractions between molecules. For example, the difference in electronegativities between atoms in the C–H atom is 0.4, compared to 0.9 for the Cl–H bond. Therefore, the chlorine–hydrogen bond forms a stronger dipole and results in stronger interactions between molecules. It is not necessary to calculate these values if you remember the trend in the periodic table.

What is hydrogen bonding?

Hydrogen bonding is a specific type of dipole–dipole attraction, where a partially positive hydrogen atom bonded to fluorine, oxygen or nitrogen on one molecule is attracted to the fluorine, oxygen or nitrogen atoms in another molecule. F, O and N are the three most electronegative elements and all have at least one lone pair of electrons that the partially positive hydrogen atom is attracted to.

This interaction is stronger than a regular dipole–dipole attraction because it involves the most electronegative (F, O and N) and electropositive (H) atoms on the periodic table. The small size of the partially positive hydrogen atom also contributes to the strength of hydrogen bonding. It can be positioned very close to the lone pair of electrons on the F, O or N.

For example, hydrogen bonds form between molecules of HF, H_2O and NH_3 (Figure 10). In water, the partially negative oxygen on one molecule attracts the partially positive hydrogen on a neighbouring molecule, resulting in a hydrogen bond. Can you identify the partially negative and partially positive atoms in hydrogen fluoride and ammonia?

hydrogen bonding
the strongest form of dipole–dipole attraction in which a partially negative atom (F, O or N) has an electrostatic attraction to a partially positive hydrogen atom that is bonded to an F, O or N in another molecule

Study tip

Although the name has “bonding” in it, hydrogen bonding is not a chemical bond. It is much weaker than metallic, ionic and covalent bonds.

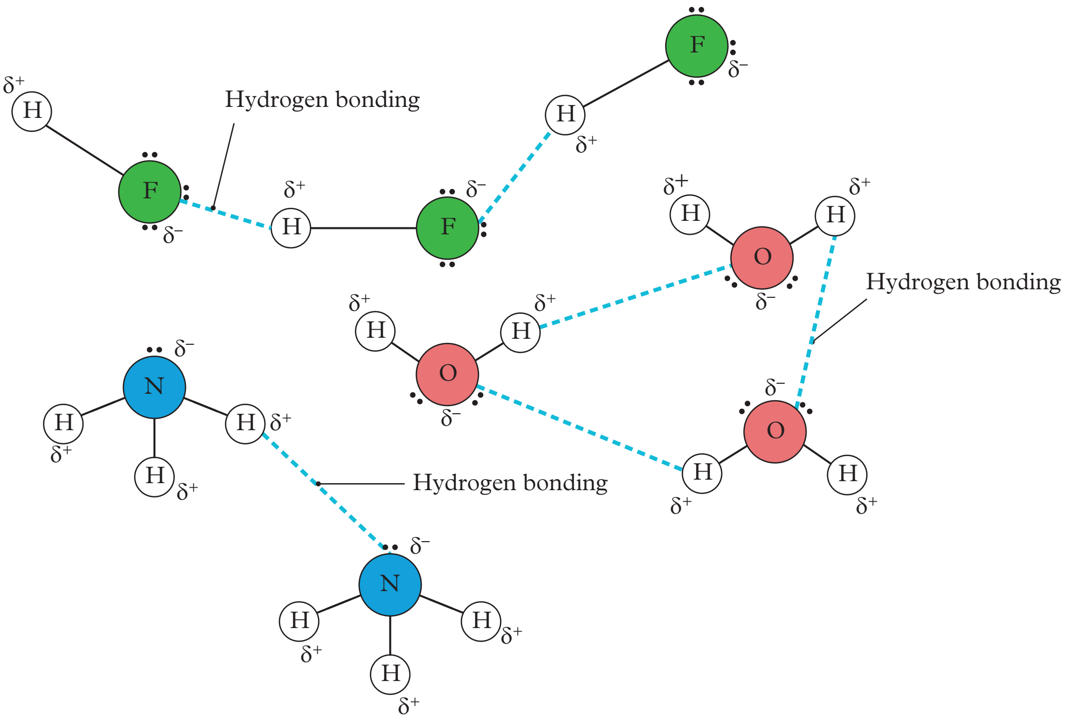


FIGURE 10 Hydrogen bonding exists between the partially positive atoms of one molecule and the partially negative atoms of neighbouring molecule.

Skill drill

Analysing and evaluating polarity data

Science inquiry skill(s): Processing and analysing data (Lesson 1.7) and Evaluating evidence (Lesson 1.8)

The polarity of a solvent is indicated by its polarity index. For example, pentane is a non-polar substance and has a polarity index of 0.0; water is a polar substance and has a polarity index of 10.0.

The polarity index of three solvents (methanol, acetonitrile and toluene) are determined experimentally by measuring their resistance when voltage is applied. Three replicate measurements are taken. The results are shown in Table 2 along with the theoretical polarity indexes.

TABLE 2 Experimental and theoretical polarity indexes of three solvents

Solvent	Experimental polarity index			Theoretical polarity index
Methanol (CH_4O)	5.1	5.3	4.9	5.1
Acetonitrile (C_2H_3N)	5.4	5.2	6.5	5.8
Toluene (C_7H_8)	3.7	3.8	3.8	2.4

Practise your skills

1 **Calculate** the mean experimental polarity index for each solvent. (1 mark)

2 **Construct** a graph to represent the means you calculated in question **1**. (2 marks)

3 **Evaluate** the results, by discussing the accuracy and precision of the measurements. (3 marks)

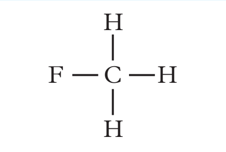
Worked example 9.3B

Determining the intermolecular forces present

Worked example 9.3B: Watch a video that shows how to solve this problem.

Analyse each of the following molecules and **determine** the strongest intermolecular forces that they would experience.

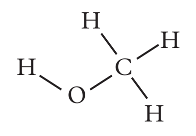
a



(1 mark)

FIGURE 11 Structure of CH_3F

b



(1 mark)

FIGURE 12 Structure of CH_3OH

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Analyse" means to examine or consider something in order to explain and interpret it. "Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to determine their polarity of each molecule and look at the atoms they consist of. The questions are worth 1 mark each, so we analyse each molecule and correctly judge the interactions they would experience.
Step 2: Determine whether the molecule is polar or non-polar. If you need a refresher, review Worked example 9.2A.	**a** CH_3F is polar. **b** CH_3OH is polar.
Step 3: If it is non-polar, there are only dispersion forces present. If it is polar, look at the atoms in each molecule to determine whether they have an F, O or N atom with a H attached.	**a** Has an F but there is no H attached. **b** Has an O with an H attached. Also has a C attached to an O.
Step 4: Consider the intermolecular forces occurring between the molecules of each compound.	**a** Dispersion forces, dipole–dipole attractions **b** Dispersion forces, dipole–dipole attractions, hydrogen bonding
Step 5: Identify the strongest intermolecular force. If it fulfils the criteria set out in Step 3, the molecule can form hydrogen bonds. If not, it can form dipole–dipole attractions.	**a** Dipole–dipole attractions (1 mark) **b** Hydrogen bonding (1 mark)

Your turn

Analyse the following molecule and **determine** the strongest intermolecular force that it would experience. (1 mark)

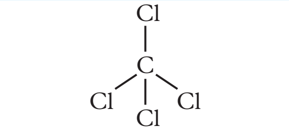


FIGURE 13 Structure of CCl_4

Real-world chemistry

Hydrogen bonding in DNA

DNA molecules hold the genetic code that builds all proteins required for the normal functioning of the human body. DNA molecules are shaped like a ladder. The outer bars or handrails of the ladder (light pink in Figure 14A) are made up of alternating sugar and phosphate groups. The sugar groups (dark pink in

Figure 14B) connect to the rungs (or steps) of the ladder, which are nitrogen-containing bases. These nitrogen-containing bases are called guanine, cytosine, adenine and thymine (GCAT). To connect both sides of the ladder, the bases form hydrogen bonds. This hydrogen bonding is between a partially negative nitrogen or oxygen and a partially positive hydrogen.

Guanine and cytosine form three hydrogen bonds, whereas adenine and thymine form two hydrogen bonds. Three hydrogen bonds hold the two bases closer together than two do. Therefore, the cytosine–guanine base pair is both stronger and held closer together. The ladder twists because of the length in the hydrogen bonds.

The body makes a copy of DNA by cutting the ladder down the middle, by breaking the hydrogen bonds, and making a copy of both sides. Therefore, it is important that the forces that hold the nitrogenous bases together are strong, so that the DNA does not break down within the body. However, they cannot be so strong that they cannot be pulled apart when a copy of the DNA is required to make a protein. Hydrogen bonding provides the perfect amount of strength to hold the molecule together. All other bonds within the DNA molecule are covalent.

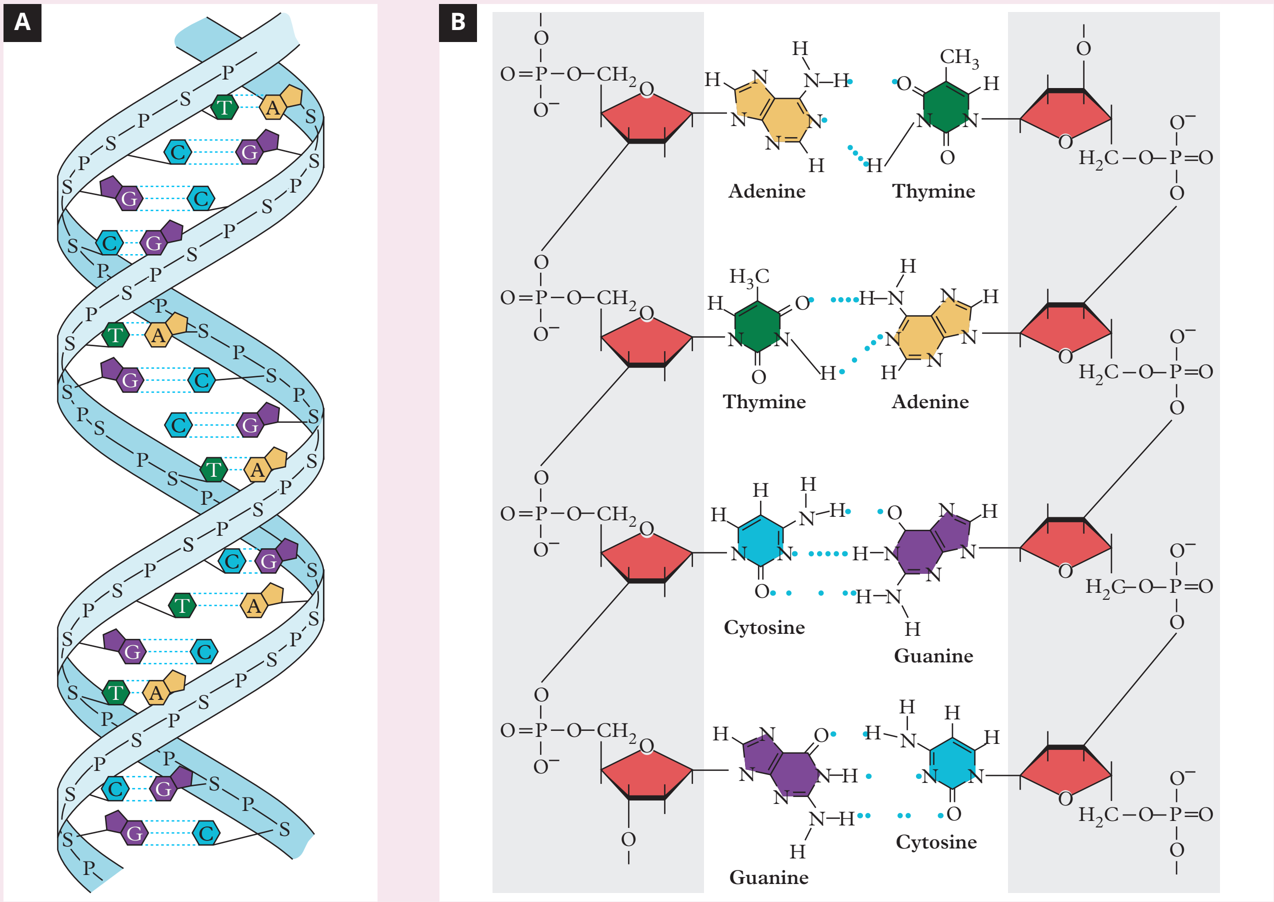


FIGURE 14 (A) The double helix structure of DNA; (B) The structure of DNA, demonstrating hydrogen bonding between nitrogen-containing bases

Apply your understanding

1 **Explain** why it is essential for hydrogen bonds between nitrogenous bases to be weaker than the covalent bonds within the molecule. (2 marks)

2 DNA is often described using the analogy of a ladder. **Explain** why this analogy may be misleading, with reference to intermolecular forces. (2 marks)

3 **Develop** a more accurate analogy for DNA that would be appropriate for a primary school student. (3 marks)

Check your learning 9.3

Check your learning 9.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** polarity and **explain** why a polar molecule has a permanent dipole. (3 marks)

Apply, analyse and interpret

2 **Compare** hydrogen bonds and dipole–dipole attractions. (3 marks)

3 **Determine** whether the following molecules interact by dispersion forces, dipole–dipole attractions and/or hydrogen bonding.
  - **a** Butane (C_4H_{10}) (1 mark)
  - **b** Ethyne (C_2H_2) (1 mark)
  - **c** Ethanol (CH_3CH_2OH) (1 mark)
  - **d** Butanoic acid (C_3H_7COOH) (1 mark)
  - **e** Ammonia (NH_3) (1 mark)
  - **f** Methyl chloride (CH_3Cl). (1 mark)

4 **Determine** which intermolecular forces could exist between different pairs of molecules from question **3**. (15 marks)

Knowledge utilisation

5 **Consider** the structure of DNA in Figure 14. **Justify** the various types of bonding and intermolecular forces involved. (3 marks)

6 A student claims that the name “hydrogen bonding” is misleading and should be changed. **Evaluate** this statement. (2 marks)

Lesson 9.4

Intermolecular forces and physical properties

Key ideas

→ When intermolecular forces between the molecules of a substance are significant, their melting point and boiling points are high.

→ When intermolecular forces between the molecules of a substance are strong, their vapour pressure is lower.

→ When two different substances have similar polarities, they are likely to mix and dissolve in each other.

What physical properties are related to intermolecular forces?

Learning intentions and success criteria

The physical properties of covalently bonded molecules depend on the combination of the intermolecular forces between molecules. An increase in the size of a molecule results in increased dispersion forces (for example), which explains why HBr has a higher boiling point than HCl, even though the dipole–dipole attractions are greater for HCl.

The properties that depend on the strength and type of intermolecular forces present include:

- melting point
- boiling point
- vapour pressure
- solubility.

Study tip

You will learn more about the concept of solubility in Module 13.

How do intermolecular forces affect melting and boiling points?

All particles have **kinetic energy**. Kinetic energy is a measure of the motion of particles. The greater the kinetic energy of a particle, the faster it moves.

kinetic energy the energy involved with the motion of particles; can be increased by adding heat energy

When investigating the melting and boiling points of a substance, two factors must be considered:

- kinetic energy of the particles: the more kinetic energy a particle has, the more it moves around and breaks away from other particles
- energy involved with the intermolecular forces holding the molecules together: the greater the intermolecular forces holding the particles together, the more energy will be required to break them apart.

Temperature is a measure of the average kinetic energy of a system. To represent the energy of a collection of particles or molecules, the **Maxwell–Boltzmann distribution** can be used. This demonstrates the kinetic energies of all particles in a system and forms a skewed distribution (Figure 1).

temperature a measure of the average heat energy of the particles within a system

Maxwell–Boltzmann distribution a representation of the spread of kinetic energies of particles in a system

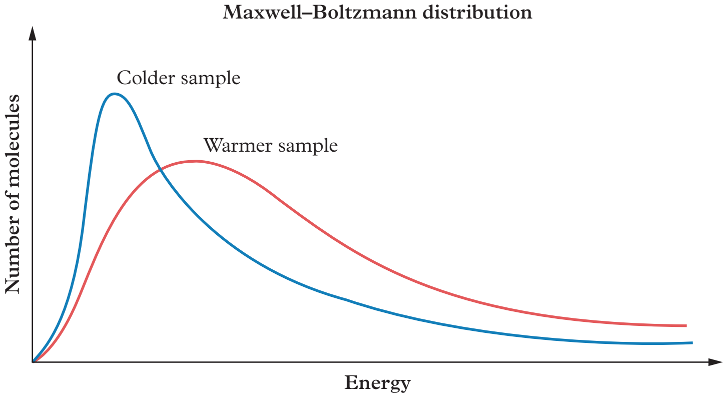


FIGURE 1 The Maxwell–Boltzmann distribution of particles energies within a system

According to the distribution, there will be a small number of particles that, randomly, have a high amount of energy. At higher temperatures, the distribution shifts to the right and a greater portion of particles have a higher energy.

Melting point

The melting point of a substance is the temperature at which a solid turns into a liquid. In a solid state, the intermolecular forces between molecules are so great that movement of the molecules cannot break them apart. The molecules are held together in a rigid structure, and their motion is so small that they only vibrate.

When heat energy is applied, molecules begin to move faster, and soon have enough energy to start moving away from neighbouring molecules. The molecules are still attracted to one another but are no longer held in fixed positions; they begin to move around. The temperature at which the heat energy is sufficient to break molecules free of their rigid structure, or turn a solid into a liquid, is called the melting point.

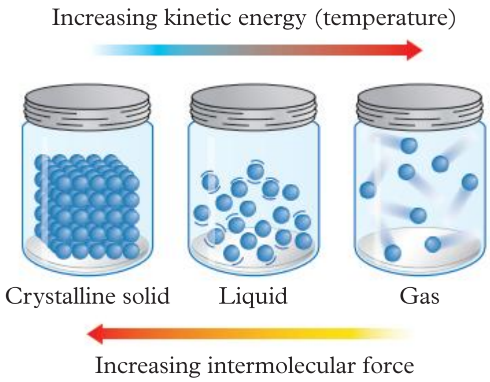


FIGURE 2 The relationship between kinetic energy and intermolecular forces in solid, liquid and gaseous states

Boiling point

The boiling point of a substance is the temperature at which a liquid turns into a gas. In a liquid state, the intermolecular forces between molecules are great enough that molecules are held together and cannot fully break apart from each other. Instead, they move over and around each other. Their kinetic energy is not great enough for them to break free of each other.

When heat energy is applied, molecules begin to move faster, and soon have enough kinetic energy to start breaking free from neighbouring molecules. The molecules are no longer able to interact with each other because their kinetic energies are greater than the energy involved with their intermolecular forces. The boiling point is the temperature at which the heat energy is sufficient to break molecules free, or turn a liquid into a gas.

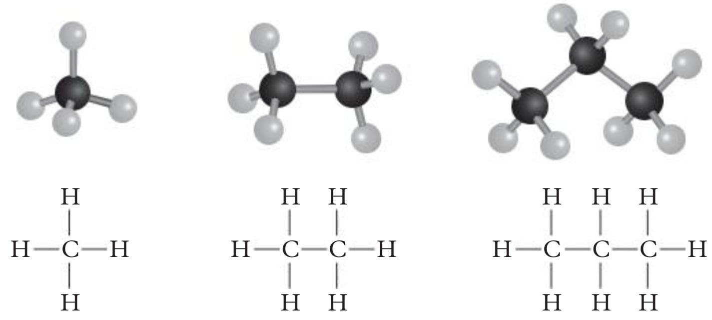
Trends in melting and boiling points

When the intermolecular bonding between molecules is stronger (dipole–dipole attractions and hydrogen bonding), more kinetic energy, and therefore heat, is required to break them apart into separate molecules. If molecules only interact through dispersion forces, which are very weak, less kinetic energy, and therefore heat, is required to break them apart compared to polar molecules of the same size and mass. The melting points of covalent molecules display similar trends as boiling points based on the strength of the intermolecular forces between molecules.

Non-polar molecules

We can observe this trend by considering non-polar molecules such as alkanes. Alkanes are a family of organic compounds that consist only of the elements carbon and hydrogen.

The carbons are covalently bonded in a long chain of single bonds, and the remaining bonds are to hydrogen. Some common alkanes are shown in Figure 3.



Methane (CH_4) **Ethane (C_2H_6)** **Propane (C_3H_8)**

FIGURE 3 Structures of the alkanes methane, ethane and propane

Alkanes experience only dispersion forces because there is no overall polarity throughout the molecules. All carbons are bonded and there are no lone pairs of electrons. Figure 4 shows the relationship between the number of carbons in an alkane and its boiling point.

As the number of carbon atoms increases, the boiling point of the molecule also increases. This is because increasing the number of carbon atoms means an increase in dispersion forces, which means more heat energy is needed to separate the weakly interacting molecules.

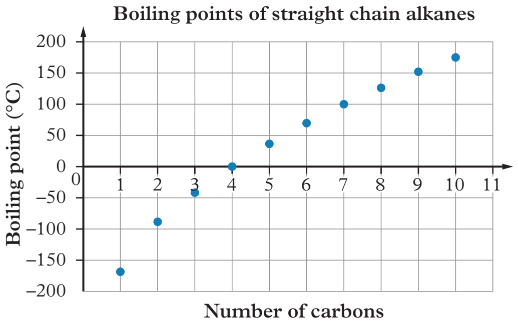
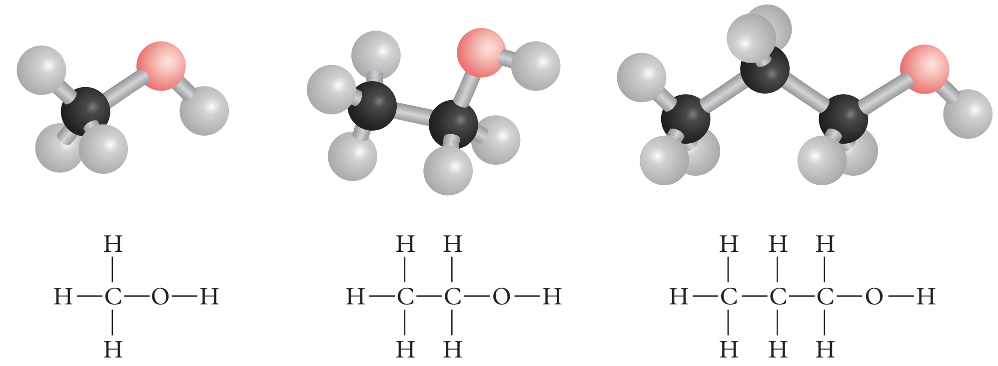


FIGURE 4 Boiling points of straight chain alkanes containing 1–10 carbon atoms

Polar molecules

Primary alcohols are a family of molecules with an OH group attached to an end or terminal carbon (Figure 5).



Methanol (CH_3OH) **Ethanol (C_2H_5OH)** **1-propanol (C_3H_7OH)**

FIGURE 5 Structures of the primary alcohols methanol, ethanol and 1-propanol

The presence of the oxygen atom and the shape of the molecule make primary alcohols polar and they undergo hydrogen bonding with other alcohol molecules due to the presence of the OH groups. The C–O bond is polar and dipole–dipole attractions also occur. However, because they contain a carbon chain, one end of the molecule experiences a small amount of dispersion forces. All of the molecules in this family contain the same amount of hydrogen bonding because they have the same polar OH group and the same dipole–dipole attractions near the C–O bond. Therefore, any differences in boiling points are not due to the polarity of the OH group, but to the length of the carbon chain and therefore the amount of dispersion forces (Figure 6).

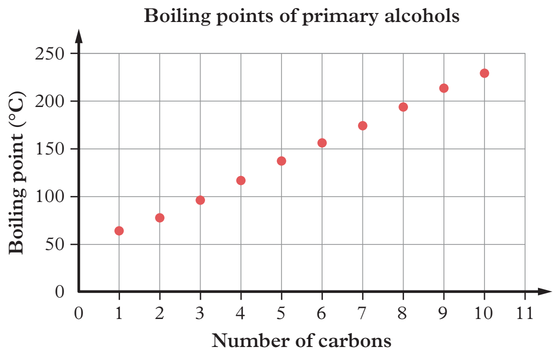


FIGURE 6 Boiling points of primary alcohols containing 1–10 carbon atoms

Comparing polar and non-polar molecules

When the two sets of data are plotted on the same graph, the trend between molecules that have different intermolecular forces can be seen. Because polar molecules that experience hydrogen bonding are more attracted to one another, they are harder to break apart than molecules that contain dispersion forces. Therefore, converting molecules with dispersion forces only (alkanes) from a liquid to a gaseous state requires less energy, and they will have lower boiling points than primary alcohols (Figure 7).

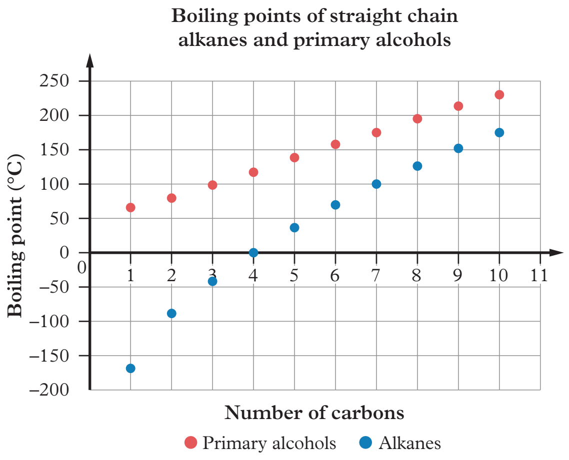


FIGURE 7 Boiling points of straight chain alkanes, interacting through dispersion forces only, and primary alcohols, interacting through hydrogen bonding, dipole–dipole attractions and dispersion forces

Study tip

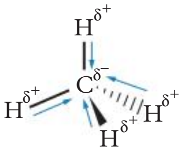
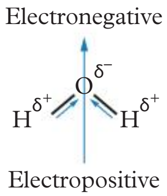
When asked to compare the properties of molecules, comparison or linking words such as “whereas” or “while” must be used. You should also reference the difference in strength or energy required using words such as “less/more/higher/lower/weaker/stronger”. If you say they have “different” energy, you are not demonstrating that you understand what the difference is.

Worked example 9.4A

Comparing molecules to determine melting points

Worked example 9.4A: Watch a video that shows how to solve this problem.

Compare the melting points of methane (CH_4) and water (H_2O). (3 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Compare" means to display recognition of similarities and differences and their significance. We need to find the similarities and differences between the intermolecular forces and therefore melting points of the two molecules. The question is worth 3 marks, so we must analyse the molecules and determine how they are similar and different.
Step 2: Determine the intermolecular forces present. If you need a refresher, review Worked example 9.3B.	Methane is non-polar because its symmetrical tetrahedral structure cancels out the effect of the four C–H bonds, which have small electronegativity differences. Because it is non-polar, it only experiences dispersion forces.  Water is polar because oxygen pulls electrons in the two O–H bonds towards itself, resulting in a partial negative charge on the oxygen end and a partial positive charge on the hydrogen end. The bent shape means the dipoles do not cancel each other out completely, allowing dipole–dipole attractions to form. Because of the presence of the O–H bond, water can hydrogen bond and interact through dispersion forces. (1 mark) 
Step 3: Consider the strength of the intermolecular forces and determine which molecule has the higher melting point.	The dispersion forces between methane molecules are much weaker than the hydrogen bonding between water molecules. (1 mark) Therefore, it takes less energy to disrupt the intermolecular forces between methane molecules, resulting in a lower melting point. (1 mark)
Step 4: Finalise your answer.	Methane is non-polar and molecules interact through by weaker dispersion forces, while water is polar and its molecules interact through stronger hydrogen bonding. Water molecules require more energy to disrupt some of the hydrogen bonding between molecules in the solid, while methane molecules require less energy to separate molecules. Therefore, methane has a lower melting point than water.

Your turn

Compare the boiling points of hydrogen sulfide (H_2S) and ammonia (NH_3). (3 marks)

How do intermolecular forces affect vapour pressure?

Vapour pressure is the pressure of a vapour in contact with its solid or liquid form. The evaporation of water is a common example of this phenomenon. Evaporation is the phase change from a liquid to a gas that occurs on the surface of a liquid, typically at a temperature lower than the boiling point. Water evaporates off the top of liquid water to form water vapour (gaseous water). If vapour pressure is high, it is easier for a liquid to evaporate (or **vaporise**).

A substance that readily vaporises is called a **volatile substance**. It has more molecules entering the gas phase and therefore exerts a higher vapour pressure. The more molecules in the gas phase, the higher the pressure that they exert.

In a container of water that is below boiling point, only the particles on the surface of the water can vaporise. If a water molecule at the bottom of the container randomly gained enough energy to vaporise, it would still need to move to the surface before it could leave the liquid phase to become a gas. However, the molecules at the surface of the water would experience the same distribution of energies.

When a water molecule vaporises, it enters the space just above the water. As more water molecules vaporise, this space contains more and more water molecules.

Molecules within the atmosphere are gaseous and move with high kinetic energy. Low-energy water particles within the atmosphere can **condense** from a gas back into a liquid. This happens more often in a **closed system**, where a lid is placed on the container and none of the water vapour can escape. This can be represented using the Maxwell–Boltzmann distribution (Figure 1). The blue line represents a colder sample where there are more particles with lower kinetic energy. As the temperature of the system increases, the number of particles remains the same but there are now more particles with a higher kinetic energy. The peak of both curves represents the mode of the kinetic energy, the kinetic energy that the majority of the particles contain. The position of the peak and the shape of the curve are related to the temperature of the particles.

Eventually, the rates at which water molecules vaporise (liquid to gaseous) and condense (gaseous to liquid) are equal. The pressure that the water vapour exerts when these rates are equal is called vapour pressure (Figure 8).

vapour pressure
the pressure of a vapour in contact with its solid or liquid form in an enclosed container

vaporise
to change from a liquid to a gas

volatile substance
a substance that easily evaporates from a liquid to a gas

condense
to change from a gas to a liquid

closed system
a system from which gas particles cannot escape

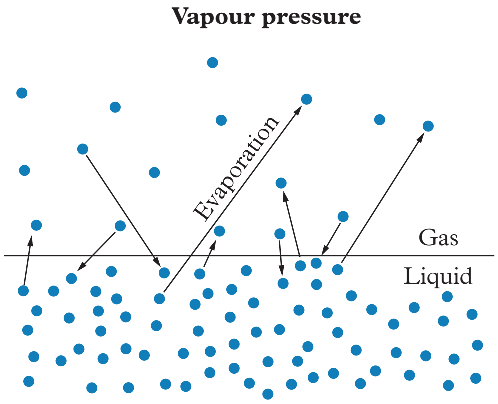


FIGURE 8 Molecules moving from liquid to gas phase and back at the same rate, resulting in vapour pressure

Trends in vapour pressure

Covalent molecules are the more likely to evaporate when they have:

- high kinetic energy
- weak intermolecular forces.

As dispersion forces are the weakest forces, covalent molecules that interact by dispersion forces are more likely to vaporise than molecules that also interact by stronger dipole–dipole attractions or hydrogen bonding.

Real-world chemistry

Water only boils at 100°C at sea level

Sometimes, water boils below 100°C. This is a result of vapour pressure and atmospheric pressure. Atmospheric pressure is the collision of air molecules with each other and against the surface of Earth. The more collisions there are, the more pressure is exerted. At higher altitudes, pressure decreases because the air is less dense and there are fewer collisions.

When water molecules evaporate, they push up against the gas molecules in the atmosphere, reducing the amount of pressure pushing down on the water. This makes it easier for the water to form a bubble of gas below the surface and for the bubble to rise to the surface. Therefore, boiling is more likely when vapour pressure increases (allowing bubbles to rise) or atmospheric pressure decreases (exerting less pressure on the water). The boiling point is reached when vapour pressure is equal to the pressure above the liquid. Different substances have different intermolecular forces and therefore vapour pressures, so they also have different boiling points.

The boiling point of water, 100°C (or more accurately, 99.61°C), is only theoretical and occurs when the vapour pressure of water is equal to the atmospheric pressure at sea level, which is 1 bar or 100 kPa. The graph in Figure 9 demonstrates the relationship between vapour pressure and the boiling point of water. If atmospheric pressure is reduced, the boiling point of water also decreases.

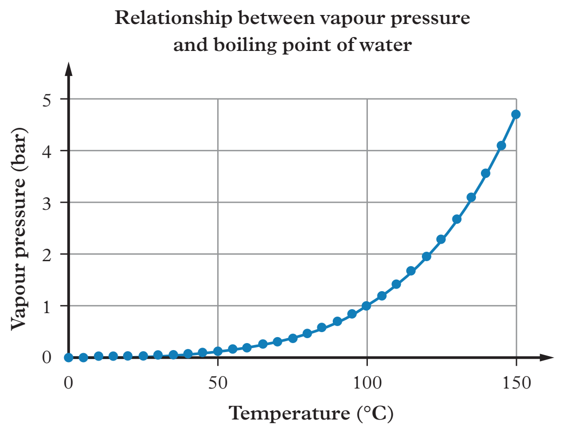


FIGURE 9 The relationship between vapour pressure (bar) and boiling point (°C) of water

Apply your understanding

1 **Describe** the relationship between boiling point and vapour pressure, using Figure 9 to **justify** your response. (2 marks)

2 **Explain** whether a substance with stronger intermolecular forces than water would have a higher or lower vapour pressure at the same temperature. (2 marks)

3 **Predict** how a graph for the relationship between boiling and vapour pressure in ammonia (NH_3) would be different to that for water. (2 marks)

How do intermolecular forces affect solubility?

The ability for a solute to form strong intermolecular forces with solvent molecules and form a **solution** is called its **solubility**. A **solute** will **dissolve** into a **solvent** to form a solution if the solute and solvent have similar intermolecular properties. This is called the "like dissolves like" rule. A polar solute dissolves in a polar solvent because they both contain partial charges. Therefore, polar substances dissolve in water because their permanent dipoles strongly attract each other.

For example, water is a covalent molecule and a common solvent. Water molecules are strongly attracted to other water molecules because of their polarity and ability to form hydrogen bonds.

solution
a mixture of a solute dissolved in a solvent

solubility
the ability of a substance to dissolve in another to form a solution; the maximum amount of solute that dissolves in a known quantity of solvent at a given temperature

solute
the minor component of a solution; a substance dissolved in a solvent to form a solution

Non-polar molecules form layers on top of or underneath polar molecules (depending on their relative densities) because the polar molecules experience greater interaction with each other than they do with the non-polar molecules. Figure 10 shows how oil forms a layer on top of water.

This happens for two reasons:

- The hydrogen bonding between the water molecules is too strong for any oil molecule to break through and mix into the water.
- Oil, despite being a larger molecule, has a lower density than water and forms the top layer. This is because the hydrogen bonding between water molecules makes then more strongly attracted to each other and therefore they are packed tighter, making water more dense.

Non-polar substances require non-polar solvents in order to dissolve. For example, turpentine is a non-polar solvent and is used to dissolve oil-based paint residue on paintbrushes. Can you think of any other examples?

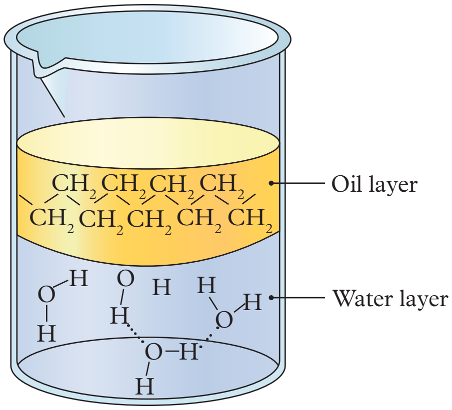


FIGURE 10 Layers form when polar water molecules attract more strongly to themselves than to non-polar oil molecules.

dissolve
when a substance goes into solution by forming new interactions with the solvent

solvent
the major component of solution; a substance that can dissolve other substances

Study tip

The "like dissolves like" rule will help you remember that substances dissolve in solvents with similar properties, but must not be used in your responses. You must state the intermolecular forces involved and whether these forces result in interactions between molecules.

Worked example 9.4B

Determining whether substances dissolve based on their intermolecular forces

Worked example 9.4B: Watch a video that shows how to solve this problem.

Explain whether ethanol (C_2H_5OH) dissolves in water. (3 marks)

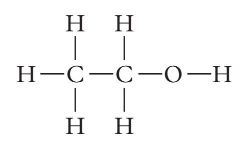


FIGURE 11 Structure of ethanol

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Explain" means to make an idea or situation plain or clear by describing it in more detail or revealing relevant facts. We need to identify the key properties of the two molecules and apply the "like dissolves like" rule. The question is worth 3 marks, so we must analyse the molecules and determine whether one dissolves in the other.
Step 2: Determine the intermolecular forces present. If you need a refresher, review Worked example 9.3B.	Ethanol is polar because it contains an OH group on one end. It has a molecular dipole and is capable of hydrogen bonding. (1 mark) Water is polar because it contains two O–H bonds and its shape results in a partial negative charge on oxygen and partial positive charges on the hydrogen atoms. It has a molecular dipole and is capable of hydrogen bonding. (1 mark)
Step 3: Recall the "like dissolves like" rule to determine whether one molecule will be soluble in the other.	Ethanol and water can interact with one another by hydrogen bonding, so the ethanol will be soluble in water. (1 mark)
Step 4: Finalise your answer.	Ethanol (the solute) dissolves in water (the solvent) because they interact by hydrogen bonding to form a solution of diluted ethanol.

Your turn

Explain whether octane (C_8H_{18}) dissolves in water (H_2O). (3 marks)

Check your learning 9.4

Check your learning 9.4: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** why kinetic energy affects the ability of a substance to melt or boil. (5 marks)

2 **Explain** the trend in melting and boiling points of covalent molecules, using examples. (4 marks)

3 **Explain** how vapour pressure relates to the boiling point of a substance. (2 marks)

Analytical processes

4 **Consider** why F_2 has a boiling point of −188°C, but HF has a boiling point of 19.5°C. (2 marks)

5 The structure of the glucose molecule is shown in Figure 12.

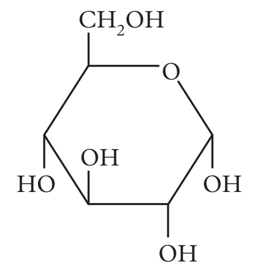


FIGURE 12 Structure of glucose

Use this structure to **deduce** whether glucose will dissolve in:

a water (H_2O) (2 marks)

b decane ($C_{10}H_{22}$). (2 marks)

Knowledge utilisation

6 **Determine** which of the following molecules has a greater solubility in water. **Justify** your answer. (4 marks)

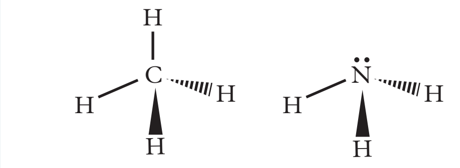


FIGURE 13 Lewis structures of methane and ammonia

7 **Discuss** the trend in boiling points of the following hydrogen halide compounds. (4 marks)

TABLE 1 Boiling points of hydrogen halides

Molecule	Boiling point (°C)
HF	19.5
HCl	–85.05
HBr	–66
HI	–35.36

MODULE 9

Lesson 9.5

Review: Intermolecular forces

Summary

9.1

- The shape and bond angles of a covalent molecule is determined by the valence shell electron pair repulsion (VSEPR) theory.
- Covalent compounds can have a linear, bent, trigonal planar, tetrahedral or pyramidal shape.

9.2

- Practical: Constructing 3D models of molecules

9.3

- The polarity of a covalent bond depends on the electronegativity of the bonded atoms.
- The polarity of a covalent molecule depends on its molecular shape and the sum of the polarity of its bonds.
- Intermolecular forces between molecules include dispersion forces, dipole–dipole attractions and hydrogen bonding.

9.4

- When intermolecular forces between the molecules of a substance are strong, their melting point and boiling points are high.
- When intermolecular forces between the molecules of a substance are strong, their vapour pressure is lower.
- When two different substances have similar polarities, they are likely to mix and dissolve in each other.

Review questions 9.5A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 The molecule ammonia is

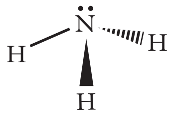


FIGURE 1 Lewis structure of ammonia

A tetrahedral and polar.
B pyramidal and non-polar.
C pyramidal and polar.
D tetrahedral and non-polar.

2 In liquid hydrogen (H_2), the intermolecular forces present are

A dispersion forces.
B dipole–dipole attractions.
C hydrogen bonding.
D ionic bonding.

3 Vapour pressure is greatest when

A a substance has minimal evaporation.
B atmospheric pressure is increased.
C a substance has weak intermolecular forces.
D a substance has low kinetic energy.

4 To melt a covalent substance, it is best to

A maximise kinetic energy and intermolecular forces.
B maximise kinetic energy and minimise intermolecular forces.
C minimise kinetic energy and maximise intermolecular forces.
D minimise kinetic energy and minimise intermolecular forces.

5 What is the bond angle in the water molecule?

A 104.5°
B 105.5°
C 106.5°
D 107.5°

6 Which of the following elements has the highest electronegativity?
- **A** Carbon
- **B** Iron
- **C** Lithium
- **D** Magnesium

7 Which of the following compounds is polar?
- **A** Ammonia
- **B** Methane
- **C** Hydrogen gas
- **D** Carbon tetrachloride

Review questions 9.5B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

8 **Construct** a Lewis structure for the following molecules, indicating which atoms are electropositive and which are electronegative, the direction of electron pull and the overall effect. **Determine** whether the molecules are polar or non-polar. (4 marks)
- **a** CCl_4 (4 marks)
- **b** HF (4 marks)
- **c** HCN (4 marks)
- **d** BH_3 (4 marks)
- **e** CH_3NH_2 (4 marks)

9 **Describe** how a covalent bond is different from an intermolecular force. (2 marks)

10 **Explain** the difference between melting point and boiling point by using the terms “kinetic energy” and “intermolecular forces”. (7 marks)

11 **Identify** and **explain** which type of intermolecular force is the strongest. (2 marks)

12 **Explain** how non-polar molecules are able to interact if they have no overall charges. (3 marks)

13 **Explain** the difference in melting points of sodium chloride (NaCl) 801°C, water (H_2O) 0°C and methane (CH_4) −182°C. (6 marks)

14 **Explain** whether substances with higher boiling points have stronger or weaker intermolecular forces. (2 marks)

15 **Explain** the relationship between the length of an alkane and its boiling point. (2 marks)

16 **Explain** the intermolecular forces present in the molecule DNA and why it is essential to the replication of the molecule. (3 marks)

17 **Explain** why ethanol has a higher vapour pressure compared to water at any given temperature. (3 marks)

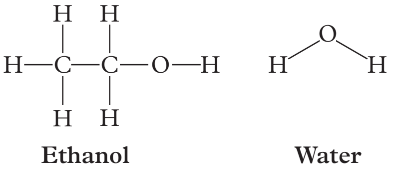


FIGURE 2 Structures of ethanol and water

18 **Explain** why methane (CH_4) has a boiling temperature of −161.5°C and is a gas at room temperature, whereas octane (C_8H_{18}) has a boiling temperature of 125.7°C and is a liquid at room temperature. (4 marks)

19 **Explain** why there is little difference in melting points of fluorine (F_2) −220°C, oxygen (O_2) −218.8°C and nitrogen (N_2) −210°C. (3 marks)

Analytical processes

20 The skeletal structures of four fatty acids are shown. Each bend represents a carbon atom; hydrogen atoms are not represented.
- **a** **Explain** how the VSEPR model has helped chemists understand the physical properties of fatty acids. (4 marks)
- **b** **Consider** the benefit in drawing a skeletal structure shown below over the 2D structure. (2 marks)

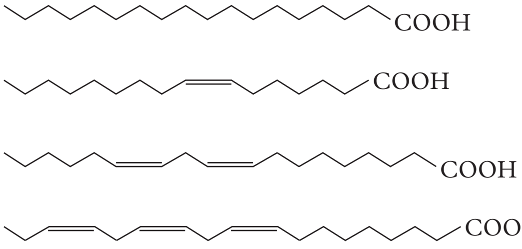


FIGURE 3 Skeletal structures of four fatty acids

21 **Analyse** the substances listed and **identify** their shape, polarity, and the types of intermolecular forces they form. (27 marks)

- HF
- H_2S
- PF_3
- CO_2
- CCl_4
- BF_3
- O_2
- SO_2
- NH_3

22 **Consider** which of the molecular shapes discussed in Lesson 9.1 result in a polar compound, using an example for each shape. (10 marks)

23 **Determine** the type of intermolecular forces that exist between the following pairs of molecules.

a H_2S and H_2O

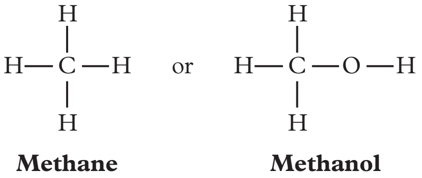
b CO_2 and N_2 (1 mark)

c NH_3 and H_2O (1 mark)

d CH_3OH and CH_3CH_2COOH (1 mark)

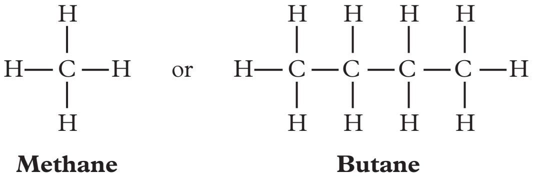
24 **Determine** which of the following pairs of molecules would have a higher boiling point. **Explain** your reasoning.

a



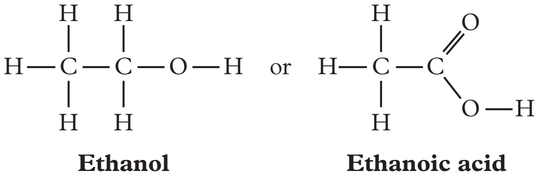
(5 marks)

b



(3 marks)

c



(4 marks)

Knowledge utilisation

25 A student claims that all substances experience intermolecular forces. **Evaluate** whether they are correct. (3 marks)

26 **Propose** an explanation for why fluorine (F) is the most electronegative element on the periodic table. (2 marks)

27 The boiling point of 1-propanol is 97°C and that of ethyl methyl ether is 7.4°C. The two molecules have the same number of carbon, oxygen and hydrogen atoms. **Discuss** the difference in their boiling points. (5 marks)

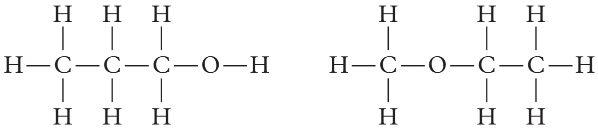


FIGURE 4 Structures of 1-propanol and ethyl methyl ether

28 Cellulose is a polymer of the molecule glucose.

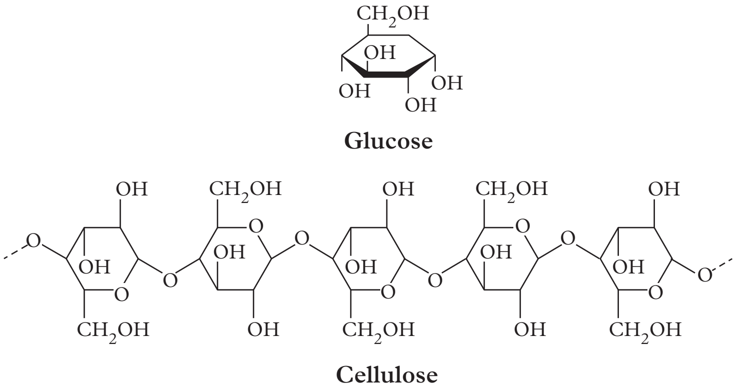


FIGURE 5 Structures of glucose and cellulose

Cellulose makes up the cell wall of plants and forms a very rigid structure. This structure is insoluble in water, whereas the glucose molecule is highly soluble in water. **Hypothesise** why is glucose soluble, but cellulose is insoluble. (2 marks)

29 **Use** your knowledge of the structure of water and its intermolecular forces to **propose** an explanation for why water expands when it is in solid form and contracts when it is in liquid form. (5 marks)

30 Carbon dioxide is a gas at room temperature and a solid (dry ice) at –78.5°C. When dry ice is left on a laboratory bench, it turns into a gas in a process called sublimation. **Propose** a reason why carbon dioxide turns from a solid to a gas, skipping the liquid phase. (2 marks)

FIGURE 6 Dry ice

Data drill

Trends in physical properties

The melting and boiling points for a series of alkanes are shown in Table 1.

TABLE 1 Melting and boiling points of alkanes

Name	Formula	Melting point (°C)	Boiling point (°C)	State at room temperature
Methane	CH_4	–182.5	–161.5	gas
Ethane	C_2H_6	–183.3	–88.6	gas
Propane	C_3H_8	–187.7	–42.1	gas
Butane	C_4H_{10}	–138.4	–0.5	gas
Pentane	C_5H_{12}	–129.7	36.3	liquid
Hexane	C_6H_{14}	–95.3	68.7	liquid
Heptane	C_7H_{16}	–90.6	98.4	liquid
Octane	C_8H_{18}	–56.8	125.7	liquid
Nonane	C_9H_{20}	–51	151	liquid
Decane	$C_{10}H_{22}$	–30	174	liquid

Apply understanding

1 **Identify** the trend in the boiling point data. (1 mark)

2 **Determine** which alkane has the strongest intermolecular forces and **identify** the types of intermolecular forces present. (2 marks)

Analyse data

3 **Organise** the melting and boiling point data from Table 1 into a scatterplot. (3 marks)

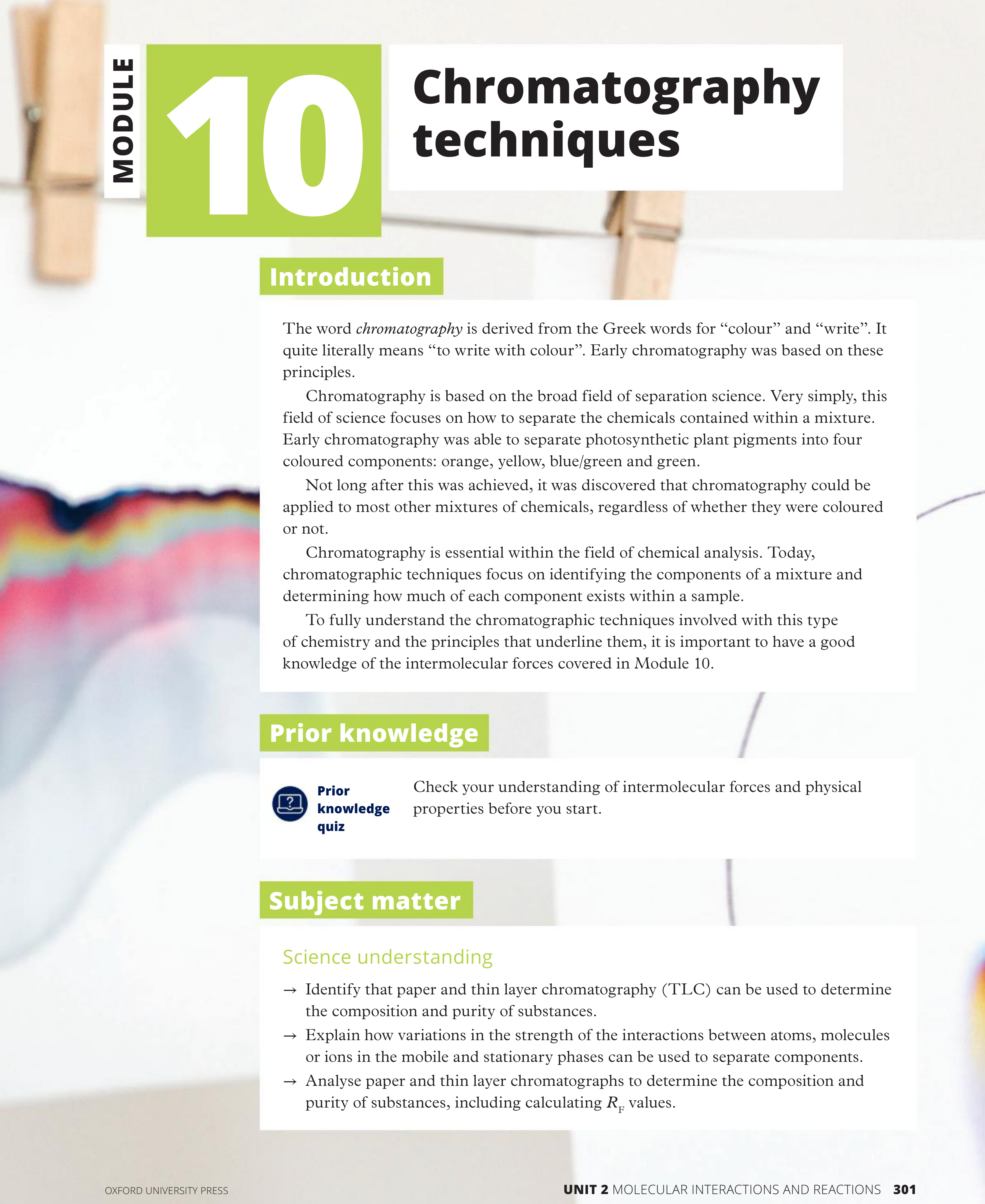
Evaluate evidence

4 **Determine** one trend in the melting point data and **infer** why this trend may exist. (2 marks)

5 **Draw a conclusion** about the data. (2 marks)

Module 9 checklist: Intermolecular forces

Quizlet: Revise key terms online to test your understanding



MODULE 10 Chromatography techniques

Introduction

The word *chromatography* is derived from the Greek words for “colour” and “write”. It quite literally means “to write with colour”. Early chromatography was based on these principles.

Chromatography is based on the broad field of separation science. Very simply, this field of science focuses on how to separate the chemicals contained within a mixture. Early chromatography was able to separate photosynthetic plant pigments into four coloured components: orange, yellow, blue/green and green.

Not long after this was achieved, it was discovered that chromatography could be applied to most other mixtures of chemicals, regardless of whether they were coloured or not.

Chromatography is essential within the field of chemical analysis. Today, chromatographic techniques focus on identifying the components of a mixture and determining how much of each component exists within a sample.

To fully understand the chromatographic techniques involved with this type of chemistry and the principles that underline them, it is important to have a good knowledge of the intermolecular forces covered in Module 10.

Prior knowledge

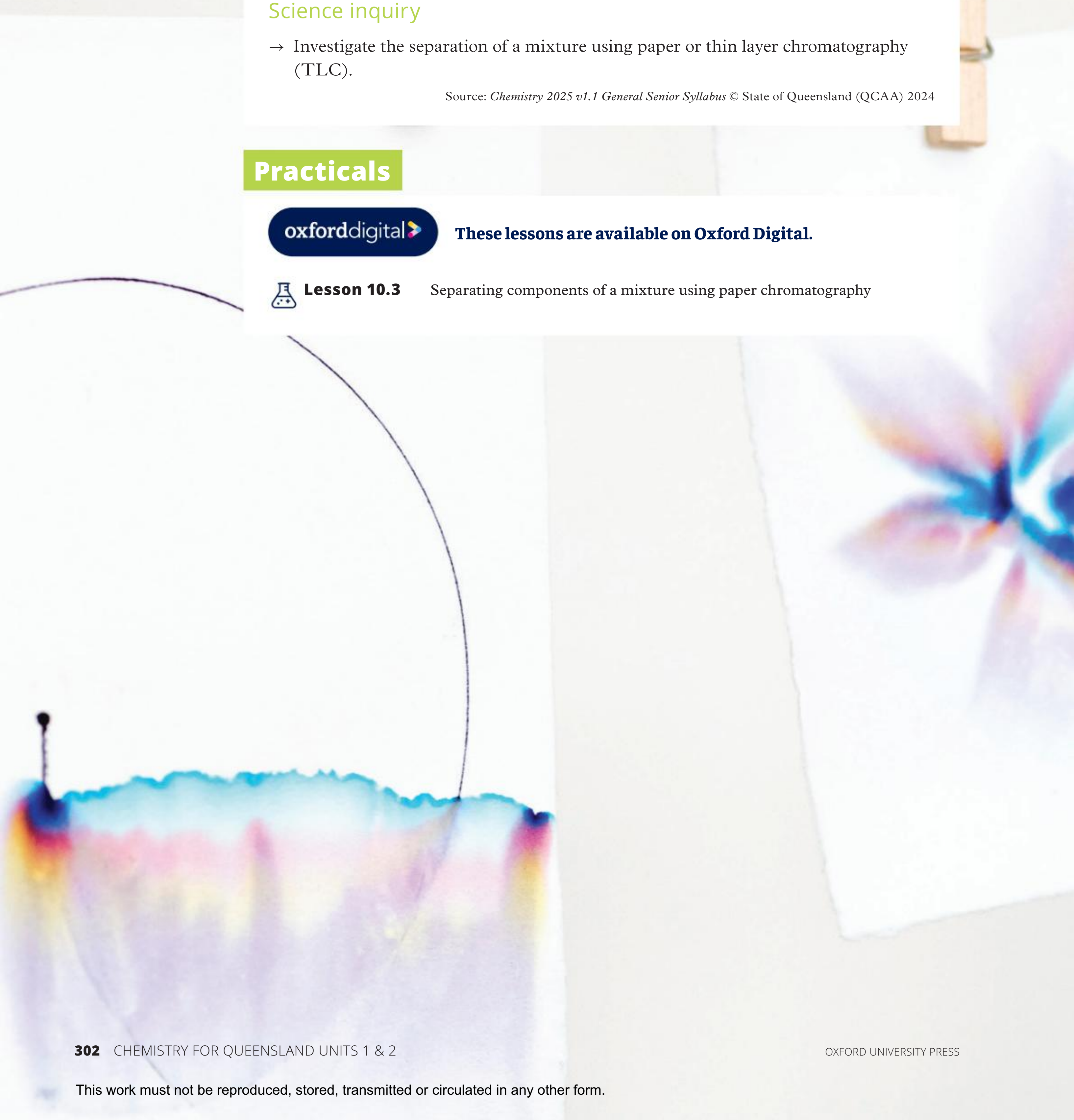
Prior knowledge quiz

Check your understanding of intermolecular forces and physical properties before you start.

Subject matter

Science understanding

- → Identify that paper and thin layer chromatography (TLC) can be used to determine the composition and purity of substances.
- → 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 paper and thin layer chromatographs to determine the composition and purity of substances, including calculating R_F values.



Science as a human endeavour

- → Appreciate that science relies on chemical processes to analyse materials in order to determine their identity, nature or source of the material.
- → Explore how chromatography techniques, including gas and high-performance liquid chromatography, can be used to determine the composition and purity of substances.

Science inquiry

- → Investigate the separation of a mixture using paper or thin layer chromatography (TLC).

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

oxforddigital **These lessons are available on Oxford Digital.**

Lesson 10.3 Separating components of a mixture using paper chromatography

Lesson 10.1

Principles of chromatography

Key ideas

- → Chromatography separates components in a mixture based on the intermolecular interactions between atoms, molecules or ions in the mobile and stationary phases.
- → The adsorption and desorption of components of a mixture depends on the strengths of their intermolecular forces.

What is chromatography?

Chromatography is an analytical technique that can be used for qualitative or quantitative analysis. It can both identify components in a mixture (qualitative analysis) and determine how much of each component is present (quantitative analysis). It is a useful form of separation science that uses intermolecular forces – which you learnt about in the previous module – to separate the components of a mixture according to their properties.

Chromatography has a wide range of uses, including in:

- pharmaceutical science – identifying elements and molecules in drugs as well as measuring the purity of medicines
- environmental science – analysing water quality and gases in the air and the impact they may have on Earth
- police work – analysis such as breathalysers and drug detection
- forensic science and archaeology – analysing materials such as blood and hair samples
- biomedical research – analysing material such as proteins in cancer research
- food science – analysing the nutritional quality of food and food spoilage.

Learning intentions and success criteria

chromatography an analytical technique that separates the components of a sample mixture based on the properties of the molecules

FIGURE 1 Chromatography has many uses.

What are the stationary and mobile phases?

Chromatography uses two phases to separate a mixture:

- **mobile phase** – a solvent that moves over or through the stationary phase; this can be a liquid or gas
- **stationary phase** – a substance that stays still (either a solid with a high surface area or a liquid coated onto a solid support).

The mobile phase moves through and carries a sample (often a mixture of solutes) over the stationary phase.

mobile phase
the phase that flows in chromatography, moving the components of a sample at different rates over the stationary phase

stationary phase
the phase to which the components of a chromatographic sample are adsorbed

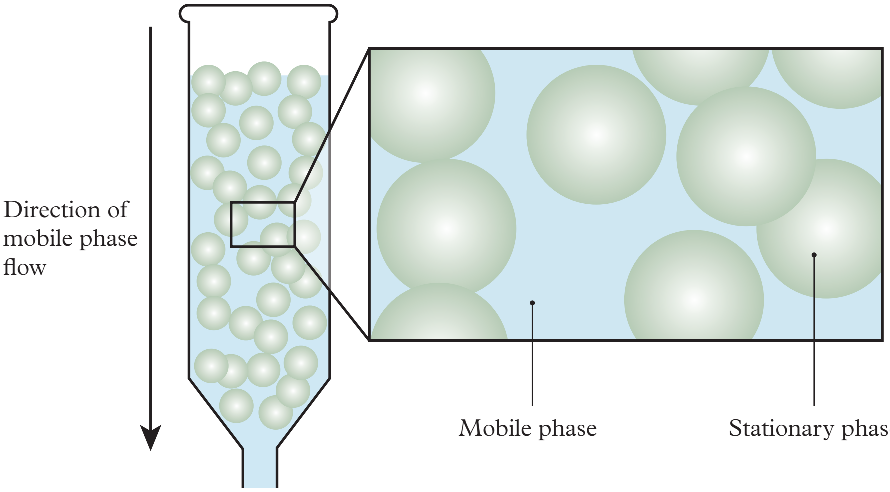


FIGURE 2 The mobile and stationary phases in chromatography

How are components separated in chromatography?

Components of the sample are more attracted to either the mobile or the stationary phase, depending on their intermolecular forces. The strengths of these interactions are referred to as the components' **affinity** for each phase. High affinity means that the component interacts strongly with the mobile or stationary phase.

The different affinities that each component has for the phases means they move at different rates and therefore, the components separate.

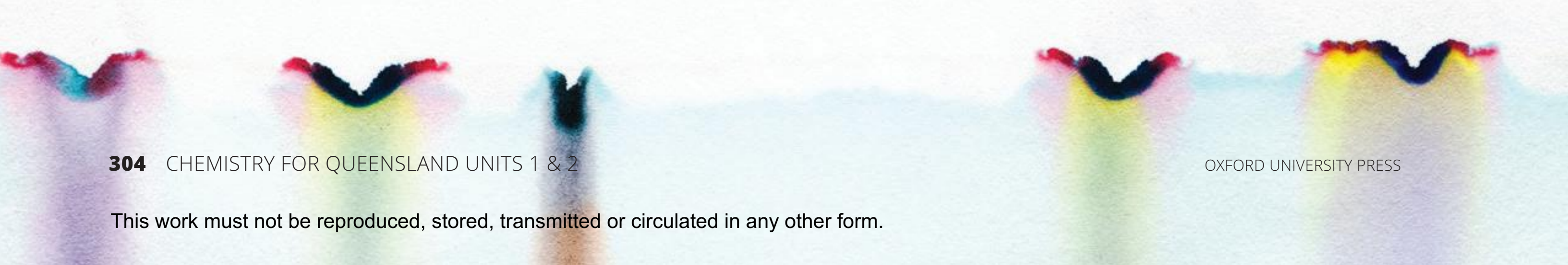
- **Adsorption**: components of the sample adsorb (stick) onto the stationary phase from the mobile phase if they are attracted to, or have an affinity for, the stationary phase.
- **Desorption**: components of the sample desorb off (are released from) the stationary phase into the mobile phase if they are attracted to, or have an affinity for, the mobile phase.

affinity
the interaction of a substance within a sample with the mobile or stationary phases

adsorption
the attraction of a substance within a sample to the stationary phase

desorption
the release of a substance within a sample from the stationary phase into the mobile phase

FIGURE 3 Chromatography separates components of mixtures based on their relative affinities for the mobile and stationary phases.



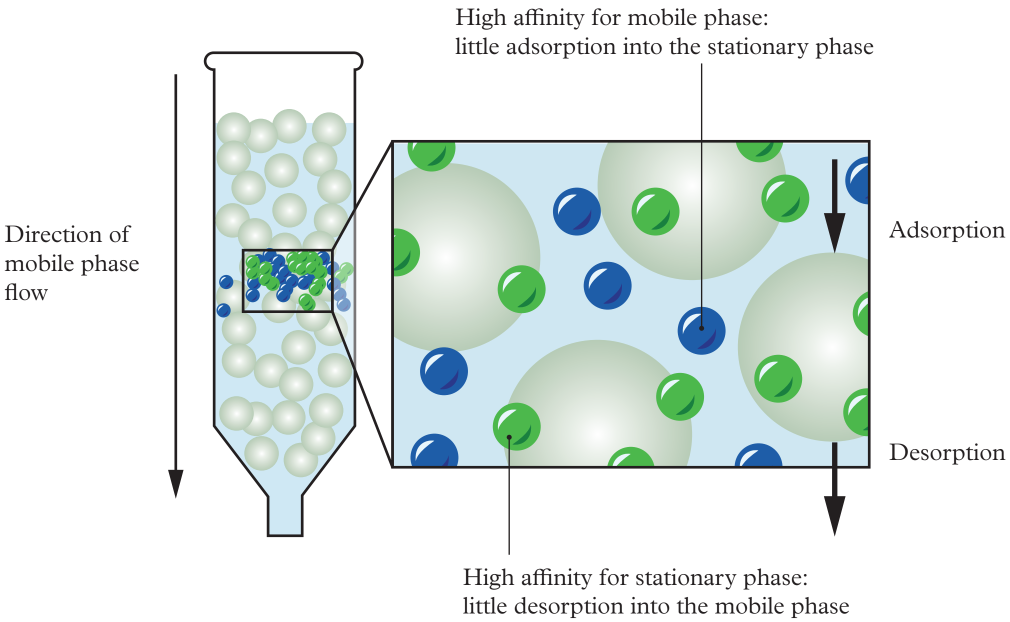


FIGURE 4 The blue and green components in this sample can be separated depending on their affinity for the mobile and stationary phases.

Study tip

*Ad*sorption is when a substance sticks to the surface to form a film. *Ab*sorption is when a substance penetrates a solid or liquid.

The effect of different component affinities for the mobile and stationary phases can be seen in Figure 5.

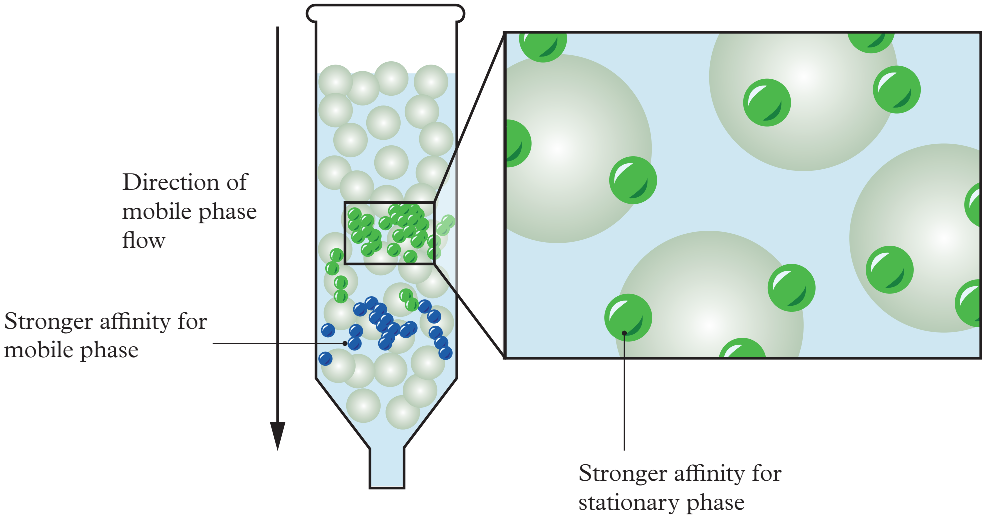


FIGURE 5 The components separate out according to their differing affinities.

Study tip

Chromatography is based on intermolecular forces, so, when answering questions, start by determining which intermolecular forces are involved. Once you have determined the intermolecular forces, explain their affinity to the mobile or stationary phase. Then, explain the effect this will have on the data you obtain.

The "like dissolves like" rule

Generally, atoms, molecules and ions are more attracted to the mobile or stationary phase that has the same or similar properties. This is the "like dissolves like" rule, which you learnt about in the previous module. For example, ions contain one or more whole charges. Because of this, they are more attracted to the mobile phase if it contains charged or partially charged molecules or ions, such as water (a polar molecule).

Skill drill

Extending an experiment about intermolecular forces

Science inquiry skill: Evaluating evidence (Lesson 1.8)

To determine whether molecules are polar, a chemist dissolves them in water in a test tube. She hypothesises that if the solute is polar, it will dissolve in the water solvent, and if it is non-polar, it will form layers. The chemist tests six substances and records her results in Table 1.

TABLE 1 Results of the experiment

Solute	Result
Sodium chloride	dissolves into water
Potassium nitrate	dissolves into water
Methanol	dissolves into water
Glucose	dissolves into water
Ethanol	dissolves into water
Octane	forms a layer with water

Practise your skills

1 **Identify an error** in the chemist's hypothesis. (1 mark)

2 **Explain** why the data collected is qualitative and not quantitative. (1 mark)

How does carbon chain length or molar mass affect separation?

Chemists often need to separate mixtures of molecules that have similar intermolecular forces such as mixtures of alcohols. In a mixture of methanol, ethanol, propan-1-ol and butan-1-ol (Figure 6), all of the molecules have the same polar –OH group on the terminal, or end, carbon. Therefore, this mixture will not be separated into its components through intermolecular forces due to the –OH group. Rather, the mixture is separated by the size of the carbon chain – the longer the carbon chain, the greater the dispersion forces and the higher the interaction with the mobile or stationary phases.

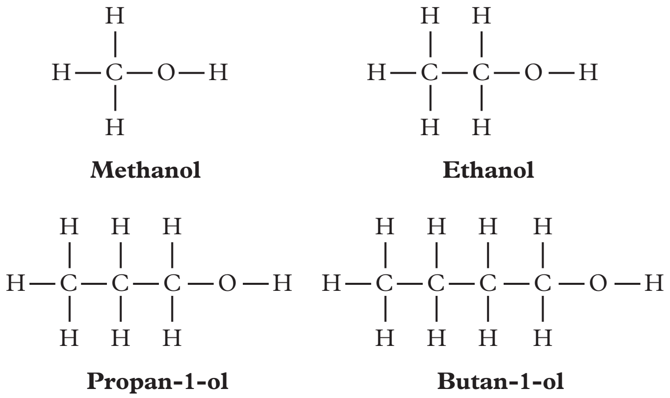


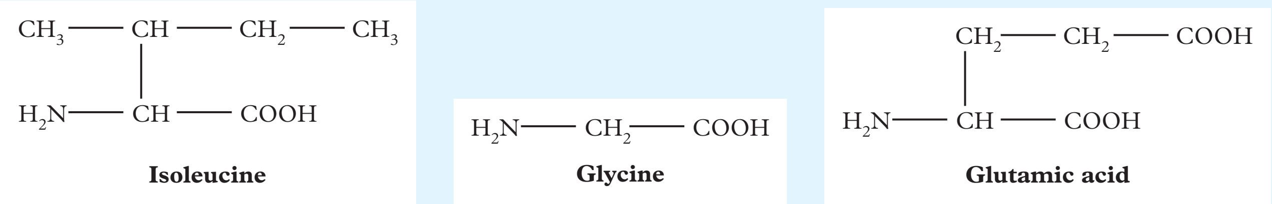
FIGURE 6 Methanol, ethanol, propan-1-ol and butan-1-ol all have a terminal polar OH group.

Worked example 10.1A

Assessing the relative affinity of components to the stationary phase

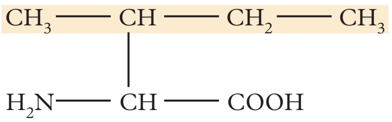
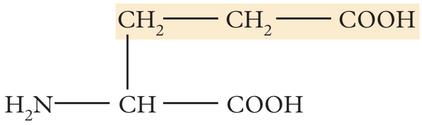
Worked example 10.1A: Watch a video that shows how to solve this problem.

A mixture of amino acids is separated using a polar stationary phase and a less polar mobile phase. **Explain** which will have the highest and lowest affinity to the stationary phase. (4 marks)



Isoleucine **Glycine** **Glutamic acid**

FIGURE 7 Structures of isoleucine, glycine and glutamic acid.

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Explain" means to make an idea or situation plain or clear by describing it in more detail or revealing relevant facts. We need determine what intermolecular forces each substance can form and compare them to the intermolecular forces in the mobile and stationary phases. The question is worth 4 marks, so we must analyse the substances, select the two correct substances, and provide explanations for the decisions.
Step 2: Determine the difference in the intermolecular forces experienced by the molecules.	All of the molecules are polar as they contain a COOH and NH_2 functional group; they are capable of hydrogen bonding. Therefore, any difference in their properties is due to the carbon chain at the top of the molecule (highlighted in the structures below). Glycine has a H group, which is non-polar and experiences weak dispersion forces. $H_2N—CH_2—COOH$ Isoleucine has a $CH_3CHCH_2CH_3$ chain, which is non-polar and also experiences dispersion forces but these are stronger than for glycine due to the length of the chain.  Glutamic acid has a CH_2CH_2COOH group, which is polar due to the COOH and hydrogen bonds. 
Step 3: Determine which will substances have a stronger affinity to the stationary phase based on the strength of their intermolecular forces.	The molecules with stronger intermolecular forces will experience greater interactions with a polar stationary phase and therefore have a higher affinity to it.
Step 4: Finalise your answer.	Glutamic acid will have the greatest affinity to the stationary phase (1 mark) as it has stronger intermolecular forces and will experience a greater interaction to a polar stationary phase (1 mark), whereas isoleucine (1 mark) will have the lowest as it has the weakest intermolecular forces (1 mark).

Your turn

A mixture if alcohols and acids is separated using a slightly polar stationary phase and a more polar mobile phase. **Explain** which will have the highest and lowest affinity to the stationary phase. (4 marks)

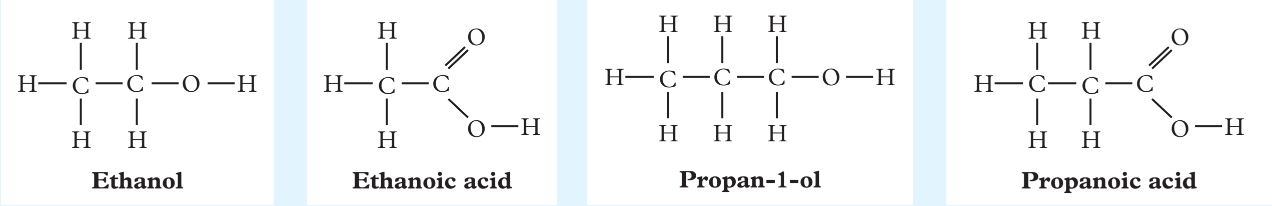


FIGURE 8 Structures of ethanol, ethanoic acid, propan-1-ol and propanoic acid

Check your learning 10.1

Check your learning 10.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define** the following terms and **explain** why they are essential to the separation of mixtures.
  - **a** mobile phase (2 marks)
  - **b** stationary phase (2 marks)
  - **c** affinity (2 marks)
  - **d** intermolecular forces. (2 marks)

2 **Explain** whether polar components of a mixture are more attracted to a polar or non-polar mobile phase, using an example. (3 marks)

Analytical processes

3 **Compare** the processes of adsorption and desorption. (2 marks)

4 A chemist is separating a mixture containing ethyl ethanoate and ethanol. **Determine** which substance will have a stronger affinity for a polar mobile phase. **Explain** why. (3 marks)

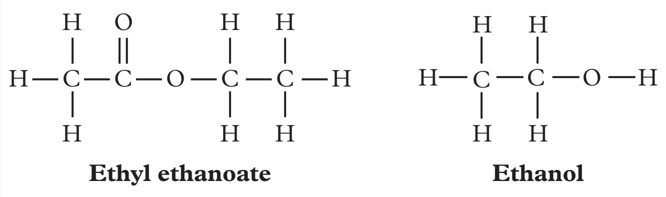


FIGURE 9 Structures of ethyl ethanoate and ethanol.

5 A chemist needs to separate a sample of dyes that has different polarities (Table 2).

TABLE 2 Polarities of components A, B and C

Component	Polarity
A	Slightly polar
B	Polar
C	Least polar

  - **a** **Determine** which component would have the strongest affinity for a mobile phase consisting of a polar solvent. **Explain** your response. (2 marks)
  - **b** **Determine** which component would have the strongest affinity for a non-polar stationary phase. **Explain** your response. (2 marks)
  - **c** **Determine** which component would adsorb most readily to a polar stationary phase. **Explain** your response. (2 marks)
  - **d** **Determine** which components would desorb most readily to a polar mobile phase. **Explain** your response. (2 marks)

Knowledge utilisation

6 **Discuss** which intermolecular forces, studied in Module 9, could be used to separate a mixture of substances. (3 marks)

7 A mobile phase chosen in an analysis contains molecules that only interact through dispersion forces. **Determine** what intermolecular forces the stationary phase should form. **Justify** why. (2 marks)

Lesson 10.2

Paper and thin layer chromatography

Key ideas

→ Paper and thin layer chromatography are the simplest chromatographic techniques for separating a sample. Both are qualitative – they can be used to determine the purity and/or composition of substances.

→ Chromatographs are analysed by calculating the R_F value.

Learning intentions and success criteria

What is paper chromatography?

Paper chromatography is a type of chromatography in which a sample containing a mixture of substances is applied to the stationary phase as a spot. The spot is carried up the stationary phase by the mobile phase. The resulting **chromatograph** is usually qualitative in nature – it can only be used to identify substances, not determine how much is present.

The steps of paper chromatography are summarised in Figure 1.

1 The stationary phase is a thin strip of absorbent paper, which is cut to fit inside a container that holds the mobile phase. A pencil line is ruled across the bottom of the paper to mark where the sample will be placed – this is called the **origin**. Pen or ink should not be used to mark the origin because inks will separate and contaminate the sample.

2 The sample is placed on the paper and standards of known identity can also be added to help identify the substances in the sample. In Figure 1, the standards are marked A, B and C.

3 The mobile phase is added to the container. The paper is then placed into the container so that the origin marked on the paper is above the level of the mobile phase.

4 The mobile phase, along with the sample, travels up the piece of paper, separating the sample into its components. Components of the sample that have a strong affinity for the stationary phase move slowly. Components that have a higher affinity for the mobile phase move faster.

5 On the resulting chromatograph, the components have moved further and are located closer to the **solvent front**.

paper chromatography
an analytical technique for separating and identifying mixtures; the stationary phase is a thin strip of absorbent paper

chromatograph
the pattern of bands, spots or peaks formed on the chromatography paper or TLC plate demonstrating the separation of a mixture

origin
the line applied to a chromatograph to mark the point where the sample or standard was placed

solvent front
the point on a chromatograph where the mobile phase reaches before the analysis is terminated

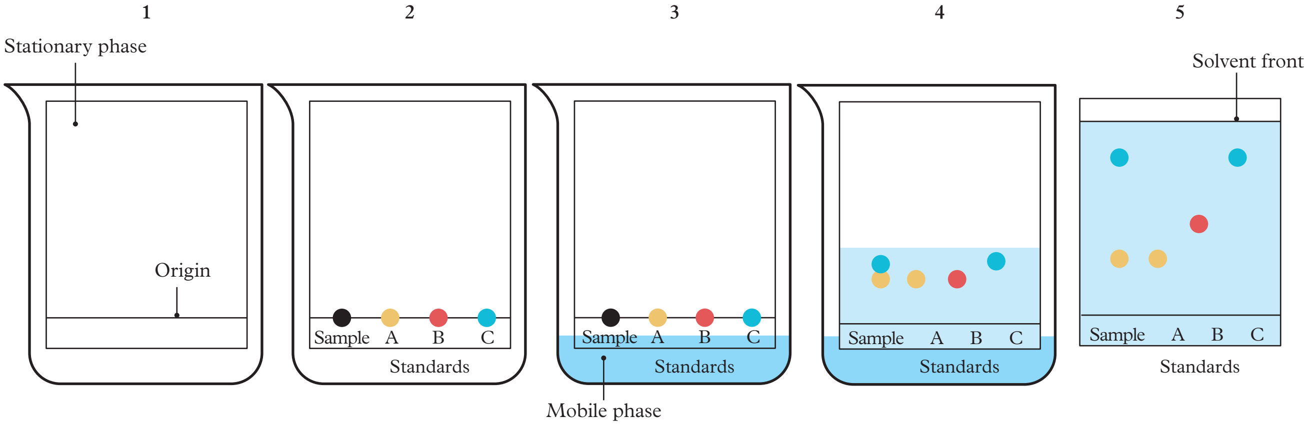


FIGURE 1 The water mobile phase moves up a piece of chromatography paper (stationary phase) to separate the ink in black marker pen (sample). Components in the sample can be identified by comparing them to known standards A–C.

How are paper chromatographs analysed?

The mobile phase is water and the stationary phase is filter paper. Paper is a derivative of cellulose, which contains many polar –OH groups. However, very few of the intermolecular forces of paper extend beyond its network of fibres (and any surface coating). Water is also a polar molecule, but more polar than the paper. Therefore, any component of the sample that has whole or partial charges, such as ionic or polar substances, will have a higher affinity, or attraction, for the water mobile phase. These components move further from the origin. Any component of the sample that experiences only dispersion forces has a higher affinity for the stationary phase and will not move as far from the origin.

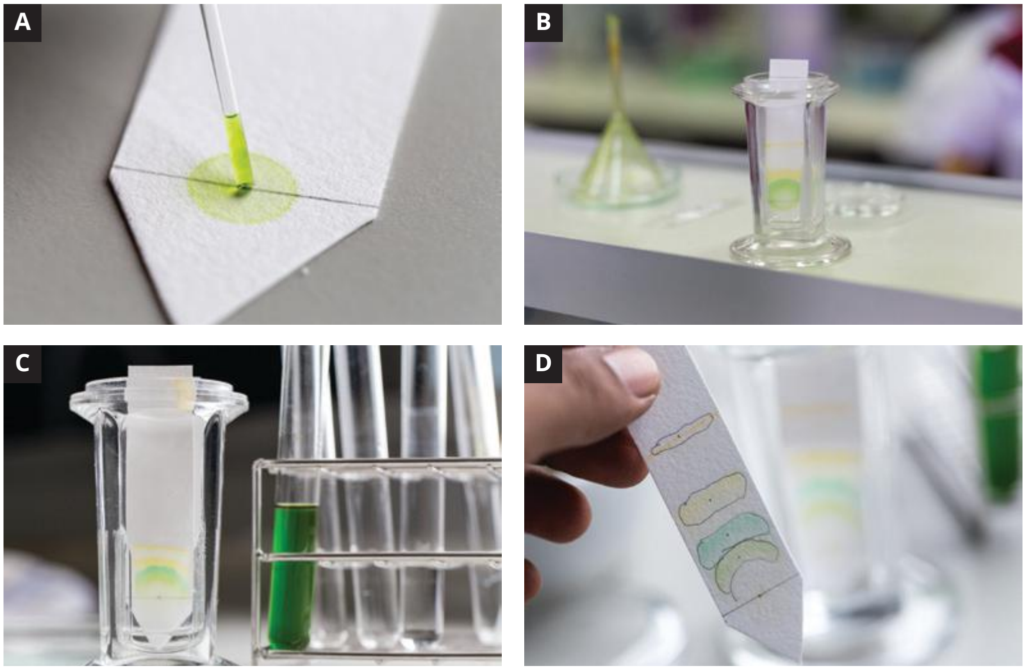


FIGURE 2 Paper chromatography can be used to separate the components of plant pigment (leaf stain). The pigment mixture is placed on the origin, and then the paper strip is placed in a container where it is in contact with the mobile phase, which is drawn up the paper by capillary action.

Study tip

Never leave R_F as a fraction. This is because it is reported on a scale of 0–1, with 0 being no affinity to the mobile phase and 1 being no affinity to the stationary phase.

retardation factor (R_F) the ratio of the distance travelled by a component of a sample, from the origin, to the distance travelled by the mobile phase

Retardation factor (R_F) calculations

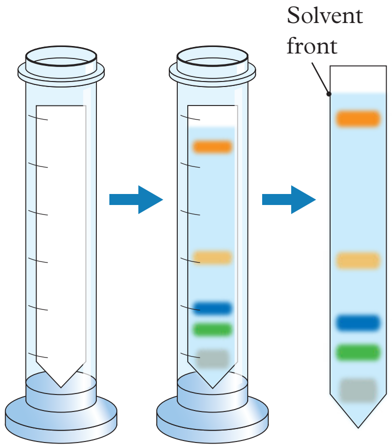
Once the separation is completed, the resulting chromatograph is analysed to identify the substances in the sample and to determine the purity of the sample. Although we can simply look at a chromatograph to judge whether two substances are identical, it is more precise to calculate the **retardation factor (R_F)** of a substance. This is the ratio of the distance moved by a substance to the distance moved by the solvent, or mobile phase.

This can be simplified to:

$$R_F = \frac{\text{distance solute moves from origin}}{\text{distance solvent moves from origin}}$$

Each atom, ion or molecule may have a different R_F depending on the properties of the mobile and stationary phases. Increasing the polarity of the mobile phase increases its affinity to charged particles within the sample. This makes the substances move further up the paper.

All R_F values must be expressed as a decimal (never a fraction) and cannot be greater than 1. If all of the calculated R_F values in a separation are 1, this indicates that the substance has not separated from the mobile phase and therefore, cannot be identified. The analysis must be run again under different mobile or stationary phase conditions.



Pigment	R_F value
Carotene	0.95
Xanthophyll	0.48
Chlorophyll a	0.30
Chlorophyll b	0.18
Leaf stain	

FIGURE 3 A mixture of plant pigments is separated into its constituent pigments. The R_F value for each of the pigments is calculated from the chromatograph.

Worked example 10.2A

Calculating R_F for a paper chromatograph

Worked example 10.2A: Watch a video that shows how to solve this problem.

A food dye was analysed against three standard dyes A–C to determine which dyes it contained. The resultant chromatograph is shown below in Figure 4.

Use the data to **determine**:

- **a** the retardation factors (R_F) of the dye standards (1 mark)
- **b** the retardation factors of all dyes within the sample (1 mark)
- **c** the dye that is most strongly attracted to the stationary phase (1 mark)
- **d** the dye that is most strongly attracted to the mobile phase (1 mark)
- **e** which dyes are in the sample (1 mark)
- **f** whether the analysis of this data is valid. (1 mark)

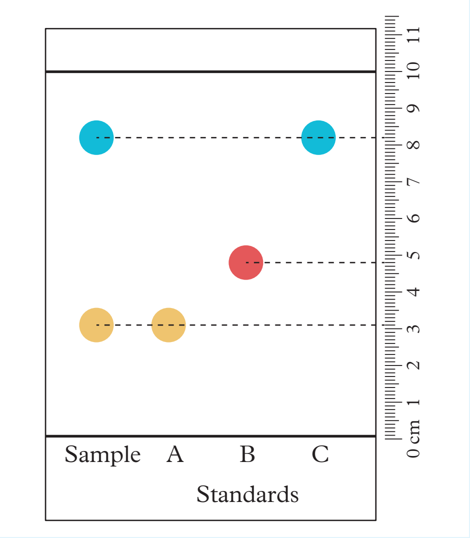


FIGURE 4 Resultant chromatograph

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to apply our understanding of retardation factors and intermolecular forces to answer the questions. The questions are worth 1 mark each, so we must recall the theory, complete the calculations and make the correct judgements.
Step 2: Use the ruler on the side of the chromatograph to read the distance travelled by the mobile phase and each standard spot.	Mobile phase = 10 cm A = 3.1 cm B = 4.8 cm C = 8.2 cm

Think	Do
Step 3: For part **a**, select the correct formula and gather any data required. Complete the calculations.	$R_F = \frac{\text{distance solute moves from origin}}{\text{distance solvent moves from origin}}$ A: $R_F = \frac{3.1\text{ cm}}{10\text{ cm}} = 0.31$ B: $R_F = \frac{4.8\text{ cm}}{10\text{ cm}} = 0.48$ C: $R_F = \frac{8.2\text{ cm}}{10\text{ cm}} = 0.82$ (1 mark)
Step 4: For part **b**, use the ruler on the side of the chromatograph to read the distance travelled by the sample spots.	3.1 cm and 8.2 cm $R_F = \frac{3.1\text{ cm}}{10\text{ cm}} = 0.31$ $R_F = \frac{8.2\text{ cm}}{10\text{ cm}} = 0.82$ (1 mark)
Step 5: For part **c**, to determine the dye that is more strongly attracted to the stationary phase, identify the standard dye with the lowest R_F value (closest to the origin).	Dye A (1 mark)
Step 6: For part **d**, to determine the dye that is more strongly attracted to the mobile phase, identify the standard dye with the highest R_F value (closest to the solvent front).	Dye C (1 mark)
Step 7: For part **e**, to determine the dyes that are in the sample, match the R_F values with those of the standards.	$R_F = 0.31$ is for dye A and $R_F = 0.82$ is for dye C. (1 mark)
Step 8: For part **f**, to determine if the analysis is valid, check that the calculated R_F values are *not* greater than or equal to 1.	All R_F values are less than 1 and greater than 0, so all of the substances have separated from the mixture and the data is valid. (1 mark)

Your turn

The ink in a Texta was analysed against three standard dyes A–C to determine the identity of its components. The resultant chromatograph is shown in Figure 5.

Use the data to **determine**:

a the retardation factors of the dye standards (1 mark)

b the retardation factors of all dyes within the sample (1 mark)

c the dye that is more strongly attracted to the stationary phase (1 mark)

d the dye that is more strongly attracted to the mobile phase (1 mark)

e which dyes are in the sample (1 mark)

f whether the analysis of this data is valid. (1 mark)

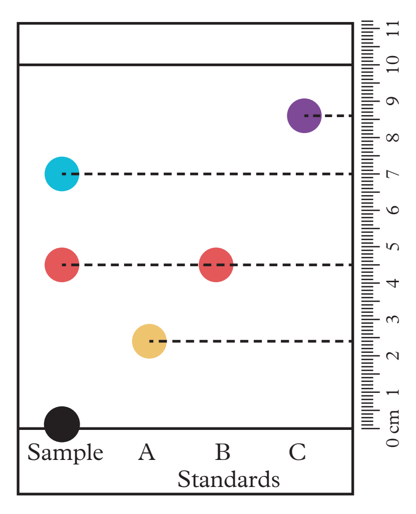


FIGURE 5 Resultant chromatograph

Real-world chemistry

How much sun are you getting?

In certain areas of the world, getting enough vitamin D is a concern due to the lack of sunlight or an indoor lifestyle. Vitamin D is fat soluble and plays a key role in the intestinal absorption of minerals such as the calcium ion (Ca^{2+}), phosphate and zinc ions (Zn^{2+}) by the body. A lack of these minerals can cause hypocalcaemia, secondary hyperparathyroidism, weakness of the bones or even rickets, where bones soften and weaken.

Humans naturally synthesise vitamin D using UVB rays from the Sun and the quantity required can be synthesised long before the skin burns. Certain regions of the world, however, do not have access to sufficient quantities of UVB from the Sun. During winter, areas near the north and south poles especially get little sunlight and therefore, little UVB.

High-performance liquid chromatography (HPLC) is widely used in numerous industries to identify and quantify chemical compounds. In HPLC, the stationary phase is less polar than the mobile phase. In reverse-phase HPLC, the polarity of the mobile and stationary phases is swapped so that the column is non-polar and the mobile phase is slightly polar.

To diagnose a patient with vitamin D deficiency, a patient's blood is collected, centrifuged to separate the blood cells from the plasma which is used for further analysis on a reverse-phase HPLC. The result is a chromatograph (Figure 6) where the components of the plasma are represented as peaks. Each has a different retention time, which represents the amount of time that it has taken to move through the column. The areas of the peaks can be calculated and used to determine the concentration of each substance.

Although highly disputed, the reference values (for normal vitamin D levels in blood) are:

- Summer: 50–300 $nmol\,L^{-1}$ (20–120 $\mu g\,L^{-1}$)
- Winter: 25–150 $nmol\,L^{-1}$ (10–60 $\mu g\,L^{-1}$)

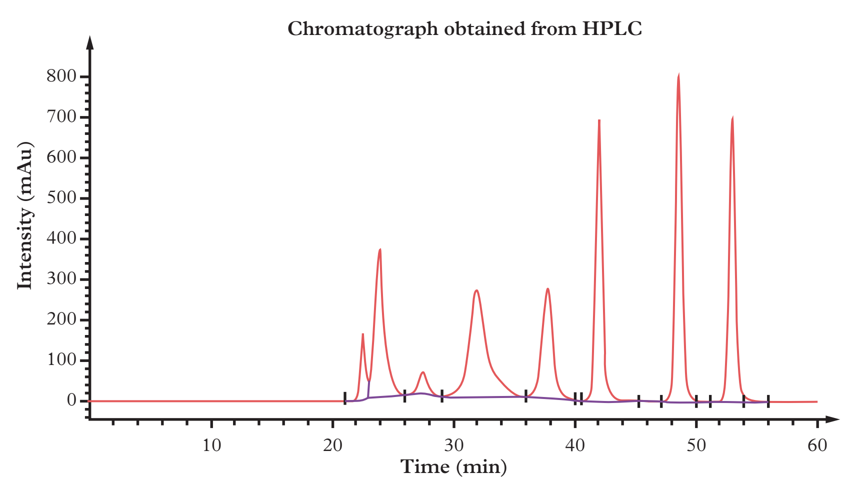


FIGURE 6 Example chromatograph from a HPLC separation

Apply your understanding

1 **Infer** why a non-polar stationary phase would be used to analyse vitamin D. (2 marks)

2 **Propose** a reason why the mobile phase is only slightly polar and not completely polar. (2 marks)

3 **Use** your knowledge of the identification of unknown substances in paper chromatography to **explain** what a chemist would need to do to confirm the identity of vitamin D in the blood sample. (3 marks)

4 **Infer** why there is a difference between the summer and winter concentrations of vitamin D. (2 marks)

What is thin layer chromatography?

thin layer chromatography (TLC)
an analytical technique for separating and identifying mixtures; the stationary phase is typically a thin layer of silica gel, aluminium oxide or cellulose supported on a piece of glass or plastic

Study tip

Both paper and thin layer chromatography are named after the stationary phase. This may help with recalling specific information about these techniques.

Study tip

If R_F = 1.0, then no separation from the stationary phase has occurred. A typical exam question may ask how to separate this component in a further analysis. Look at the properties of the stationary phase and determine what you could do to make this substance more attracted to it.

Thin layer chromatography (TLC) works by the same principles as paper chromatography with one major change that makes the process more efficient (faster and more sensitive). In TLC, the support for the stationary phase is a piece of glass or plastic, which is coated in the stationary phase, consisting of silica gel, aluminium oxide or cellulose. This coating is a thin layer spread on the surface of the plastic or glass; hence, the name "thin layer" chromatography.

The components within the sample are separated, as with paper chromatography, according to their affinity for the mobile or stationary phase.

It is important to note that paper chromatography and TLC are not limited to coloured compounds. Fluorescent TLC plates can be used under UV light to see the components of a sample that would not otherwise be visible. These are shown in the darker areas of the plate where the sample blocks the fluorescence of the plate (Figure 7).

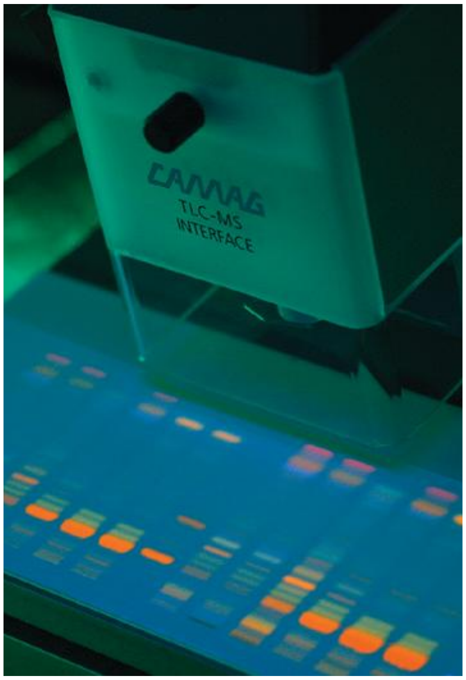


FIGURE 7 TLC performed on fluorescent plates

Two-dimensional paper or thin layer chromatography

Occasionally, a chromatography technique will not separate some components sufficiently within the sample. This is because the components interact the same amount with the mobile and stationary phases. Rather than repeating the analysis, chemists will rotate the chromatograph by 90°, so that the sample is on the bottom. They rule a new origin line and run the analysis again, using a different mobile phase with different properties (Figure 8). The chromatograph is analysed by using R_F calculations for analysis 1 and also for analysis 2. This is called two-dimensional chromatography.

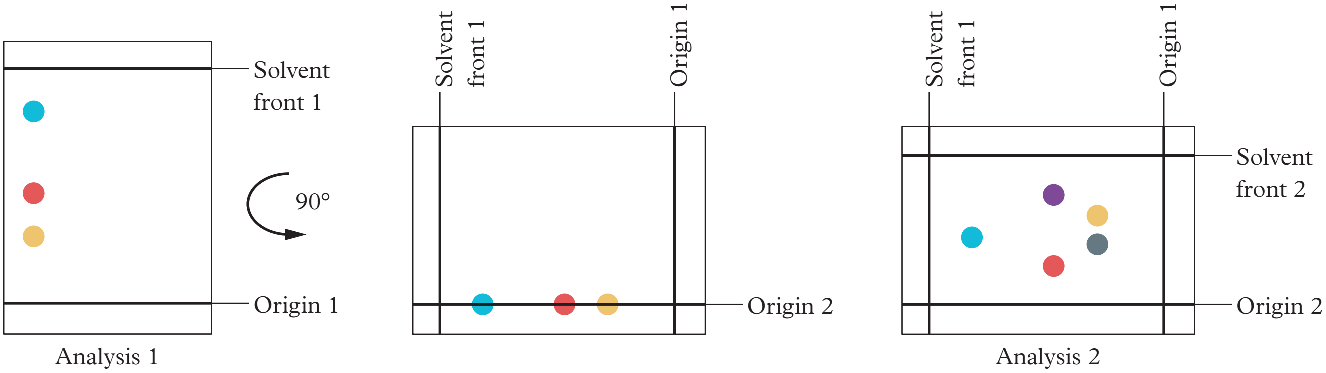


FIGURE 8 In two-dimensional paper chromatography, the chromatograph is rotated 90° after the first analysis. This gives a new origin line. The second analysis separates the green and purple dots further into their individual components. (Origin 1 and solvent front 1 refer to the first analysis. Origin 2 and solvent front 2 refer to the second analysis.)

Check your learning 10.2

Check your learning 10.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the steps involved in calculating retardation factor, using Figure 9. (3 marks)

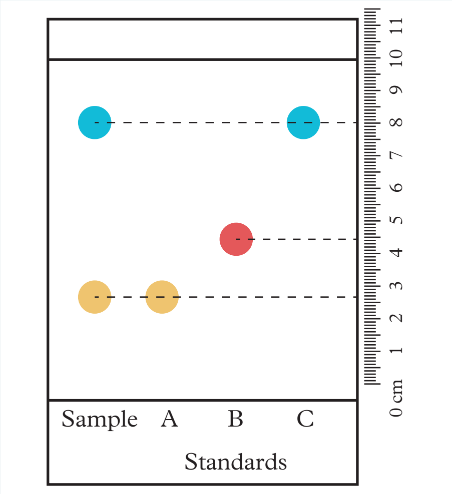


FIGURE 9 Chromatograph

2 **Explain** how components of a mixture are identified by paper or thin layer chromatography. (4 marks)

Analytical processes

Use the chromatograph in Figure 10 to answer questions 3 and 4.

3 **Analyse** the chromatograph.

- a **Calculate** the R_F values for all components. (3 marks)
- b **Determine** which component has the highest affinity for the mobile phase. (1 mark)
- c **Determine** which component has the highest affinity for the stationary phase. (1 mark)

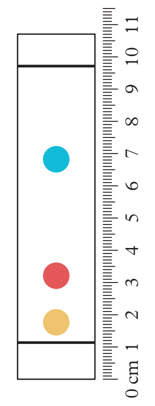


FIGURE 10 Chromatograph

4 If the chromatograph were run with a polar stationary phase, **compare** the types of intermolecular forces that may be present between each component and the stationary and mobile phases. (2 marks)

5 The R_F values of 10 food dyes are shown in Table 1. These R_F values are specific for TLC using a silica plate and a 1% aqueous ethanol mobile phase.

TABLE 1 R_F values of some food dyes in 1% ethanol

Dye	R_F
Brilliant blue FCF	0.12
Indigotine	0.23
Fast green FCF	0.28
Erythrosine	0.45
Quinoline Yellow	0.50
Carmoisine	0.60
Tartazine	0.65
Green S	0.72
Patent blue V	0.86
Ponceau 4R	0.91

Three foods A–C were tested to determine which food dyes were present. The resultant chromatograph is shown in Figure 11.

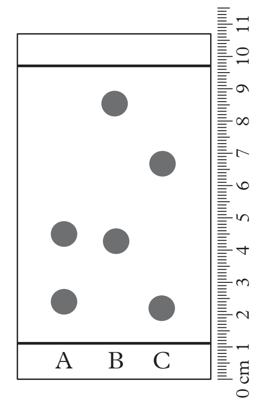


FIGURE 11 Chromatograph

- a **Calculate** the R_F values of the components of each sample and **determine** which food colourings were present in the foods. (6 marks)
- b There are two components of these samples that, on visual inspection, appear to be the same. **Consider** whether these components are the same. **Explain** your answer. (2 marks)

6 A note left at a crime scene was written in black pen. Two black pens found at the scene were analysed by forensic scientists to determine which pen had written the note. To ensure valid results, five teams were given one of the two pens to test. Each team developed a method that involved using paper chromatography with a water mobile phase. The data in Table 2 was obtained.

a **Use** the data to **calculate** the R_F values of components 1–3. (1 mark)

b **Determine** which teams analysed the same pens. (1 mark)

c **Determine** which teams analysed the pen that was responsible for the note at the crime scene. (1 mark)

d **Deduce** whether component 1 is the same substance for both pens. **Explain** your answer. (2 marks)

e **Comment** on whether it is possible to identify the components of the sample from this analysis. (2 marks)

TABLE 2 Distance travelled by different components in paper chromatography

Team	Solvent front (cm)	Component 1 (cm)	Component 2 (cm)	Component 3 (cm)
1	11	6.16	8.69	2.31
2	15	4.80	8.85	13.05
3	9	5.04	7.11	1.89
4	21	11.76	16.59	4.41
5	18	5.76	10.62	15.66
Crime scene note	25	8.00	14.75	21.75

Practical

Lesson 10.3

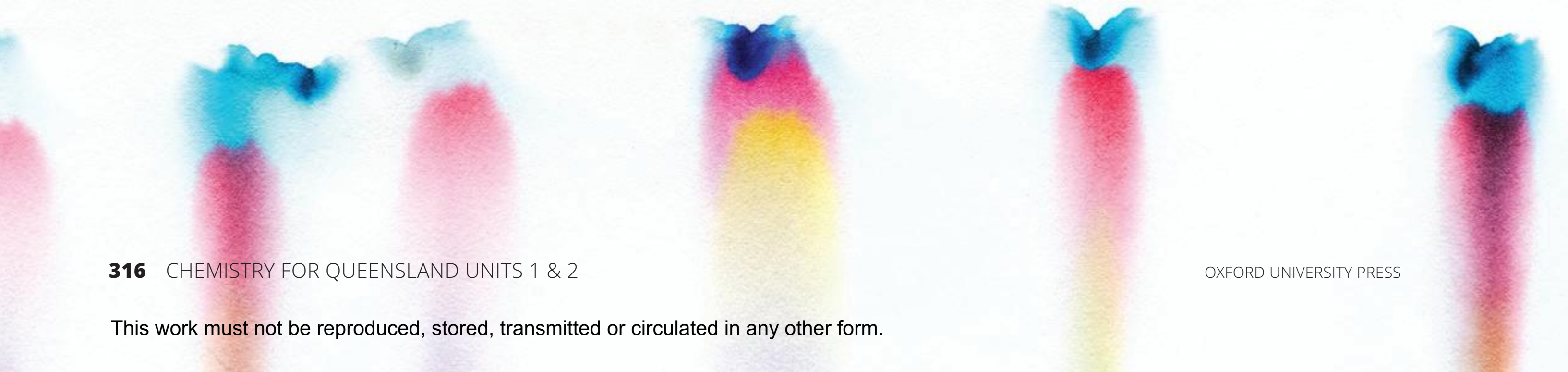
Separating components of a mixture using paper chromatography

Learning intentions and success criteria

Video demonstration

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

FIGURE 1 Paper chromatography separates mixtures.



MODULE 10

Lesson 10.4

Review: Chromatography techniques

Summary

10.1
- Chromatography separates components in a mixture based on the intermolecular interactions between atoms, molecules or ions in the mobile and stationary phases.
- The adsorption and desorption of components of a mixture depends on the strengths of their intermolecular forces.

10.2
- Paper and thin layer chromatography are the simplest chromatographic techniques for separating a sample. Both techniques are qualitative – they can be used to determine the purity and/or composition of substances.
- Chromatographs are analysed by calculating R_F.

10.3
- Practical: Separating components of a mixture using paper chromatography

Key formulas

Retardation factor	$R_F = \frac{\text{distance solute moves from origin}}{\text{distance solvent moves from origin}}$

Review questions 10.4A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 A polar solute will have a high affinity for a

- **A** polar mobile phase but not a polar stationary phase.
- **B** polar mobile phase and a polar stationary phase.
- **C** non-polar mobile phase and a polar stationary phase.
- **D** polar mobile phase and a non-polar mobile phase.

2 R_F can be calculated by

- **A** adding the distance of the solvent front to the distance moved by the solvent.
- **B** determining the difference between the distance of the solvent front and the distance moved by the solute.
- **C** dividing the distance moved by the solute by the distance of the solvent front.
- **D** determining the distance moved by the solute.

3 In a mixture of solutes, the following R_F values were calculated based on a polar mobile phase and a non-polar stationary phase. Use the results in Table 1 to choose the correct answer.

TABLE 1 R_F values for solutes A, B and C

Solute	R_F
A	0.50
B	0.72
C	0.92

- **A** A is the least polar and desorbs the most from the stationary phase.
- **B** C is the most polar and has the strongest affinity for the mobile phase.
- **C** A is the most polar and desorbs the least from the stationary phase.
- **D** C is the least polar and has the strongest affinity for the mobile phase.

4 In a TLC separation to determine the amino acids in a seaweed extract, a student determines that two amino acids have the same R_F. To fully separate these molecules, the best option would be to

A turn the separation sideways and change the properties of the mobile phase for two-dimensional chromatography.

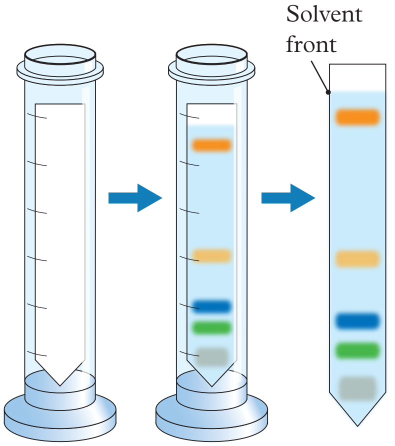
B run the separation again with a more polar mobile phase.

C run the separation again with a more polar stationary phase.

D run the separation again with paper chromatography.

Use the following information to answer questions 5 to 7.

The separation of a leaf stain into its individual pigments is shown in Figure 1. The R_F value is calculated from the chromatograph.



Pigment	**R_F value**
Carotene	0.95
Xanthophyll	0.48
Chlorophyll a	0.30
Chlorophyll b	0.18
Leaf stain	

FIGURE 1 Chromatograph from separation of a leaf stain

5 Which solute has the greatest adsorption to the stationary phase?

A Chlorophyll b (green)

B Chlorophyll a (blue)

C Xanthophyll (yellow)

D Carotene (orange)

6 Which solute has the strongest affinity for the mobile phase?

A Chlorophyll b (green)

B Chlorophyll a (blue)

C Xanthophyll (yellow)

D Carotene (orange)

7 Assuming that the mobile phase is polar, which solute has the strongest intermolecular forces?

A Chlorophyll b (green)

B Chlorophyll a (blue)

C Xanthophyll (yellow)

D Carotene (orange)

Review questions 10.4B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

8 **Explain** how chromatography separates substances from a mixture. **Use** the terms “mobile phase”, “stationary phase”, “affinity”, “adsorb” and “desorb”. (7 marks)

9 **Describe** how chromatography can be used to identify impurities in a substance. (2 marks)

10 **Explain** how non-polar molecules are able to interact if they have no overall charges. (1 mark)

11 **Explain** why it is necessary to prepare a set of standards when performing qualitative analysis in chromatography. (1 mark)

12 **Explain** why an R_F value can never be 0 or 1 in a valid separation. (2 marks)

Analytical processes

13 A separation was conducted to determine the pigments in a sample of food colouring. The chromatograph included the standards of food dye that the manufacturer claims is in the food product. The chromatograph of the pigment and standard dyes A–C are shown in Figure 2.

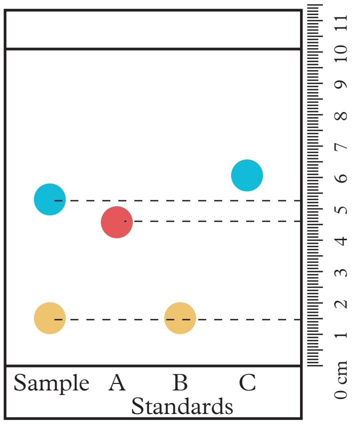


FIGURE 2 Chromatograph

- **a** **Calculate** the R_F values for the solutes in the sample. (1 mark)
- **b** **Identify** the standard dye with the highest affinity for the polar mobile phase. (1 mark)
- **c** **Identify** the standard dye which desorbs the least from the stationary phase. (1 mark)
- **d** **Determine** which standard dyes are in the food sample. (1 mark)
- **e** **Compare** the results obtained to the manufacturer's claims regarding the dyes in the food. (2 marks)

14 The chromatograph shows a separation conducted to determine the types of ink in a sample of marker pen. The stationary phase is paper and the mobile phase is 10% by volume of ethanol in water.

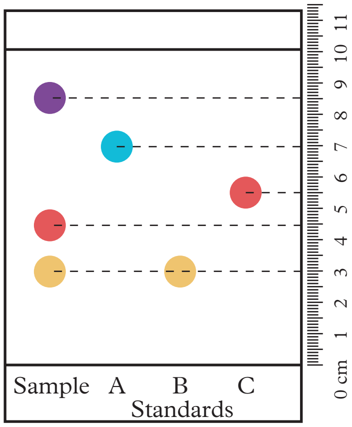


FIGURE 3 Chromatograph

- **a** **Calculate** the R_F values for the solutes in the sample. (1 mark)
- **b** **Identify** the polarity of mobile and stationary phases. **Explain** your choice. (2 marks)
- **c** **Identify** the standard dye with the strongest solute–solvent intermolecular forces. (1 mark)
- **d** **Identify** the standard dye that adsorbs the strongest to the stationary phase. (1 mark)
- **e** **Determine** which standard dyes are in the ink sample. (1 mark)
- **f** **Describe** how the separation can be improved to ensure that the identity of all dyes is determined. (2 marks)

15 **Consider** the chromatograph in Figure 4.

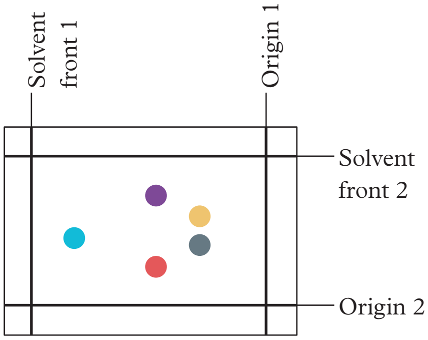


FIGURE 4 Chromatograph

- **a** **Describe** briefly how it would have been produced. (2 marks)
- **b** **Contrast** the mobile phase (solvent) in both separations. (2 marks)
- **c** **Explain** why the chromatography technique was performed twice. (1 mark)
- **d** Using two-dimensional chromatography, you cannot compare the dyes in the mixture directly with known dyes. Instead, you must make use of known R_F values.
  - **i** **Explain** why this is true. (3 marks)
  - **ii** **Explain** how an R_F value is measured. (1 mark)
- **e** **Use** a ruler to **calculate** both R_F values for all five chemical substances. (2 marks)
- **f** From the R_F values calculated, **determine** which spot has the:
  - **i** highest R_F value in solvent 1 (1 mark)
  - **ii** highest R_F value in solvent 2 (1 mark)
  - **iii** lowest R_F value in solvent 1 (1 mark)
  - **iv** lowest R_F value in solvent 2 (1 mark)
  - **v** highest affinity for solvent 1 (1 mark)
  - **vi** highest affinity for solvent 2 (1 mark)
  - **vii** highest affinity to the stationary phase. (1 mark)

16 A mixture is separated by paper chromatography. The mobile phase is a 1% aqueous ethanol solution. After the mixture is separated, the components of the sample have R_F values of 0.89, 0.62 and 0.27. **Determine**:

- **a** which component has the highest affiliation for the mobile phase (1 mark)
- **b** which component has the highest affinity for the stationary phase (1 mark)
- **c** what can be concluded about the properties of the components in the mixture. (2 marks)

17 Many mobile phases consist primarily of water. **Consider** why water is often chosen as a mobile phase. (2 marks)

Knowledge utilisation

18 The R_F values in Table 2 were calculated from a sample that has been separated using paper chromatography.

TABLE 2 Results from paper chromatography

Component	R_F	Properties
A	0.83	Most polar
B	0.74	
C	0.54	
D	0.32	

- **a** **Calculate** the distance that each solute moved from the origin, assuming that the solvent front travelled 17 cm. (1 mark)
- **b** **Determine** the polarity of the mobile and stationary phases. **Justify** your response. (4 marks)
- **c** **Identify** the component with the highest affinity for the mobile phase. (1 mark)
- **d** **Identify** the component with the strongest adsorption to the stationary phase. (1 mark)

Data drill

Paper chromatography

A chemist runs a mixture of dyes and their standards on a piece of chromatography paper. The chemist uses the following conditions for the separation:

- The mobile phase is more polar.
- The stationary phase is slightly polar.

TABLE 1 Results of the paper chromatography experiment

Component	R_F	Colour
A	0.2	Green
B+C	0.5	Orange
D	1.0	Blue

Apply understanding

1 **Use** the R_F values in Table 1 to **identify** the component that is the least polar. (1 mark)

2 **Determine** the types of intermolecular forces experienced between molecules of the mobile phase. (1 mark)

Analyse data

3 **Sequence** the components in order of increasing polarity. (1 mark)

4 **Compare** the intermolecular forces experienced by components A and D. (2 marks)

Evaluate evidence

5 With reference to the data, **justify** whether the following assumption is valid: *The orange component must be made of two dyes, as the packaging states that the mixture contain red and yellow.* (2 marks)

6 **Deduce** whether this method is a valid separation for this dye. **Justify** your response. (2 marks)

Module 10 checklist: Chromatography techniques

Quizlet: Revise key terms online to test your understanding

MODULE 11

Gases

Introduction

The earliest philosopher in ancient Greece proposed five elements that make up our world: earth, water, fire, aether and air. It was not until the 17th century that scientists developed an alternative, evidence-based theory of gases, which we use today.

Gases are one of the five states of matter, and the first to be thoroughly investigated by early scientists. Early experiments focused on the relationships between the pressure, volume and temperature of gases. The result of these experiments was the formulation of the gas laws, the earliest of which was developed in 1662 by Robert Boyle and is still used in modern chemistry.

A thorough understanding of kinetic molecular theory is essential to understanding the behaviour of gases. This theory describes the way that gas particles move, which gives an insight into their properties. Kinetic molecular theory is used to explain the early research of scientists, such as Boyle, and the relationship between pressure, volume and temperature.

Gas laws can be applied to various phenomena and occupations. Scuba divers trust the gas laws to help them to avoid decompression sickness when underwater. Similarly, aeroplane travel relies on a knowledge of gases in the atmosphere.

This module focuses on the understanding of kinetic molecular theory, and its ability to explain the gas laws and their relationship with the number of particles contained within a gaseous system.

Prior knowledge

Prior knowledge quiz

Check your understanding of the mole and kinetic theory before you start.

Subject matter

Science understanding

→ State the relationship between the volume of a gas, number of moles and molar volume at standard temperature and pressure (STP).

→ Apply the kinetic theory of gases to explain the relationships between pressure, temperature and volume of a gas.

→ Identify that the kinetic theory of gases applies to ideal gases.

- → Apply the ideal gas equation to calculate the mass of chemicals or the volume of a gas (STP) involved in a chemical reaction. (Formula: $PV = nRT$)
- → Analyse data to determine the relationships between pressure, temperature and volume of a gas.

Science as a human endeavour

- → Appreciate that safe scuba diving requires knowledge of the behaviour of gases.
- → Explore Jacques Cousteau and Emile Gagnan's role in the invention of SCUBA.

Science inquiry

- → Investigate Boyle's law or the molar volume of a gas.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

These lessons are available on Oxford Digital.

Lesson 11.2	Investigating the properties of gases
Lesson 11.4	Investigating Boyle's law

Lesson 11.1

Properties of gases

Key ideas

→ Kinetic molecular theory can be used to explain gas behaviour and properties.

→ Pressure, temperature and volume affect the behaviour of gases.

What is the kinetic theory of gases?

Learning intentions and success criteria

kinetic molecular theory the theory that states that all particles are in constant random motion

Kinetic molecular theory explains the movement and behaviour of particles in solids, liquids and gases. The more kinetic energy they have, the more they will move. This can also be used to explain the behaviour of gases. In the gaseous state, particles do not remain together. They have a high level of kinetic energy so they separate and move freely.

The kinetic molecular theory of gases can be summarised into the following key points.

- Gas particles are in constant, random, straight-line motion.
- Gas particles have very little attraction to one another because the energy of their movement is greater than the energy of attraction between particles.
- The space between gas particles is much larger than the space that the particles themselves occupy. This means that the gas can be compressed and has a low density.
- Kinetic energy is conserved in gas particle collisions. This means that when gas particles collide, the combined energy of the two particles is the same after the collision as it was before.
- The average kinetic energy of a gaseous system depends on its temperature. The higher the temperature, the higher the average kinetic energy.

  The behaviour of gases is affected by:
  - pressure
  - temperature
  - volume.

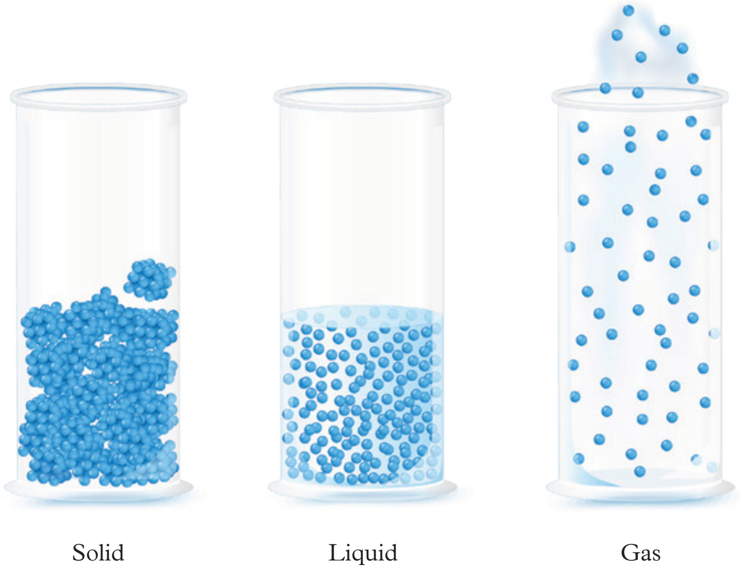


FIGURE 1 Gases expand to occupy the whole space of a container, unlike solids and liquids.

How does pressure affect gas behaviour?

pressure
the force exerted, per unit area, by one substance upon another substance

Pressure is defined as the force exerted, per unit area, by one substance on another substance. For gases, this is the pressure exerted by gas particles when they collide with the walls of a container. This relationship can be represented mathematically as:

$$\text{pressure} = \frac{\text{force}}{\text{area}} \text{ or } P = \frac{F}{A}$$

A good analogy to help you visualise pressure involves a punching bag. If a punching bag is hit only once, a small force is exerted, and the bag may move a little. If a punching bag is hit continually, with many hits, a larger force is exerted.

Gas particles that have more kinetic energy will move more rapidly, greatly increasing the average force of each collision. There is also a slight increase in the number of collisions. An increase in the energy of the collisions with the walls of the container will increase the pressure exerted on the container.

FIGURE 2 Pressure is generated when air molecules collide with the internal walls of a balloon.

The SI unit for pressure is the pascal (Pa). This is the force of one newton per square metre of substance ($N\,m^{-2}$). Often, we deal with pressure in the magnitude of kilopascals (kPa). In your studies, you will also encounter the unit of pressure called bar, where 1 bar is equal to 100 kPa.

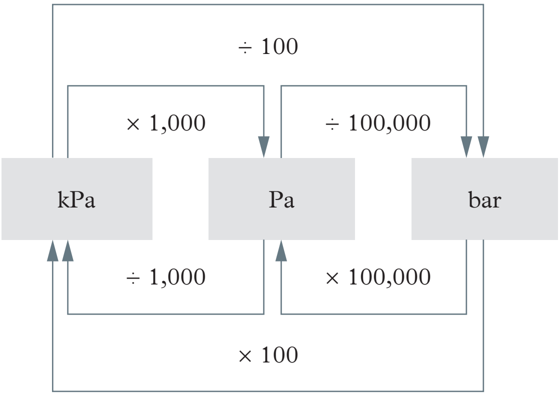


FIGURE 3 The conversion tool for converting between units of pressure

Worked example 11.1A

Converting between pressure units

Worked example 11.1A: Watch a video that shows how to solve this problem.

Use the correct conversion factors to convert the following units.

a 0.872 bar to kPa (1 mark)

b 635 kPa to bar (1 mark)

c 92,572 kPa to bar. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Use" means to apply knowledge or rules to put theory into practice. We need to apply our understanding of pressure unit conversions to express each value in the correct units. Each question is worth 1 mark, so we must recall the formulas and apply them to the values.

Think	Do
Step 2: Recall the conversion formulas and complete the calculations. See Figure 3.	**a** $0.872\,\text{bar} \times \frac{100\,\text{kPa}}{1\,\text{bar}} = 87.2\,\text{kPa}$ **b** $635\,\text{kPa} \times \frac{1\,\text{bar}}{100\,\text{kPa}} = 6.35\,\text{bar}$ **c** $92{,}572\,\text{Pa} \times \frac{1\,\text{bar}}{100{,}000\,\text{Pa}} = 0.92572\,\text{bar}$
Step 3: Finalise your answers and make sure you have expressed them using the correct units and number of significant figures.	**a** 87.2 kPa (1 mark) **b** 6.35 bar (1 mark) **c** 0.92572 bar (1 mark)

Your turn

Use the correct conversion factors to convert the following units.

a 1.39 bar to kPa (1 mark)

b 10,600 kPa to bar (1 mark)

c 0.562 bar to Pa. (1 mark)

How does temperature affect gas behaviour?

All particles in a container have different amounts of heat energy. Temperature is a measure of the average heat energy of the particles within a system. As the temperature of the system increases, the average heat energy of the system also increases. This can be represented by a Maxwell–Boltzmann distribution curve (Figure 4). Particles with more heat energy will move more in a container. Therefore, they have an increased amount of kinetic energy.

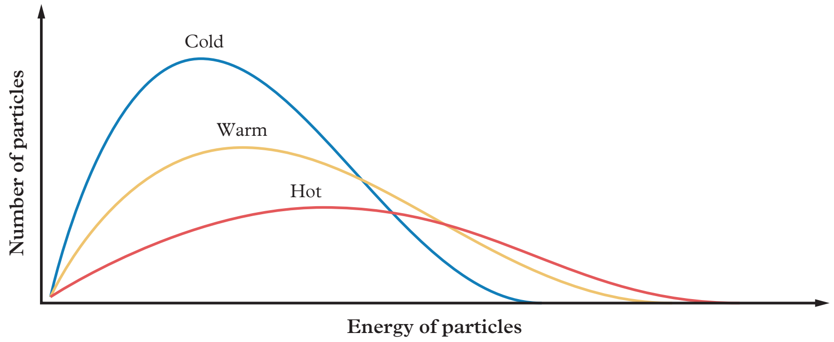


FIGURE 4 The Maxwell–Boltzmann representation of the amount of energy distributed in a reaction system at different temperatures; the curves represent the energy distributions at different temperatures.

The kinetic energy of any gaseous system is the same at a specified temperature. For kinetic energy to remain constant, the average velocity of particles must decrease when the mass of the particles increases. This means that heavier particles move slower and lighter ones move faster on average, at the same temperature.

The SI unit for temperature is the kelvin (K). One degree on the Celsius scale is equivalent to one kelvin. So:

$$0°\text{C} = 273\,\text{K or } 0\,\text{K} = -273\,°\text{C}$$

$$25°\text{C} = 273\,\text{K} + 25 = 298\,\text{K}$$

Figure 5 demonstrates how to convert between temperature units.

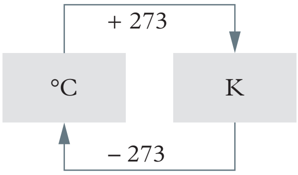


FIGURE 5 The conversion tool for converting between units of temperature

Worked example 11.1B

Converting between temperature units

Worked example 11.1B: Watch a video that shows how to solve this problem.

Use the correct conversion factor to convert the following units.

a 100°C to K (1 mark)
b 45°C to K (1 mark)
c –124°C to K (1 mark)
d 345 K to °C (1 mark)
e 123 K to °C (1 mark)
f 526 K to °C. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Use" means to apply knowledge or rules to put theory into practice. We need to apply our understanding of temperature unit conversions to express each value in the correct units. Each question is worth 1 mark, so we must recall the formulas and apply them to the values.
Step 2: Recall the conversion formula and complete the calculations. See Figure 5.	**a** $100 + 273 = 373$ **b** $45 + 273 = 318$ **c** $-124 + 273 = 149$ **d** $345 - 273 = 72$ **e** $123 - 273 = -150$ **f** $526 - 273 = 253$
Step 3: Finalise your answers and make sure you have expressed them using the correct units and number of significant figures.	**a** 373 K (1 mark) **b** 318 K (1 mark) **c** 149 K (1 mark) **d** 72°C (1 mark) **e** –150°C (1 mark) **f** 253°C (1 mark)

Your turn

Use the correct conversion factor to convert the following units.

a 18°C to K (1 mark)
b 45°C to K (1 mark)
c –45°C to K (1 mark)
d 596 K to °C (1 mark)
e 0 K to °C (1 mark)
f 124 K to °C. (1 mark)

How does volume affect gas behaviour?

volume
a measure of the space occupied by a substance

Gases occupy the whole container. Therefore, the **volume** of a container is the equivalent of the volume of the gas inside that container. Two types of gaseous systems are considered when discussing gas behaviour:

- **Fixed volume or rigid wall systems** are usually contained within gas cylinders because their walls will not expand or collapse. This means that the volume that the gas occupies within the cylinder is always constant.

FIGURE 6 A balloon is an example of a variable volume system because it can adjust to the volume of gas.

- **Variable volume systems** have walls that can adjust to the volume of the gas such as a balloon or a syringe. As more gas is added into the balloon, the elastic of the balloon stretches to accommodate the gas (Figure 6). When the gas leaves the balloon, the elastic compresses to a smaller size.

The SI unit for volume is cubic metres (m^3). Litres (L) is also commonly used, where:

$$1\,m^3 = 1{,}000\,L = 1{,}000{,}000\,mL$$

$$0.001\,m^3 = 1\,L = 1{,}000\,mL$$

Figure 7 demonstrates how to convert between volume units.

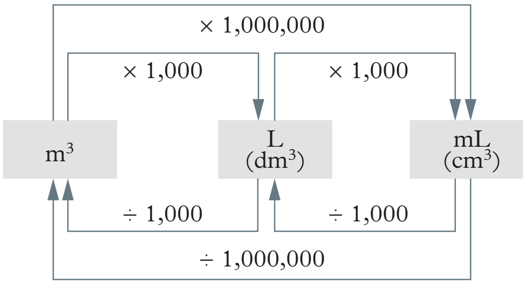


FIGURE 7 The conversion tool for converting between units of volume

Study tip

1 m³ is the volume of a cube with sides measuring 1 m each. 1 L is the volume of a cube with sides measuring 10 cm each. Try and visualise these volumes in your head.

Worked example 11.1C

Converting between volume units

Worked example 11.1C: Watch a video that shows how to solve this problem.

Use the correct conversion factors to convert the following units:

a 25 mL to m^3 (1 mark)

b 45.672 L to m^3 (1 mark)

c 2.5 m^3 to L. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Use” means to apply knowledge or rules to put theory into practice. We need to apply our understanding of volume unit conversions to express each value in the correct units. Each question is worth 1 mark, so we must recall the formulas and apply them to the values.
Step 2: Recall the conversion formulas and complete the calculations. See Figure 7.	**a** $25 \times \frac{1\,m^3}{1{,}000{,}000\,mL} = 2.5 \times 10^{-5}\,m^3$ **b** $45.672 \times \frac{1\,m^3}{1{,}000\,L} = 0.045672\,m^3$ **c** $2.5 \times \frac{1{,}000\,L}{1\,m^3} = 2500\,L$
Step 3: Finalise your answers and make sure you have expressed them using the correct units and number of significant figures.	**a** $2.5 \times 10^{-5}\,m^3$ (1 mark) **b** $0.045672\,m^3$ (1 mark) **c** 2500 L (1 mark)

Your turn

Use the correct conversion factors to convert the following units:

a 0.00475 m^3 to mL (1 mark)

b 1.5 L to m^3 (1 mark)

c 0.5 m^3 to L. (1 mark)

Check your learning 11.1

Check your learning 11.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** kinetic energy and **explain** why it has an essential role in the particle model of matter. (2 marks)

2 **Define** pressure in terms of gases. (1 mark)

3 **Use** kinetic molecular theory to explain why:
  - **a** there are spaces between gas particles (2 marks)
  - **b** gas particles have more energy than particles in a solid or liquid (2 marks)
  - **c** gases occupy the space of a container. (2 marks)

4 **Use** the correct conversion factors to convert the following units.
  - **a** 315 K to °C (1 mark)
  - **b** 74°C to K (1 mark)
  - **c** 273 bar to kPa (1 mark)
  - **d** 50 kPa to bar (1 mark)
  - **e** 25 mL to m^3 (1 mark)
  - **f** 0.45 m^3 to L. (1 mark)

5 **Use** kinetic molecular theory to **explain** what happens to the kinetic energy of a gaseous system when:
  - **a** temperature is decreased (2 marks)
  - **b** pressure is increased (2 marks)
  - **c** volume is kept the same. (2 marks)

6 The kinetic energy of a gas with a mass of 2.50 g is calculated to be 625.0 J. **Calculate** the average velocity of the gas particles. (2 marks)

Knowledge utilisation

7 At 100°C, some particles still have the same amount of energy as they did at 0°C. **Discuss** how this is possible, with reference to the Maxwell–Boltzmann distribution. (3 marks)

8 When a scented candle is lit at one end of a room, the scent slowly diffuses to the opposite side of the room. **Determine** how the kinetic molecular theory explains this phenomenon. (3 marks)

Learning intentions and success criteria

Video demonstration

Practical

Lesson 11.2

Investigating the properties of gases

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 11.3
Gas laws

Key ideas

→ As temperature increases, pressure increases (Gay-Lussac's law); as temperature increases, the volume of a variable volume system increases (Charles' law); as volume increases, pressure decreases (Boyle's law).

→ Gases occupy the same molar volume under standard temperature and pressure (STP).

What are the gas laws?

The relationship between pressure, volume and temperature is summarised into three main **gas laws** that are named after the scientists who developed them. They are:

- Gay-Lussac's law, which relates temperature and pressure
- Charles' law, which relates temperature and volume
- Boyle's law, which relates pressure and volume.

Learning intentions and success criteria

gas laws
the laws that explain the properties of gases and how they behave at different temperatures, volumes and pressures

Gay-Lussac's law: how are temperature and pressure related?

Gay-Lussac's law reflects the relationship between the temperature and pressure of a gas. As temperature increases, the kinetic energy of particles within a fixed volume system also increases. This means that particles move faster and collide with the walls of the container with more force and slightly more frequently. As the walls of the container cannot move, and the container cannot expand to accommodate these excess collisions, the pressure within the container increases. This occurs when volume is constant – that is, a fixed volume system.

Study tip

Gas laws use kelvin (K) as a measurement of temperature. You need to know how to convert degrees Celsius (°C) to kelvin. Use the formula K = °C + 273.

Numerical representation of Gay-Lussac's law

The relationship between temperature and pressure can be represented numerically. Figure 1 depicts a round-bottom flask filled with a gas and attached to a barometer (used to measure pressure). The same flask is placed in three water baths at different temperatures to determine the effect of temperature on gas pressure. Table 1 summarises the data gathered in the experiment.

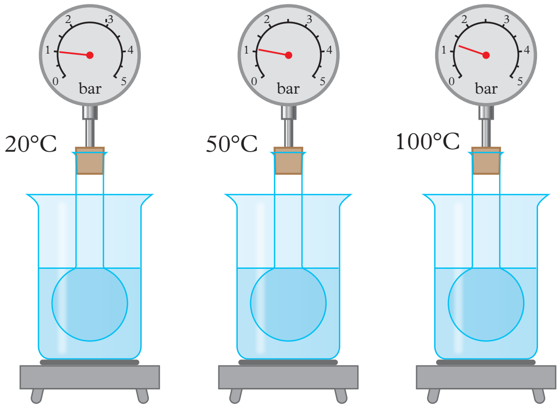


FIGURE 1 The effect of temperature on pressure in a gaseous system

TABLE 1 Data obtained from the experiment in Figure 1

Temperature (°C)	Temperature (K)	Pressure (bar)	P/T (bar °C^{-1})	P/T (bar K^{-1})
20	293	1.00	0.050	0.0034
50	323	1.10	0.022	0.0034
100	373	1.27	0.013	0.0034

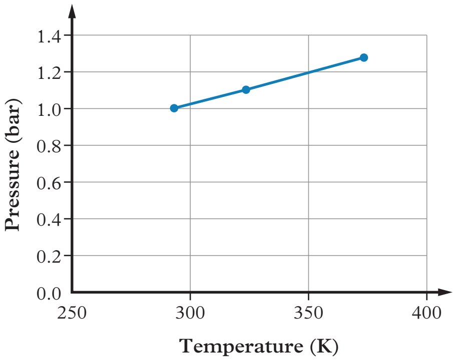


FIGURE 2 The relationship between temperature (K) and pressure in a gaseous system

The relationship between temperature and pressure is shown in the graph in Figure 2. The graph demonstrates that there is a linear relationship between temperature (in K) and pressure. Therefore, pressure is proportional to temperature in Kelvin ($P \propto T$). This relationship can therefore be represented mathematically as:

$$\frac{P}{T} = \text{a constant}$$

If a change in temperature or pressure to a gaseous system is made, new values can be calculated from the equation, as long as temperatures are in kelvin:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Worked example 11.3A

Calculating pressure and temperature changes

Worked example 11.3A: Watch a video that shows how to solve this problem.

At 293 K, the pressure in a reaction vessel is 101.3 kPa. If the temperature is raised to 308 K, **calculate** the final pressure of the vessel. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the final pressure (P_2). The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ $P_1 = 101.3\,\text{kPa}$, $T_1 = 293\,\text{K}$ $P_2 = ?$, $T_2 = 308\,\text{K}$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$\frac{101.3\,\text{kPa}}{293\,\text{K}} = \frac{P_2}{308\,\text{K}}$ (1 mark) $P_2 = 101.3\,\text{kPa} \times \frac{308\,\text{K}}{293\,\text{K}}$ $P_2 = 106.49\,\text{kPa}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	106 kPa (1 mark)

Your turn

At 298 K, the pressure in a reaction vessel is 4.0 bar. If the pressure decreases to 2.5 bar, **calculate** the final temperature of the vessel. (2 marks)

Charles' law: how are temperature and volume related?

Charles' law states that at a constant pressure, the volume of a gas is proportional to the temperature of the system. This means that when the volume of a container is variable, temperature will have an effect. As temperature increases, the kinetic energy (movement) of particles increases and they collide with the walls of the container more frequently and with more force. However, as the volume of the container can vary, these collisions will push on the walls of the container, causing the container to expand, rather than increasing the pressure. In this system, pressure is constant, i.e. the container expands, and in so doing, maintains the pressure of the system.

A balloon is one example of this phenomenon. When inflated and then placed in a beaker filled with liquid nitrogen (–196°C), the low temperature causes the volume of the balloon to decrease, seen in Figure 3.

FIGURE 3 A balloon decreases in volume when placed in a beaker filled with liquid nitrogen (–196°C).

When the balloon is removed from the beaker and the temperature starts to increase, the air inside the balloon expands back to its original volume (Figure 4).

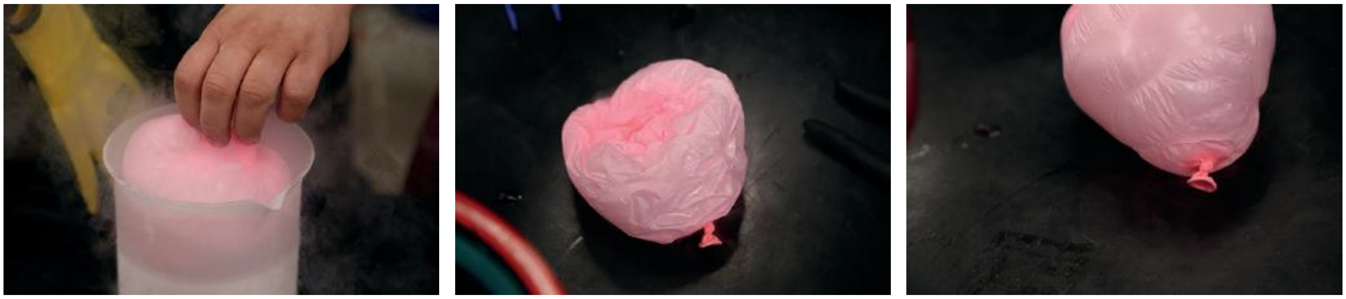


FIGURE 4 A balloon increases in volume when removed from the beaker and placed on a bench at room temperature.

Numerical representation of Charles' law

The relationship between temperature and volume is represented in Figure 5. A cylinder filled with a gas is kept at a constant pressure of 1 bar. The cylinder is heated, and the change in volume is recorded to determine the effect of temperature on the volume of a gaseous system. Table 2 summarises the data gathered in the experiment.

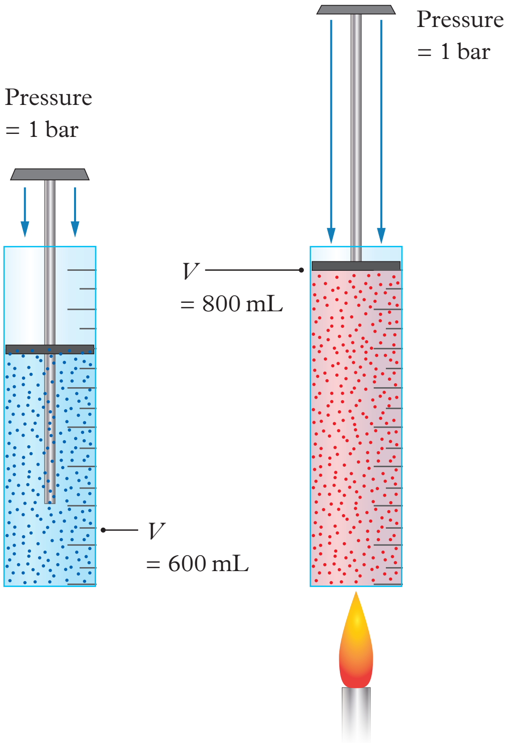


FIGURE 5 The pressure (1 bar) applied to the cylinder is identical throughout the experiment. When the temperature is increased, the volume that the gas occupies also increases.

TABLE 2 Data obtained from the experiment in Figure 5

Temperature (°C)	Temperature (K)	Volume (mL)	V/T (mL K^{-1})
–124	149	600	4.021
–99	174	700	4.021
–74	199	800	4.021
–49	224	900	4.021
–24	249	1,000	4.021
1	274	1,100	4.021
25	298	1,200	4.021
50	323	1,300	4.021
75	348	1,400	4.021
100	373	1,500	4.021

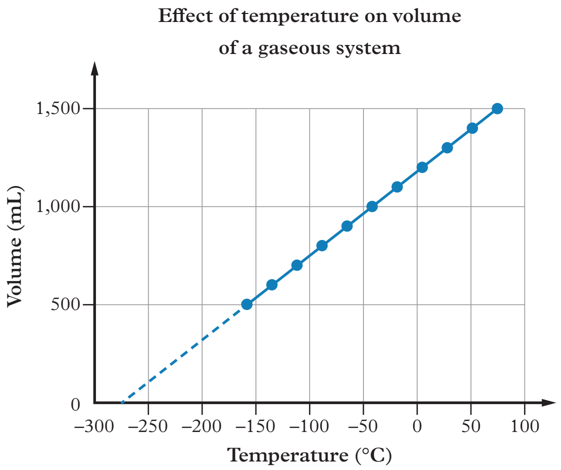


FIGURE 6 The relationship between temperature (°C) and volume in a gaseous system

Figure 6 shows the graph extrapolating to a volume of 0 mL. The *y*-axis intercept here is –273°C. This value is exactly 0 kelvin (K). There is a linear relationship between temperature and volume. Therefore, volume is proportional to temperature ($V \propto T$). This relationship can therefore be represented mathematically as:

$$\frac{V}{T} = \text{a constant}$$

Any change in temperature or volume can be calculated from the equation:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

In such calculations, temperature is in kelvin.

Worked example 11.3B

Calculating volume and temperature changes

Worked example 11.3B: Watch a video that shows how to solve this problem.

At 20°C, the volume in a cylinder is 200 mL. If the temperature is raised to 62°C, **calculate** the final volume of the cylinder. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the final volume (V_2). The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $V_1 = 200\,\text{mL}$, $T_1 = 20°\text{C}$ $V_2 = ?$, $T_2 = 62°\text{C}$
Step 3: Substitute the known values into the formula and solve for the unknown value. Complete any required conversions and make sure all temperatures are converted to kelvin.	$\frac{200\,\text{mL}}{20°\text{C} + 273} = \frac{V_2}{62°\text{C} + 273}$ (1 mark) $V_2 = 228.67\,\text{mL}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	229 mL (1 mark)

Your turn

At 18°C, the volume in a cylinder is 125 mL. If the volume is increased to 450 mL, **calculate** the final temperature of the cylinder. (2 marks)

Boyle's law: how are pressure and volume related?

In a variable volume system at a fixed temperature, the volume must decrease to increase the pressure. If the volume decreases, the molecules within the system have less space to move in. Therefore, the molecules hit the walls of the container more often, increasing the pressure. Boyle's law states that pressure is inversely proportional to volume in fixed temperature systems.

Numerical representation of Boyle's law

The relationship between pressure and volume is represented in Figure 7. A syringe filled with a gas is kept at a constant temperature. The plunger on the syringe is pushed down. As the plunger is pushed down further, more pressure is required to keep pushing it. The volume within the syringe decreases the further down the plunger is pushed and the pressure increases. Therefore, as pressure increases, volume decreases. Table 3 summarises the data gathered in the experiment.

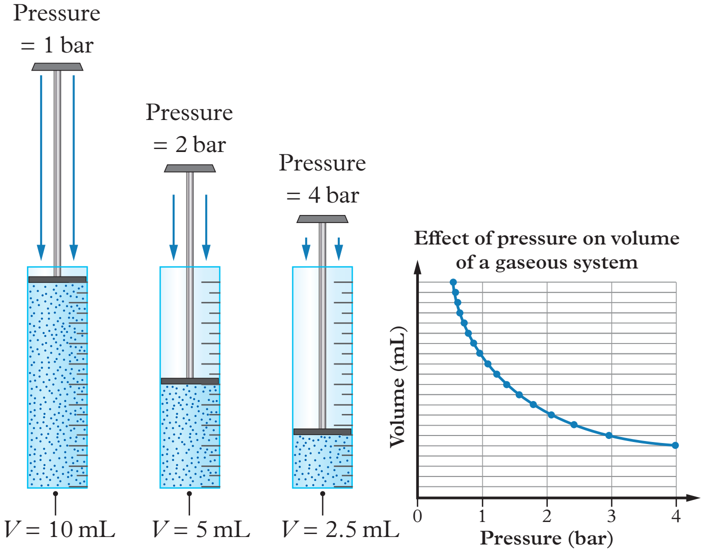


FIGURE 7 The temperature of the cylinder stays the same, but when the plunger is depressed, pressure increases and volume decreases.

TABLE 3 Data obtained from the experiment in Figure 7

Volume (mL)	Pressure (bar)	$P \times V$
10	1	1 × 10 = 10
5	2	2 × 5 = 10
2.5	4	4 × 2.5 = 10

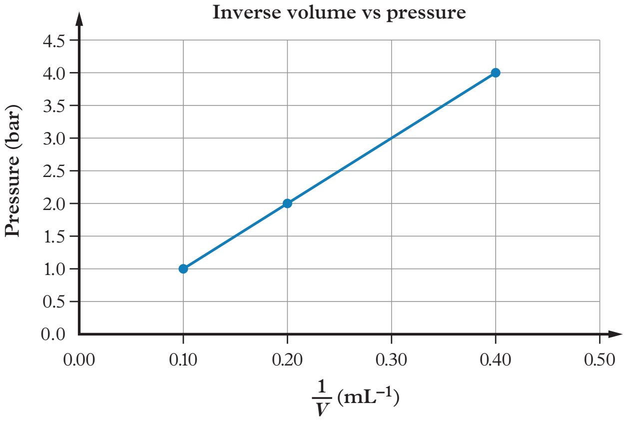


FIGURE 8 The relationship between pressure and $\frac{1}{V}$

When the pressure is graphed against $\frac{1}{V}$, a linear relationship is seen (Figure 8). The graph demonstrates that there is a linear relationship between pressure and inverse volume. Therefore, pressure is proportional to $\frac{1}{V}\left(P \propto \frac{1}{V}\right)$, or we can say that pressure is inversely proportional to volume. This relationship can be represented mathematically as:

$$P \times V = \text{a constant}$$

As with Gay-Lussac's and Charles' law, any change in pressure or volume at a constant temperature can be calculated using the following equation:

$$P_1V_1 = P_2V_2$$

Worked example 11.3C

Calculating pressure and volume changes

Worked example 11.3C: Watch a video that shows how to solve this problem.

At 1.00 bar, a sample of argon gas occupies 10 L of a cylinder. If the pressure is increased to 1.7 bar, **calculate** the volume of the cylinder that the gas will occupy. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the final volume (V_2). The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$P_1V_1 = P_2V_2$ $P_1 = 1$ bar, $V_1 = 10$ L $P_2 = 1.7$ bar, $V_2 = ?$
Step 3: Substitute the known values into the formula and solve for the unknown value.	1 bar × 10 L = 1.7 bar × V_2 (1 mark) $V_2 = 5.88235$ L
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	5.9 L (1 mark)

Your turn

A variable volume container has a volume of 10.0 L and a pressure of 4.0 bar. If the volume is decreased to 300 mL, **calculate** the final pressure of gas in the container. (2 marks)

Avogadro's law: what is molar volume?

The relationship between pressure, volume and temperature can be extended to the number of atoms or molecules in a gaseous system. Amadeo Avogadro hypothesised that under identical conditions of temperature and pressure, equal volumes of gases contain an equal number of gas particles. Therefore, one mole of any gas (6.02×10^{23} molecules), at the same temperature and pressure, will occupy a specific volume (Figure 9). This is called the **molar volume**.

molar volume
the volume occupied by 1 mole of a gas at a specified temperature and pressure

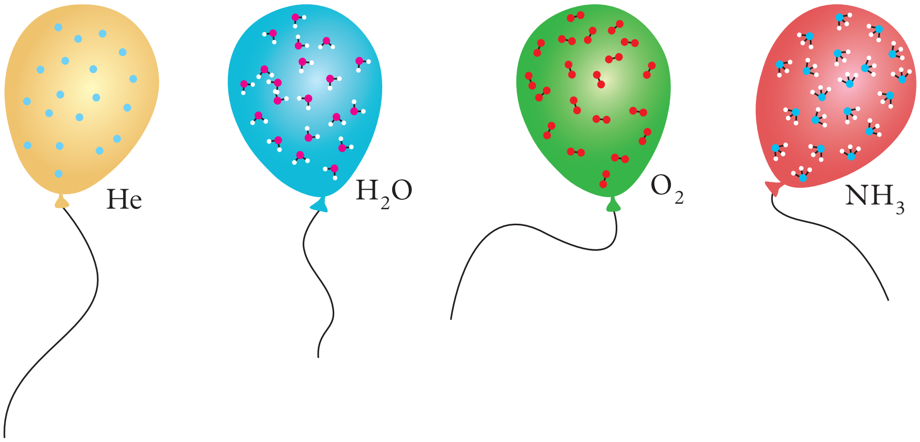


FIGURE 9 Four balloons of the same volume, at the same temperature and pressure, have the same number of gas particles.

standard temperature and pressure (STP)
standard conditions for temperature and pressure, when gases are measured at 100 kPa (1 bar) and 0°C (273 K)

standard laboratory conditions (SLC)
standard conditions for temperature and pressure under laboratory conditions, when gases are measured at 100 kPa (1 bar) and 25°C (298 K)

Standard conditions

To compare the gases in two different systems, scientists keep the systems at the same standard conditions. These are sets of conditions that are kept constant when measuring gases. By keeping temperature and pressure constant, the only variables should be volume and number of moles.

- At **standard temperature and pressure (STP)**, 0°C (273 K) and 100 kPa (1 bar), 1 mole of any gas occupies 22.7 dm^3, which is the equivalent of 22.7 L.
- At **standard laboratory conditions (SLC)**, 25°C (298 K) and 100 kPa (1 bar), 1 mole of any gas occupies 24.8 dm^3, which is the equivalent of 24.8 L.

Study tip

The data book does not state the value of SLC, so it will not be used in exam conditions unless stated. It does, however, define STP as

- $2.27 \times 10^{-2}\ m^3\ mol^{-1}$; as m^3 is 1,000 L, this is the same as kilolitres per mole
- $22.7\ dm^3\ mol^{-1}$; as dm^3 is L, this is the same as litres per mole.

Numerical representation of Avogadro's law

The relationship between moles and volume can be represented by the equation:

$$n = \frac{V}{V_m}$$

where n is the number of moles, V is the volume (L) and V_m is the molar volume ($L\,mol^{-1}$).

At STP this equation becomes:

$$n_{STP} = \frac{V}{22.7\,L\,mol^{-1}}$$

At SLC this equation becomes:

$$n_{SLC} = \frac{V}{24.8\,L\,mol^{-1}}$$

Worked example 11.3D

Calculating the amount of gas at STP

▶ **Worked example 11.3D:** Watch a video that shows how to solve this problem.

Calculate the amount (in mol) of nitrogen gas in a 40.0 L cylinder at STP. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the amount of a gas. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$n = \frac{V}{V_m}$ $V = 40.0\,L$, $V_m = 22.7\,L\,mol^{-1}$, $n = ?$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$n(N_2) = \frac{40.0\,L}{22.7\,L\,mol^{-1}}$ (1 mark) $= 1.762\,mol$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	1.76 mol (1 mark)

Your turn

Calculate the amount (in mol) of propane gas in a 11.0 L barbecue cylinder at STP. (2 marks)

Worked example 11.3E

Calculating the mass of gases at STP

▶ **Worked example 11.3E:** Watch a video that shows how to solve this problem.

Calculate the mass of helium in a 20.0 L cylinder at STP. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the mass of a gas. The question is worth 4 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$n = \frac{V}{V_m}$ $V = 20.0\,L$, $V_m = 22.7\,L\,mol^{-1}$, $n = ?$ $n = \frac{m}{M}$ $m = ?$, $M = 4.00\,g\,mol^{-1}$
Step 3: Substitute the known values into the formula and solve for moles.	$n = \frac{20.0\,L}{22.7\,L\,mol^{-1}}$ (1 mark) $= 0.88\,mol$ (1 mark)
Step 4: Substitute the known values into the formula and solve for mass.	$0.88\,mol = \frac{m}{4.00\,g\,mol^{-1}}$ (1 mark) $m = 3.524$
Step 5: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	3.52 g (1 mark)

Your turn

Calculate the mass of ethyne (C_2H_2) gas that occupies a 15 L cylinder of gas at STP. (4 marks)

Worked example 11.3F

Determining the identity of a gas

Worked example 11.3F: Watch a video that shows how to solve this problem.

A balloon is filled to 632.34 mL with 2.3344 g of an unknown monatomic gas at STP. **Determine** the identity of the gas. (4 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	“Determine” means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to determine the amount of gas, then find its molar mass in order to identify it. The question is worth 5 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Select the correct formula and gather any data required.	$n = \frac{V}{V_m}$ $V = 632.34\,\text{mL}$, $V_m = 22.7\,\text{L mol}^{-1}$ $n = \frac{m}{M}$ $m = 2.3344\,\text{g}$, $M = ?$
Step 3: Substitute the known values into the formula and solve for moles. Complete any required conversions.	$n = \frac{632.34\,\text{mL}}{22.7\,\text{L mol}^{-1}} \times \frac{1\,\text{L}}{1{,}000\,\text{mL}}$ (1 mark) $= 0.028\,\text{mol}$ (1 mark)
Step 4: Substitute the known values into the formula and solve for molar mass.	$0.028\,\text{mol} = \frac{2.334\,\text{g}}{M}$ (1 mark) $M = 83.79\,\text{g mol}^{-1}$ (1 mark)
Step 5: Finalise your answer.	The monatomic gas with a molar mass of $83.79\,\text{g mol}^{-1}$ is krypton. (1 mark)

Your turn

2.56 g of a diatomic gas fills a container to 820 mL at STP. **Determine** the identity of the gas. (4 marks)

Real-world chemistry

The invention of the Aqua-Lung

Jacques-Yves Cousteau (1910–1997), a French naval officer, oceanographer and film maker was an avid diver and environmentalist. He often stated that he was miserable out of the water. It was like he had been placed into heaven (in the water) and was then forced back onto the earth.

He wanted to go deeper into the water but was restricted by the available equipment. At the time, divers would wear lead-soled boots to weigh them down and have a pipe attached to a bubble style helmet which reaching the surface for breathing (Figure 10). But Cousteau wanted to swim and explore to explore the oceans with freedom and without the cumbrance of lead shoes.

If he could use a tank which stored enough air, then he could breathe using the air from the tank and stay underwater for longer periods of time. However, longer dive times required more air to be stored in a fixed volume tank at higher pressures. These high-pressure tanks could not be directly fitted to a hose and breathed from, as all of the gas would escape within a few seconds.

FIGURE 10 A lead-footed and bubble-headed diving suit

Through his wife's father, Cousteau met Emilie Gagnan, a French engineer. Gagnan had developed a regulator intended to power motor cars. The regulator could convert the high pressures in the air tank to more moderate pressures that could be used for breathing. This allowed more air to be contained within the tank, enabling deeper dives and longer dive times. Cousteau borrowed a regulator and mounted it on air bottles. It was designed to allow small amounts of air out of the bottles and not allow the expelled air back in.

From this, he built the first underwater breathing unit, the Aqua-Lung. Through numerous improvements to the design, he was able to dive to depths of more than 60 metres.

The fathers of modern diving, Cousteau and Gagnan are credited with the invention of the more modern SCUBA (self-contained underwater breathing apparatus) gear that we use today.

Apply your understanding

1 **Explain** why high-pressure air tanks are required when diving. (2 marks)

2 Air tanks require thick metal reinforced walls. **Justify** why these are required based on the purpose of the diving tanks. (2 marks)

Real-world chemistry

The effect of gases on the body while scuba diving

Scuba allows humans to explore and play under water, by providing the oxygen required to breathe under water. A thorough understanding of gas laws and their effect on the human body is important to avoid harm or death during scuba diving.

Atmospheric pressure at sea level is 1 atm (equivalent to 1.01325 bar). Under water, pressure increases by 1 atm every 10 m in depth. When pressure on the body increases, the body is compressed and the volume of gas within the body decreases. This phenomenon follows Boyle's law.

While descending, extra pressure is placed on the body, which damages tissues that contain or are surrounded by air spaces. This can rupture the eardrums and cause sinus pain. To avoid these issues, divers are taught how to clear their ears and equalise the pressure in their body.

The opposite occurs while ascending: the pressure on the body decreases. As this occurs, nitrogen dissolved in the bloodstream and tissues form bubbles of gas again. It is a similar effect to when opening a bottle of fizzy cordial. When the bottle is under pressure and the lid is then removed, large gas bubbles form rapidly, causing the bottle to overflow. Nitrogen gas must be removed and can only be done so naturally and gradually, by the body.

FIGURE 11 A scuba diver explores the Great Barrier Reef, near Port Douglas.

A slow ascension rate is the best method to remove the excess nitrogen. Taking breaks ("decompression stops") is also recommended as a diver ascends – the diver remains at a certain depth until it is safe to continue ascending. The nitrogen slowly moves from the tissues back into the bloodstream from where it is exhaled. If the nitrogen within the tissues is not removed slowly, it forms bubbles that can harm the body in several ways, such as damaging the fragile tissue surrounding the air sacs in the lungs.

Apply your understanding

1 **Explain** why pressure varies when diving and what effect this has on gaseous volume within the body. (2 marks)

2 Pressure decreases and air becomes thinner with increasing altitude. **Apply** your knowledge of gas laws to **predict** what may happen to the human body at higher altitudes. (3 marks)

Check your learning 11.3

Check your learning 11.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the relationship between each of the following variables in one sentence:

- a pressure and volume (1 mark)
- b pressure and temperature (1 mark)
- c volume and temperature. (1 mark)

2 **Calculate** the volume of:

- a 3.2 mol of water vapour at STP (2 marks)
- b 2.9 g of oxygen gas at STP (4 marks)
- c 9.37×10^{21} nitrogen molecules at STP. (4 marks)

3 **Calculate** the:

- a mass of 120 mL of CO_2 at STP (4 marks)
- b number of H atoms in 525 mL of CH_4 at STP. (4 marks)

4 **Calculate** temperatures in Kelvin for the following values:

- a –25°C (1 mark)
- b 13°C (1 mark)
- c 273°C. (1 mark)

Analytical processes

5 At a constant temperature and at a pressure of 106.5 kPa, a gas sample occupies 120 mL of a container. **Deduce** what pressure would be required to increase the volume by 680 mL. (2 marks)

6 Figure 13 shows the data obtained when a volume of gas is heated at a constant pressure.

- a **Determine** the constant that represents the relationship between volume and temperature. (2 marks)

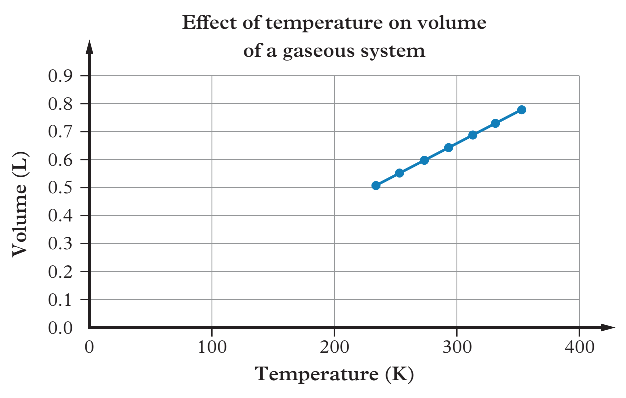


FIGURE 12 The effect on volume when temperature is raised in a variable volume system

- b The temperature of the gas continues to increase until it reaches 473 K. **Determine** what volume the gas occupies at 473 K. (2 marks)

Knowledge utilisation

7 **Discuss** why heating the air inside a hot air balloon causes the balloon to rise into the atmosphere. (2 marks)

8 Theoretically, the volume of a gas could be 0 L.

- a **Comment** on what would need to happen for this to occur. (1 mark)
- b **Discuss** why this is theoretical. (2 marks)

9 An experiment found the density of nitrogen gas to be 1.2504 $g\,L^{-1}$ at STP. **Predict** what would happen to the pressure of the system if the density of this gas was increased in a:

- a fixed volume container (2 marks)
- b variable volume container. (2 marks)

Practical

Lesson 11.4

Investigating Boyle's law

Learning intentions and success criteria

Video demonstration

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 11.5

Ideal gases

Key ideas

→ The gas laws can be combined to form the ideal gas equation.

→ The universal gas constant R is 8.31 J K^{-1} mol^{-1}.

What is the ideal gas equation?

Learning intentions and success criteria

ideal gas
a gas in which particles do not interact with one another and occupy no volume

Most gases will behave like an **ideal gas** at STP. In theory, this assumes that the gas particles do not occupy any volume and are not attracted to each other or to the walls of the container. The particles are in continuous, random motion and only change direction after a collision. All collisions are elastic, meaning that energy is not lost when particles collide. Assuming gases behave ideally at STP makes it more straightforward to predict the behaviour of gases under different conditions.

From Lesson 11.3, we can summarise the gas laws as:

- $\frac{P}{T}$ = a constant
- $\frac{V}{T}$ = a constant
- $P \times V$ = a constant
- $V_m = \frac{V}{n}$ where molar volume (V_m) is a constant at a specific temperature and pressure.

These laws can be combined to generate the ideal gas equation:

$$R = \frac{P \times V}{n \times T}$$

universal gas constant (R)
the proportionality constant that defines gas behaviour under ideal conditions

where R is the **universal gas constant** and temperature (T) is expressed in Kelvin.

The ideal gas equation, also called the universal gas equation, provides chemists with a simpler method of determining experimental values in gas experiments without having to complete multiple calculations.

Under standard conditions, where pressure is 100,000 Pa (100,000 N m^{-2}), 1 mole is contained within a volume of 22.7 L (0.0227 m^3), temperature is 273 K, the universal gas constant, R, can be derived:

$$\begin{aligned} R &= \frac{100{,}000\,\text{N}\,\text{m}^{-2} \times 0.0227\,\text{m}^3}{1\,\text{mol} \times 273\,\text{K}} \\ &= 8.31\,\text{N}\,\text{m}\,\text{mol}^{-1}\,\text{K}^{-1} \end{aligned}$$

As 1 N m is equal to 1 J, the universal gas constant is most commonly expressed as $R = 8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$.

The form in which you will use the ideal gas equation is:

$$PV = nRT$$

where P is pressure (kPa), V is volume (L), n is the number of moles, T is temperature (K) and $R = 8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$.

Study tip

The ideal gas constant (R) can be found in the Formula and data book: $8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$

Worked example 11.5A

Calculating volume using the ideal gas equation

Worked example 11.5A: Watch a video that shows how to solve this problem.

Use the ideal gas equation to **calculate** the volume that 0.94 mol of nitrogen (N_2) gas occupies at STP. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Use" means to apply knowledge or rules to put theory into practice. "Calculate" means to determine or find a number or answer by using mathematical processes. We need to apply the ideal gas equation to determine the volume. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$PV = nRT$ $P = 1$ bar or 100 kPa, $V = ?$ $n = 0.94$ mol, $R = 8.31\ \text{J mol}^{-1}\ \text{K}^{-1}$, $T = 0°\text{C}$ or 273 K
Step 3: Substitute the known values into the formula and solve for the unknown value. Make sure you use the correct units.	$100\ \text{kPa} \times V = 0.94\ \text{mol} \times 8.31\ \text{J mol}^{-1}\ \text{K}^{-1} \times 273\ \text{K}$ (1 mark) $V = 21.33$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	21 L (1 mark)

Your turn

Use the ideal gas equation to **calculate** the volume that 3.60 mol of carbon dioxide (CO_2) gas occupies at STP. (2 marks)

Using the formula that relates moles, mass and molar mass, the ideal gas equation can be expanded to:

$$PV = \frac{mRT}{M}$$

Worked example 11.5B

Calculating the mass of a gas using the ideal gas equation

Worked example 11.5B: Watch a video that shows how to solve this problem.

Use the ideal gas equation to **calculate** the mass of carbon dioxide that occupies a 2.00 L cylinder at STP. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Use" means to apply knowledge or rules to put theory into practice. "Calculate" means to determine or find a number or answer by using mathematical processes. We need to apply the ideal gas equation to determine the volume. The question is worth 4 marks, so we must gather the correct information, correctly substitute into the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$PV = nRT$ $P = 100\,\text{kPa}$, $V = 2.00\,\text{L}$ $n = ?$, $R = 8.31\,\text{J}\,\text{mol}^{-1}\,\text{K}^{-1}$, $T = 0°\text{C}$ $n = \frac{m}{M}$ $m = ?$, $M = 44.01\,\text{g}\,\text{mol}^{-1}$
Step 3: Substitute the known values into the formula and solve for the amount (in moles). Make sure you use the correct units.	$100\,\text{kPa} \times 2.00\,\text{L} = n \times 8.31\,\text{J}\,\text{mol}^{-1}\,\text{K}^{-1} \times (0°\text{C} + 273)$ (1 mark) $n = 0.08816\,\text{mol}$ (1 mark)
Step 4: Substitute the known values into the formula and solve for mass.	$0.08816\,\text{mol} = \frac{m}{44.01\,\text{g}\,\text{mol}^{-1}}$ (1 mark) $m = 3.8799\,\text{g}$
Step 5: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	3.88 g (1 mark)

Your turn

Use the ideal gas equation to **calculate** the mass of helium gas in a 5.00 L weather balloon at STP. (4 marks)

Study tip

There can be many ways of answering these types of questions. As long as you work through the question logically and correctly complete your calculations, the answer will still be marked as correct.

Study tip

Spaced practice is a useful technique to recall prior learning. Build time in your study schedule to practise problems from earlier in the course. This is very helpful for stoichiometry problems.

How do we apply the ideal gas equation to gases in chemical reactions?

The ideal gas equation can be used in stoichiometric calculations to calculate quantities of gaseous reactants and products. If pressure, volume and temperature of a gas are known, the amount (in moles) can be calculated and used. If you are able to calculate the amount (in moles) of gas involved in a reaction, then you can easily find the volume it will occupy. Recall how to write chemical equations and calculate amounts in reactions from Unit 1.

Worked example 11.5C

Using the ideal gas equation in stoichiometry

Worked example 11.5C: Watch a video that shows how to solve this problem.

Petrol is a mixture of hydrocarbons. For this question, we will use the formula of octane C_8H_{18} to approximately represent petrol.

a **Determine** the balanced chemical equation for the complete combustion of octane. (1 mark)

b **Calculate** the volume of carbon dioxide at STP produced by the combustion of 703 g of octane, which approximates to 1 L of petrol. (4 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. "Calculate" means to determine or find a number or answer by using mathematical processes. We need to write a chemical equation and then apply the ideal gas equation to determine the volume of gas produced. The questions are worth a variety of marks, so we must gather the correct information, write the equation, correctly substitute into the formula and use stoichiometry to complete the calculation.
Step 2: For part a, write a balanced the chemical equation for the reaction, including states.	$C_8H_{18}(l) + \frac{25}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g)$ (1 mark)
Step 3: For part b, select the correct formula and gather any data required.	$n = \frac{m}{M}$ $n = ?$, $m(C_8H_{18}) = 703\,g$, $M(C_8H_{18}) = 114.26\,g\,mol^{-1}$ $PV = nRT$ $P = 100\,kPa$, $V = ?$ $n = ?$, $R = 8.31\,J\,K^{-1}\,mol^{-1}$, $T = 0°C$
Step 4: Substitute the known values into the formula and solve for the amount (in mol) of octane.	$n(C_8H_{18}) = \frac{703\,g}{114.26\,g\,mol^{-1}}$ (1 mark) $= 6.1526\,mol$ (1 mark)
Step 5: Use the mole ratio to find the amount (in mol) of carbon dioxide.	$n(CO_2) = \frac{8}{1} \times n(C_8H_{18})$ $= \frac{9}{1} \times 6.1526\,mol$ $= 49.2211\,mol$
Step 6: Substitute the known values into the formula and solve for the volume of carbon dioxide.	$100\,kPa \times V = 49.2211\,mol \times 8.31\,J\,mol^{-1}\,K^{-1} \times (0°C + 273)$(1 mark) $V = 1116.6\,L$
Step 7: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	$1.12 \times 10^3\,L$ (1 mark)

Your turn

During photosynthesis, plants use carbon dioxide from air, together with liquid water in their cells, to make an aqueous solution of glucose $C_6H_{12}O_6$. Oxygen is a byproduct.

a **Determine** the balanced chemical equation for the photosynthesis reaction. (1 mark)

b **Calculate** the volume of carbon dioxide at STP used in the synthesis of 100 g of glucose. (4 marks)

Skill drill

Investigating the effect of vaping on lung capacity

Science inquiry skill: Evaluating evidence (Lesson 1.8)

To determine the effect of vaping on lung capacity in 25-year-olds, a scientist sets up an experiment where participants blow a breath into the clear plastic tube, which is captured in the plastic jug and measured (Figure 1).

The scientist first tests the lungs of participants who did not vape or smoke. The average lung capacity reported was 6.0 L. The next day, which is approximately 10°C hotter, the scientist is forced to use a different laboratory at a higher altitude. On this day, the participants who vape recorded an average lung capacity of 6.0 L.

The scientist concluded that vaping had no effect on lung capacity.

Practise your skills

1 **Identify** the independent, dependent and controlled variables. (3 marks)

2 **Construct** a hypothesis for the experiment. (3 marks)

3 **Comment** on the accuracy and reliability of the data. (4 marks)

3 **Propose** two ways in which the experiment could be improved. (2 marks)

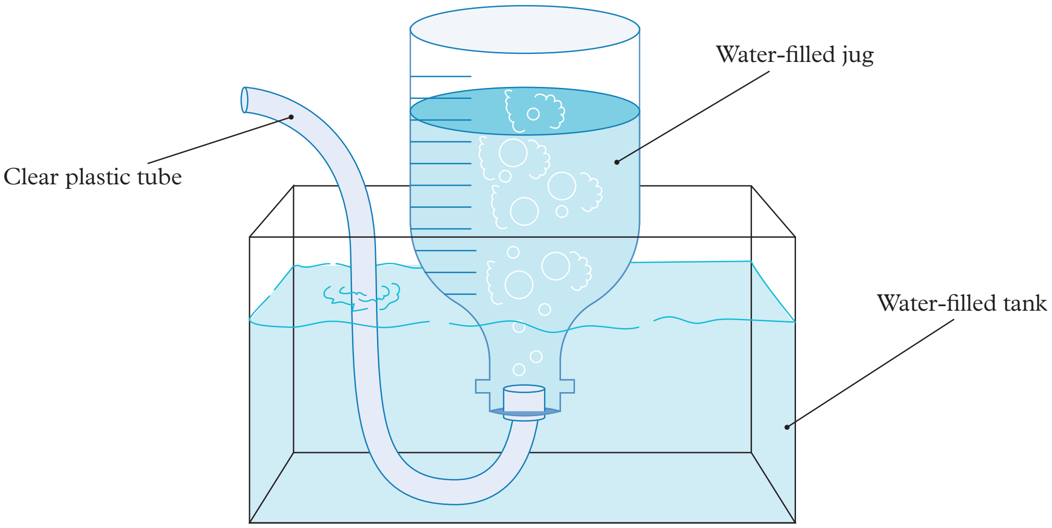


FIGURE 1 Lung capacity apparatus

Check your learning 11.5

Check your learning 11.5: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Recall** the pressure and temperature at STP. (1 mark)

2 **Describe** four characteristics of an "ideal" gas. (4 marks)

3 The ideal gas equation relates the amount, pressure, volume and temperature of a gas with the universal gas constant.

- a **Calculate** the pressure 2.0 mol of H_2 gas would exert on a 250 mL cylinder at 10°C. (2 marks)
- b A 0.10 g sample of He is used to fill a balloon. The sample has a pressure of 1.3 bar at 19°C. **Calculate** the volume of the balloon. (2 marks)
- c At STP, **calculate** the mass of xenon gas that can be used to fill a 1.4 L cylinder. (4 marks)

Analytical processes

4 **Compare** the following samples and **determine** which has the greater mass: 1.3 L of N_2 at 0.6 bar and 15°C or 0.5 L of water vapour at 123.8 kPa and 303.6 K. (5 marks)

Knowledge utilisation

5 A 1.82 g sample of propane gas occupies 1.00 L at 1 bar and 21.1°C. **Prove** that the molecular formula of propane is C_3H_8. (3 marks)

6 During cellular respiration, dissolved glucose ($C_6H_{12}O_6$) reacts with oxygen. Carbon dioxide gas and liquid water are the products.

- a **Determine** the balanced chemical equation for the respiration equation. (1 mark)
- b **Calculate** the volume of oxygen at STP used in the respiration of 1.00 g of glucose. (4 marks)

7 1.0 tonne of nitrogen gas and 0.200 tonne of hydrogen gas are allowed to react to form ammonia gas (1.00 tonne = 1×10^6 g)

- a **Determine** the balanced chemical equation for the reaction. (1 mark)
- b **Determine** the limiting reactant. (5 marks)
- c **Calculate** the theoretical volume of ammonia at 500°C and 30,000 kPa that could be produced from these quantities of reactions. (2 marks)
- d The actual volume of ammonia obtained under these conditions is 1,700 L. **Calculate** the percentage yield of ammonia obtained. (2 marks)

FIGURE 2 The volume of oxygen gas used in cellular respiration can be calculated using stoichiometry and the ideal gas equation.

MODULE 11

Lesson 11.6

Review: Gases

Summary

- 11.1 Kinetic molecular theory can be used to explain gas behaviour and properties.
- 11.2 Pressure, temperature and volume affect the behaviour of gases.
- Practical: Investigating the properties of gases
- 11.3 As temperature increases, pressure increases (Gay-Lussac's law); as temperature increases, the volume of a variable volume system increases (Charles' law); as volume increases, pressure decreases (Boyle's law).
- Gases occupy the same molar volume under standard temperature and pressure (STP).
- 11.4 Practical: Investigating Boyle's law
- 11.5 The gas laws can be combined to form the ideal gas equation. The universal gas constant R is $8.31\,\mathrm{J\,mol^{-1}\,K^{-1}}$.

Key formulas

Gay-Lussac's law	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Charles' law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Boyle's law	$P_1V_1 = P_2V_2$
Molar volume	$n = \frac{V}{V_m}$
	$n_{STP} = \frac{V}{22.7\,\mathrm{L\,mol^{-1}}}$
	$n_{SLC} = \frac{V}{24.8\,\mathrm{L\,mol^{-1}}}$
Ideal gas equation	$PV = nRT$
	$PV = \frac{mRT}{M}$

Review questions 11.6A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Kinetic molecular theory states that

A not all particles are in motion.

B particle movement is predictable.

C particle movement is circular.

D particle collisions are elastic and conserve energy.

2 According to Boyle's law, as the volume of a container decreases,

A pressure decreases.

B temperature increases.

C pressure increases.

D temperature decreases.

3 The pressure of a container of gas at 50°C and 5 bar is increased by 97 kPa. What is the temperature increase of the container?

A 62°C
B 59.57°C
C 50.02°C
D 9.57°C

4 Which of the following has the greatest number of molecules?

A 1.00 L of N_2 at STP
B 2.5 L of O_2 at 2°C and 0.5 bar
C 0.9 L of H_2O at STP
D 0.5 L of CO_2 at 50°C and 1.5 bar

5 Which of the following includes the correct units for the ideal gas equation?

A Pressure in kPa, volume in L, temperature in K
B Pressure in Pa, volume in L, temperature in K
C Pressure in kPa, volume in mL, temperature in K
D Pressure in kPa, volume in mL, temperature in °C

6 73 K is equal to

A –200°C.
B 0°C.
C 346°C.
D 200°C.

7 What is the amount (in mol) of oxygen gas in a 32.0 L cylinder at STP?

A 1.43 mol
B 0.775 mol
C 1.29 mol
D 1.41 mol

8 A syringe containing gas is compressed by pushing in the plunger to halve the volume. Which of the following statements about the pressure is true?

A It is doubled because the particles are colliding with more force due to being closer together.
B It is doubled because the particles are closer together causing more collisions on the surface area.
C It is halved because pressure is proportional to volume and it has been halved.
D It is halved because some of the air leaked out of the syringe when the plunger was pushed in.

9 Gas–Lussac's Law relates pressure and temperature. The main reason that pressure increases as temperature rises is that at a higher temperature, the particles move faster and

A collide with more force.
B there are many more collisions.
C cause the volume of the container to change, affecting the pressure.
D cause the particles to chemically decompose, producing more particles.

10 Which of the following statements is true?

A Equal volumes of gases at the same temperature and pressure have the same mass.
B Equal volumes of gases at the same temperature and pressure have the same average particle velocity.
C Equal volumes of gases at the same temperature and pressure have the same number of particles.
D Equal volumes of gases at the same temperature and pressure have the same molar mass.

Review questions 11.6B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

11 A variable volume container of gas is prepared. **Explain** how:

a increasing the volume will affect the pressure of the gas, assuming temperature is kept constant. (2 marks)
b decreasing the temperature will affect the volume of the gas. (2 marks)

12 A pressure cooker is used to increase the pressure in a pot. **Explain** how it is able to cook food. (2 marks)

FIGURE 1 Rice cooker

13 **Describe** two effects of increasing the number of molecules of a gas in a variable volume system. (2 marks)

14 **Describe** the effect of an increase in pressure on the temperature of a fixed volume system. (1 mark)

15 **Calculate** the volume (in L) of:

- **a** 5.3 g of water vapour at STP (4 marks)
- **b** 4.44 moles of oxygen gas at STP (2 marks)
- **c** 19.0 g of neon gas at STP (4 marks)
- **d** 4.522×10^{22} nitrogen molecules at STP (4 marks)
- **e** 3.2 g of carbon dioxide at STP (4 marks)
- **f** 5.2 g of argon gas at STP. (4 marks)

16 Cylinders of oxygen gas used in hospitals store oxygen at a pressure of 13,700 kPa. As it flows from the cylinder, the oxygen pressure decreases to 100 kPa. **Calculate** the volume of oxygen that could be obtained from a 20 L cylinder, assuming that temperature is the same. Remember that the last 20 L in the cylinder will not flow out. (3 marks)

17 **Use** your understanding of gas laws to complete the following.

- **a** A 50 mL vessel has a pressure of 104.5 kPa. If the volume is increased to 420 mL, **calculate** the pressure of the system in bar. (2 marks)
- **b** **Calculate** the change in temperature needed to increase the pressure of a 117.3 kPa container at 20°C to 134.2 kPa. (3 marks)

18 **Use** your understanding of gas laws to complete the following.

- **a** A 40 L cylinder of O_2 at 13°C has a pressure of 98 kPa. **Calculate** the mass of the gas. (2 marks)
- **b** **Calculate** the volume occupied by 12 g of water vapour at a pressure of 67 kPa and a temperature of 123 K. (2 marks)
- **c** **Calculate** the pressure of a 4.00 L gas container containing 67.0 mol of He at 5.00°C. (2 marks)

Analytical processes

19 **Use** the kinetic molecular theory, to **contrast** the molar volume of a gas at SLC and STP. (4 marks)

20 **Use** the following information to **determine** the identity of the monatomic gases.

- **a** 8.21 g of an unknown monatomic gas fills a balloon to 847 mL at STP. (3 marks)
- **b** 380 mg of an unknown monatomic gas fills a balloon to 216 mL at STP. (3 marks)
- **c** 1.29 g of an unknown monatomic gas fills a balloon to 733 mL at STP. (3 marks)

21 Weather balloons are strong helium balloons inflated to a volume approximately 2 m in diameter. They carry instruments to measure the atmosphere at high altitudes. As they rise up through the atmosphere, they expand to approximately 20 m across, and eventually burst, dropping the instruments back to Earth where they can be recovered. **Apply** your understanding of the gas laws and kinetic theory to **explain** the increase in volume and the eventual bursting of the balloon. (3 marks)

22 Petrol combustion engines ignite the fuel with a spark. Diesel engines do not use a spark. In a diesel engine, air is introduced into the cylinder. The air in the cylinder is compressed rapidly, then the diesel is introduced and ignites. **Apply** your understanding of the variables in the ideal gas equation and **use** them to **explain** how ignition could occur without a spark. (2 marks)

23 A student is designing an experiment to determine the molar volume of a gas. They plan to trap the hydrogen produced when magnesium reacts with hydrochloric acid in a gas tube. The volume of hydrogen at STP needs to be close to but less than the 50.0 mL tube. **Determine** the maximum mass of magnesium that could be used assuming it is the limiting reactant. (5 marks)

Knowledge utilisation

24 A student takes a plastic pump water bottle to school on a snowy winter's day. The student travels to school by car and the car heater is on. It takes almost an hour to get to school. On arriving at school, the student has a mouthful of water. A small amount of gas exits the bottle when it is opened. At lunchtime, the student takes the water bottle outside and accidentally leaves it there. The sides of the bottle become crumpled in. **Discuss** what has happened to the water bottle to cause both effects. (4 marks)

25 Ping pong balls can become dented throughout the course of a game. **Propose** how you could remove the dent by using your knowledge of gas laws. (2 marks)

FIGURE 2 Ping pong ball

26 A turkey timer is used to indicate when a turkey is cooked. **Investigate** the turkey timer and **use** gas laws to **explain** how it works. (3 marks)

FIGURE 3 Turkey timer

27 Tyres are a key safety feature of a car, responsible for the friction between the car and road, which allows for acceleration and deceleration, or breaking.

- a **Discuss** why tyre pressure should be checked on a cool day. (2 marks)
- b **Predict** what would happen if you over-filled a car tyre (without it exploding). **Justify** your answer. (2 marks)
- c **Predict** what would happen if you under-filled a car tyre. **Justify** your answer. (2 marks)

28 **Propose** a reason why the ideal gas equation only applies to ideal gases. (2 marks)

Data drill

Analysing the quantity of gas produced in a reaction

To investigate the effect of temperature on the quantity of gas produced in a reaction, three identical pieces of magnesium ribbon, each with a mass of 3.64 g, are prepared. On three separate days, one of the pieces of magnesium ribbon is reacted with 20.00 mL of 2.0 M HCl in a conical flask which has a gas adaptor attached to a syringe (Figure 1).

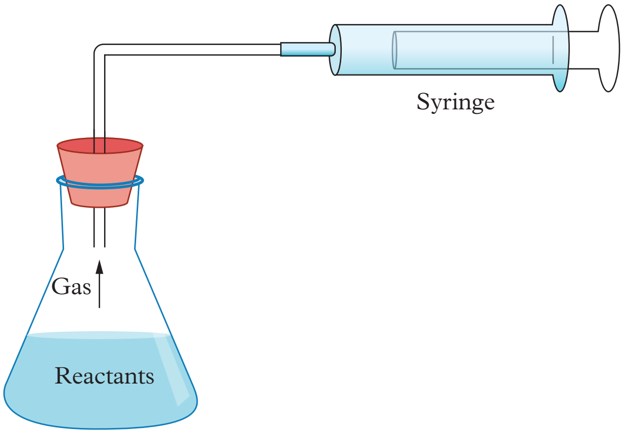


FIGURE 1 Set-up for the gas production experiment

Hydrogen gas produced is captured as it fills the syringe.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

The results are shown in Table 1. Note: The laboratory is located at sea level at a pressure of 100 kPa and HCl is the limiting reactant.

TABLE 1 Results of the gas production experiment

Trial	Temperature (°C)	Volume of gas produced (mL)
Day 1	25.7	480
Day 2	22.4	472
Day 3	26.2	489

Apply understanding

1 **Calculate** the theoretical volume of the gas produced at STP. (4 marks)

2 **Calculate** the average volume of gas produced in the reaction. (1 mark)

Analyse data

3 **Identify** any differences between the theoretical and average experimental values. (1 mark)

4 **Organise** the data into a graphical format. (2 marks)

Evaluate evidence

5 **Justify** whether the data is accurate. (2 marks)

6 **Draw a conclusion** about the relationship between temperature and volume of gas produced. (1 mark)

Module 11 checklist: Gases

Quizlet: Revise key terms online to test your understanding

UNIT 2 Topic 1 review

Multiple choice

(1 mark each)

1 Element X is in group 14 of the periodic table. Element Y is in group 16. The most likely formula for a compound formed between X and Y is

  A XY.

  B X_2Y.

  C XY_2.

  D X_2Y_3.

2 Consider the following diagrams:

In a fluorine molecule, there are

  A six non-bonding electron pairs and one bonding electron pair.

  B eight non-bonding electron pairs and one bonding electron pair.

  C twelve non-bonding electron pairs and one bonding electron pair.

  D six non-bonding electron pairs and two bonding electron pairs.

3 A knowledge of electronegativity, dipoles and molecular shape will lead to the conclusion that ammonia contains

  A dipoles and is polar.

  B dipoles but is non-polar.

  C no dipoles and is non-polar.

  D covalent bonds but no dipoles.

4 Carbon dioxide (CO_2) is

  A bent with polar bonds and is polar overall.

  B linear with polar bonds and is polar overall.

  C bent with two polar bonds but is non-polar overall.

  D linear with two polar bonds but is non-polar overall.

5 When comparing the three molecules H_2, H_2O and H_2S, the likely order of boiling points, from lowest to highest, is

  A H_2, H_2O and H_2S.

  B H_2, H_2S and H_2O.

  C H_2O, H_2 and H_2S.

  D H_2S, H_2 and H_2O.

6 Sample components that have a higher R_F will have

  A greater desorption from the stationary phase into the mobile phase.

  B greater desorption from the stationary phase and desorption into the mobile phase.

  C greater adsorption to the stationary phase and less desorption into the mobile phase.

  D greater adsorption to the stationary phase and greater adsorption into the mobile phase.

Use the following information to answer questions 7 to 9.

Figure 1 was obtained from paper chromatography.

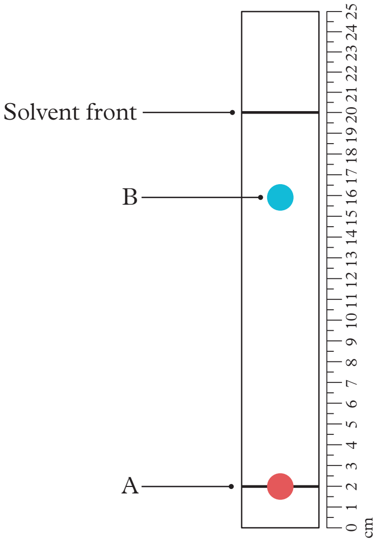


FIGURE 1 Chromatogram obtained from paper chromatography

7 Calculate the R_F of sample component B.

  A 0.11

  B 0.10

  C 0.80

  D 0.78

8 Which statement is true of the affinity that A and B have to the stationary phase?

  A A has a higher affinity to the stationary phase.

  B B has a higher affinity to the stationary phase.

  C Both components are equally attracted to the mobile phase.

  D Both components are equally attracted to the stationary phase.

9 A student suspects that the spot labelled A is composed of more than one substance. What change could they make to improve the separation?

A Increasing the affinity of A to the mobile phase
B Increasing the affinity of B to the mobile phase
C Increasing the affinity of A to the stationary phase
D Increasing the affinity of B to the stationary phase

10 A polar stationary phase was used in thin layer chromatography, with hexane as the solvent. Samples A–C were run using these phases to generate the following chromatogram.

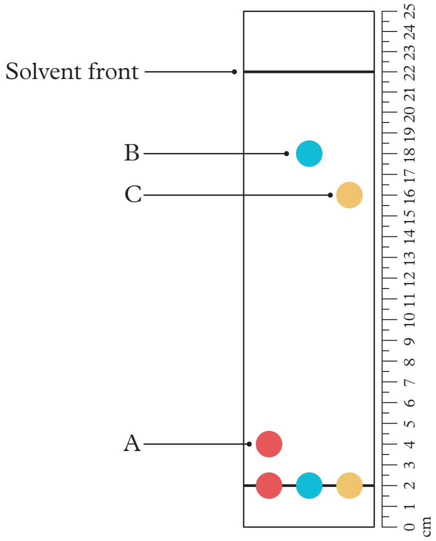


FIGURE 2 Chromatogram

Which sample has the least polar molecules?

A Sample A
B Sample B
C Sample C
D There is not enough information to determine which sample has the most polar molecules.

11 Rebreather apparatus used by scuba divers removes carbon dioxide by reacting it with strong bases, such as a mixture of calcium hydroxide and sodium hydroxide. The overall reaction can be represented in this equation:

$$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$$

Typically, a minimum of 0.25 L of carbon dioxide is produced per minute at 25°C and 100 kPa. The mass of calcium hydroxide required to remove this amount of carbon dioxide is

A 0.030 g
B 0.748 g
C 0.594 g
D 8.917 g

12 In a reaction between hydrochloric acid and aluminium sulfide, a highly toxic gas hydrogen sulfide is formed according to the following equation

$$Al_2S_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2S(g)$$

An excess of hydrochloric acid is added to 0.200 mol of aluminium sulfide. What is the volume of hydrogen sulfide produced at standard temperature and pressure (STP)?

A 1.65 L
B 14.9 L
C 1.51 L
D 13.6 L

13 An aerosol can with a volume of 200.0 mL above the liquid contents contains 4.20 g of propane, C_3H_8, gas as a propellant. The can exploded one day in a hot car when the temperature reached 58°C. What pressure had been reached inside the can at this temperature?

A 1.31×10^3 kPa
B 2.29×10^2 kPa
C 1.01×10^4 kPa
D 1.01×10^1 kPa

14 A soft-drink manufacturer leaves 25 mL of head space between the top of the soft-drink and the lid. The pressure of carbon dioxide gas in this space is 225 kPa at 15°C. The amount of carbon dioxide, in moles, in this space is

A 45.1.
B 2.35.
C 0.0451.
D 0.00235.

15 Two identical flasks are prepared at standard temperature and pressure (STP). Flask A contains 4.987 g of N_2 gas and flask B contains 14.400 g of an unknown gas. What is the identity of the gas in flask B?

A H_2
B SO_2
C HBr
D C_4H_{10}

Short response

16 **Explain** the difference in boiling point between octane (125.7°C) and water (100°C), with reference to their structures and intermolecular interactions. (4 marks)

17 **Contrast** the strengths of the intermolecular forces: dipole–dipole attractions, hydrogen bonding and dispersion forces. (3 marks)

18 **Consider** the molecules; H_2S, NH_3 and CH_4.

- a **Use** the VSEPR model to draw the molecules and **determine** their shape. (3 marks)
- b **Identify** each of the molecules as polar or non-polar. (3 marks)
- c **Recognise** the intermolecular forces experienced by the molecule when it interacts with identical molecules. (3 marks)
- d **Justify** which molecule has the highest boiling point. (3 marks)

19 In an experiment to determine the vapour pressure of a series of molecules, the following data was obtained. The units for vapour pressure in Figure 3 are millimetres of mercury (mmHg).

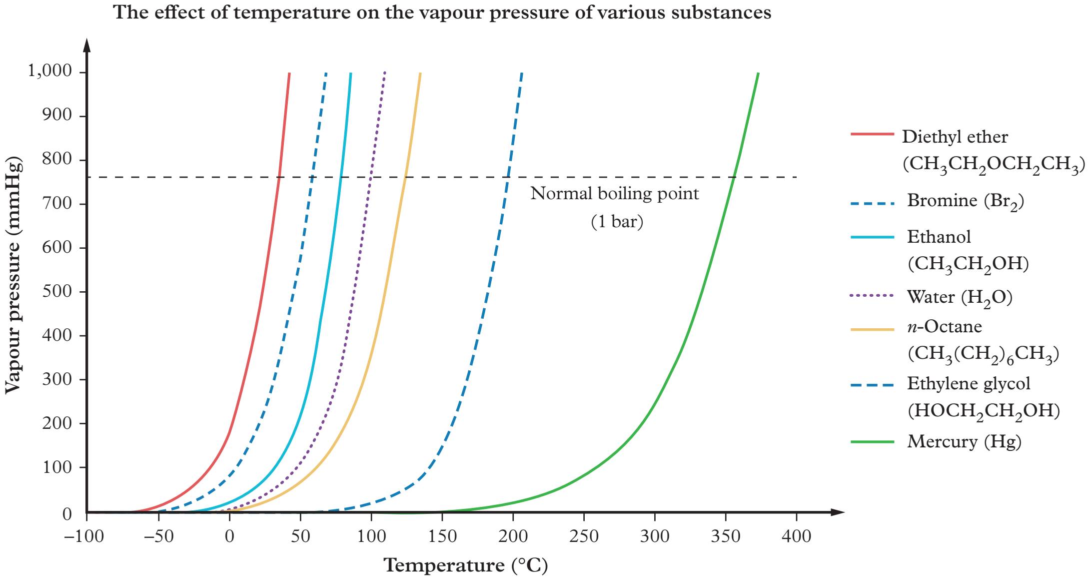


FIGURE 3 Effect of temperature on the vapour pressure of substances

- a **Explain** what vapour pressure is and how it is affected by the strength of the intermolecular forces. (3 marks)
- b **Consider** the vapour pressure of diethyl ether.
  - i **Determine** whether diethyl ether is polar or non-polar, and what intermolecular forces it experiences. (2 marks)
  - ii Based on the presence of the oxygen in the molecule, **explain** whether the vapour pressure of diethyl ether in the graph is what you expected. (2 marks)
- c **Determine** whether octane and water are polar or non-polar, and what intermolecular forces they experience. (2 marks)
  - i **Explain** the difference in vapour pressure between water and octane. (3 marks)

20 Figure 4 demonstrates the trends in the boiling points of group 14 to 17 hydrides.

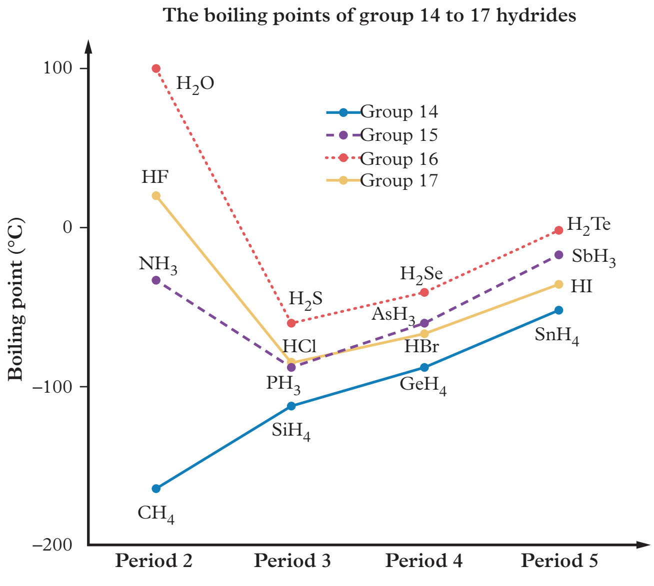


FIGURE 4 Trends in the boiling points of group 14 to 17 hydrides

- **a** **Consider** water (H_2O) and hydrogen sulfide (H_2S). Both oxygen and sulfur are in group 16 on the periodic table.
  - **i** **Identify** the shape of the molecules. (1 mark)
  - **ii** **Determine** whether the molecules are polar or non-polar. (1 mark)
  - **iii** **Determine** the type/s of intermolecular forces that the molecules interact through. (2 marks)
  - **iv** **Explain** the difference in boiling points of the molecules using the data from the graph. (3 marks)
- **b** **Consider** water (H_2O) and hydrogen fluoride (HF).
  - **i** **Describe** the shape of the molecules. (1 mark)
  - **ii** **Determine** whether the molecules are polar or non-polar. (1 mark)
  - **iii** **Determine** the type/s of intermolecular forces that the molecules interact through. (2 marks)
  - **iv** **Explain** the difference in boiling points of the molecules using the data from the graph. (3 marks)
- **c** **Consider** ammonia (NH_3) and methane (CH_4).
  - **i** **Describe** the shape of the molecules. (1 mark)
  - **ii** **Determine** whether the molecules are polar or non-polar. (1 mark)
  - **iii** **Determine** the type/s of intermolecular forces that the molecules interact through. (2 marks)
  - **iv** **Explain** the difference in boiling points of the molecules using the data from the graph. (3 marks)

21 **Explain** how chromatography can be used to separate the components of a sample to determine its composition and purity. (8 marks)
In your response, you must refer to:

- the mobile and stationary phases, including their properties
- the affinity to the two phases
- the strength of intermolecular interactions
- how conclusions can be drawn on composition and purity.

22 The data in Table 1 was obtained from a separation of three ink samples conducted using paper chromatography. The mobile phase used was 10% ethanol in water.

TABLE 1 Data of three ink samples

Sample	Component	R_F
1	Blue	0.56
1	Yellow	0.76
2	Blue	0.32
2	Red	0.91
3	Green	0.23
3	Orange	0.23

- **a** **Sketch** a chromatograph of the resulting separation. (3 marks)
- **b** **Explain** which component has a greater affinity to the mobile phase. (2 marks)
- **c** **Explain** which component has a greater affinity to the stationary phase. (2 marks)
- **d** Two blue components are observed. **Justify** whether they are the same dye. (2 marks)
- **e** **Evaluate** whether the separation is valid and **describe** how it can be improved. (2 marks)

23 A quality control chemist is required to determine the composition and purity of three dyes used in various foods. Using a thin layer chromatography (TLC) plate, the samples are separated using a water mobile phase. Figure 5 demonstrates the results.

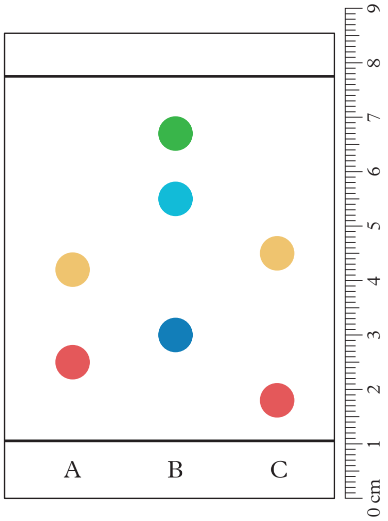


FIGURE 5 Composition and purity of three dyes

- **a** Before they are run on the TLC plate, samples A and C look to be the same colour. After separation, more detail is identified. **Explain** whether the two samples are the same and **justify** your answer using the R_F values of the yellow and red dyes. (6 marks)
- **b** **Explain** which component of the three samples has a greater affinity to the mobile phase. (2 marks)
- **c** **Explain** which component of the three samples has a greater affinity to the stationary phase. (2 marks)
- **d** **Evaluate** whether the identity of the components of the dyes can be determined using the separation shown. (3 marks)
- **e** A food chemist knows that each of the dyes should contain only two components. **Comment** on the purity of each of the three samples. (2 marks)

24 **Use** the kinetic theory of gases to **explain** the relationship between

- **a** temperature and volume in variable volume systems (2 marks)
- **b** temperature and pressure in fixed volume system (2 marks)
- **c** number of particles and pressure in a variable volume system (2 marks)
- **d** number of particles and pressure in fixed volume system (2 marks)
- **e** volume and pressure in a variable volume system. (2 marks)

25 A 2.5 L container contains pure nitrogen gas (N_2) at STP.

- **a** **Calculate** the amount, in mol, of the gas using molar volume. (2 marks)
- **b** **Calculate** the amount, in mol, of the gas using the ideal gas equation. (2 marks)
- **c** **Comment** on any differences in the values that you have calculated in parts **a** and **b**. (1 mark)
- **d** **Explain** why molar volume is used at STP. (2 marks)

26 Chlorine gas can be produced by the combustion of hydrogen chloride gas (HCl) where oxygen is in excess. The equation for the combustion is:

$$4HCl(g) + O_2(g) \rightarrow 2H_2O(l) + 2Cl_2(g)$$

- **a** **Calculate** the volume of Cl_2 produced from burning 500 mL of HCl gas at STP. (4 marks)
- **b** **Calculate** the volume of O_2 consumed in the process at STP. (2 marks)

27 A tank of biogas collected at a sewerage treatment plant had a total volume of 80,000 L at 18°C and 105 kPa. The biogas consisted of 90% CH_4.

- **a** **Calculate** the mass of methane. (4 marks)
- **b** When methane is combusted, 890.6 kJ of energy is released per mole of methane. **Use** this information to **construct** a balanced thermochemical equation for the combustion of methane. (2 marks)
- **c** **Calculate** the mass of CO_2 generated by the complete combustion of the tank. (2 marks)
- **d** **Calculate** the energy generated by the complete combustion of methane in the tank. (2 marks)

28 In an experiment, a student wishes to determine the relationship between temperature and pressure in a fixed volume system. They set up the experiment as shown in Figure 6, where three identical flasks are placed in a beaker of water at room, medium and hot temperatures.

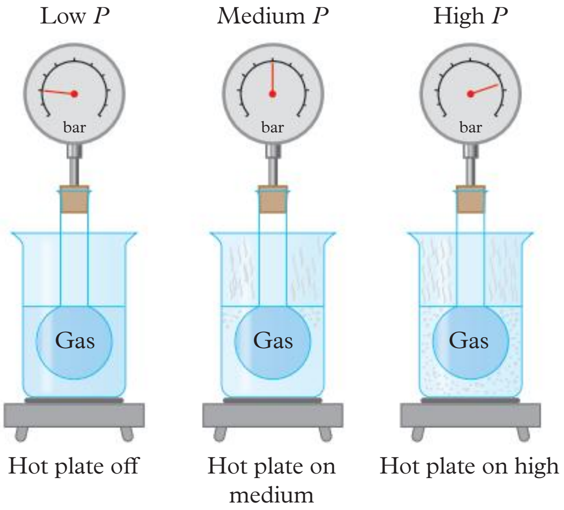


FIGURE 6 Three flasks placed in a beaker of water at different temperatures

- **a** **Identify** the independent variable. (1 mark)
- **b** **Identify** the dependent variable. (1 mark)
- **c** **Generate** a hypothesis for the reaction. (3 marks)

The results of the experiment are summarised in Table 2.

TABLE 2 Experiment results

Temperature	Pressure
Room	Low
Medium	Medium
Hot	High

- **d** **Justify** whether the experiment is qualitative or quantitative. (2 marks)
- **e** **Describe** the relationship between the temperature and pressure of gases and **explain** the observations using the kinetic theory of gases. (2 marks)
- **f** **Evaluate** whether the experiment is valid. (2 marks)
- **g** **Propose** two ways in which the experiment could be improved to generate more accurate and precise data. (2 marks)

TOTAL MARKS

/147

MODULE 12

Aqueous solutions and molarity

Introduction

Water is one of the most abundant and important molecules on Earth. Approximately 71% of Earth's surface is covered in water, with more than 96% of that water held in the oceans and seas. This is salt water so it is not suitable for human consumption unless the salt is removed. Incredibly, the remaining 4% (or just under) is all of the freshwater used to sustain many of the lifeforms on our planet. Of this 4%, approximately one-third is trapped in the polar ice caps.

As water is one of the most essential molecules on Earth, it is important to recognise its structure and properties.

In 1766, British chemist and physicist Henry Cavendish noted that water formed when various gases in the atmosphere reacted with one another. Before this, water was believed to be an element, but Cavendish demonstrated that the reaction involved two different elements and therefore water must be a compound. Cavendish did not know at that time that the elements involved were hydrogen and oxygen.

French chemist Antoine Lavoisier, the "father of modern chemistry", built on Cavendish's work to name hydrogen, which derived from the Greek *hydro*, meaning "water", and *gen*, meaning "something produced". Therefore, hydrogen was named for its ability to form water.

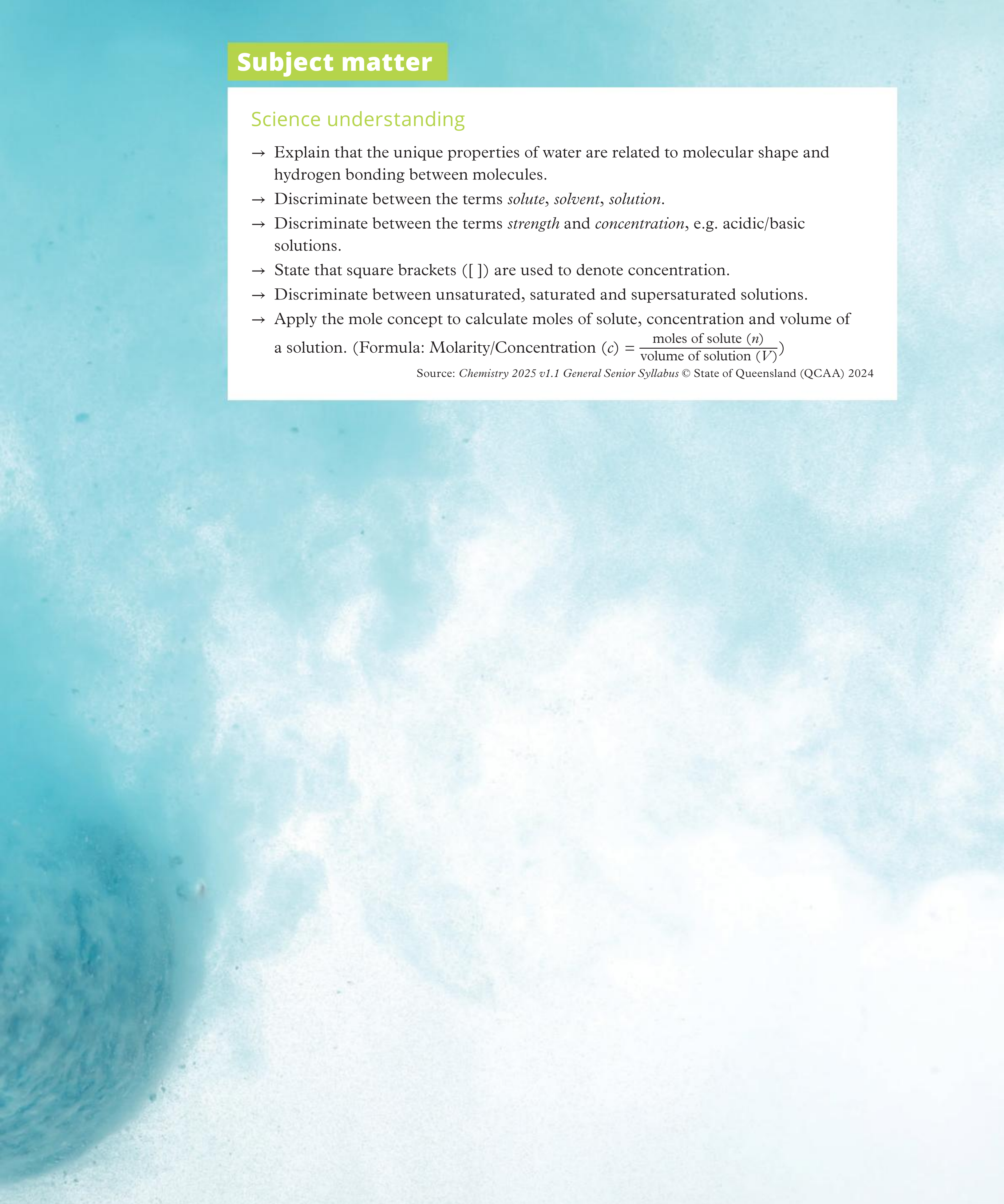
Not long after this, Cavendish hypothesised that water must be two parts hydrogen and one part oxygen. This was experimentally confirmed several years later when German chemist Johann Wilhelm Ritter passed an electric current through water, in a process called electrolysis. He was able to generate hydrogen gas, H_2, at one electrode and oxygen gas at the other. Because there was twice as much hydrogen as oxygen, O_2, he concluded that water must be made of two parts hydrogen and one part oxygen.

Since then, water's behaviour has been researched extensively and scientists have discovered much about its properties and its involvement in reactions. This module explores some of water's properties and its role in the chemistry of aqueous solutions.

Prior knowledge

Prior knowledge quiz

Check your understanding of intermolecular forces and solutions before you start.



Subject matter

Science understanding

- → Explain that the unique properties of water are related to molecular shape and hydrogen bonding between molecules.
- → Discriminate between the terms *solute*, *solvent*, *solution*.
- → Discriminate between the terms *strength* and *concentration*, e.g. acidic/basic solutions.
- → State that square brackets ([ ]) are used to denote concentration.
- → Discriminate between unsaturated, saturated and supersaturated solutions.
- → Apply the mole concept to calculate moles of solute, concentration and volume of a solution. (Formula: Molarity/Concentration $(c) = \frac{\text{moles of solute } (n)}{\text{volume of solution } (V)}$)

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Lesson 12.1

Properties of water

Key ideas

→ Water has a bent shape, making it a polar molecule with a dipole or a dipole moment. This allows it to form hydrogen bonds with other water molecules.

→ Water's ability to form strong hydrogen bonds with other water molecules gives rise to its unique properties.

Learning intentions and success criteria

What is the polarity and shape of water?

Water is a unique and incredibly useful compound. You should recall from Module 9 that molecules have shapes based on the electrons contained within their valence shells. Oxygen has two lone pairs of electrons when it bonds covalently to two hydrogen atoms. This means that the water molecule adopts a bent structure due to the lone pairs of electrons on the oxygen, as seen in Figure 1.

Oxygen is electronegative, so it pulls on the electrons in the covalent bonds, meaning that they spend more time closer to the oxygen nucleus. This means that in a O–H bond, the negative charge is closer to the oxygen atom than the hydrogen atom. Because water is bent, a dipole is formed in which oxygen has a partial negative charge and each hydrogen has a partial positive charge. The presence of partial charges arranged asymmetrically in a molecule makes it a polar molecule.

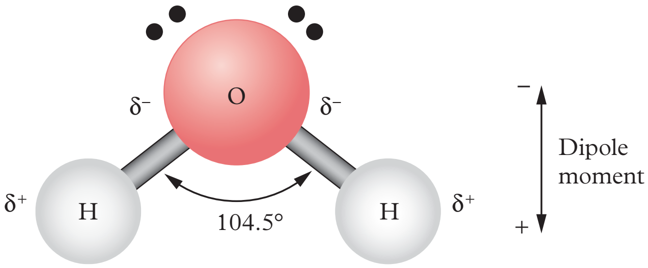


FIGURE 1 The bent shape of a water molecule (H_2O) and resulting polarity due to the electronegativity of oxygen

What intermolecular forces does water form?

Hydrogen bonds form by the electrostatic attraction between a partially positive hydrogen atom and a partially negative oxygen, nitrogen or fluorine atom. In water, the partial positive charge on the hydrogen of one molecule is attracted to the partial negative charge on the oxygen of a different molecule (Figure 2), forming a hydrogen bond.

What are the properties of water?

The shape, polarity and intermolecular forces can be used to explain the properties of water, such as the:

- boiling point of water
- density of water in the solid and liquid states
- surface tension of water.

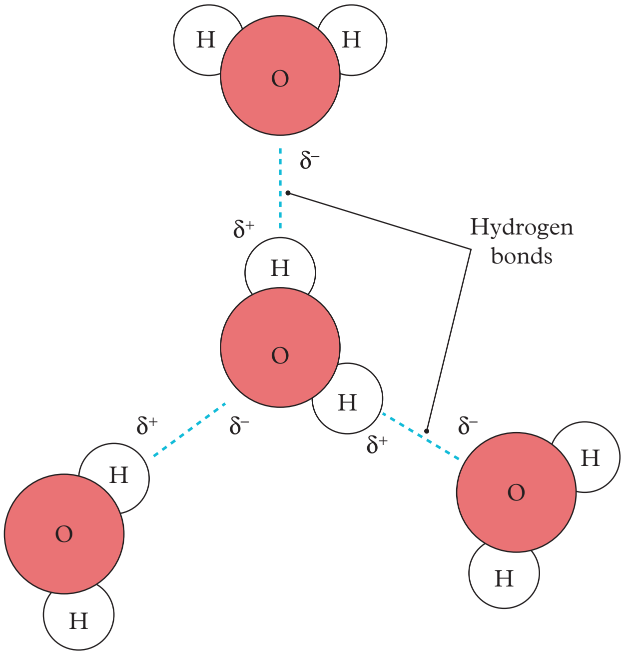


FIGURE 2 Hydrogen bonding between water molecules

Boiling point

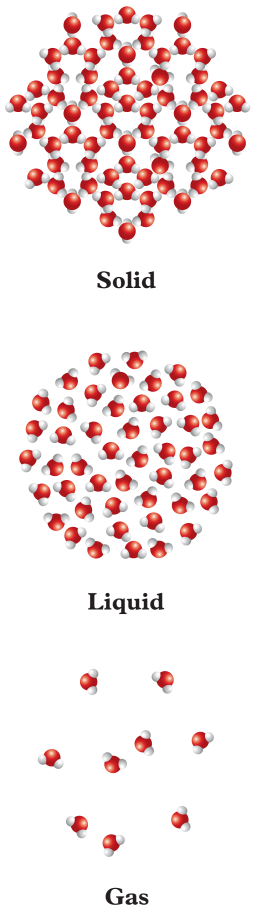


FIGURE 3 The interaction of water particles in solid, liquid and gaseous states

The boiling point of water is the temperature at normal atmospheric pressure at which it changes state from a liquid to a gas by forming bubbles under the surface. In different states, the particles exhibit different behaviours (Figure 3).

- In the solid state, particles experience so little movement that they only vibrate. They pack tightly together because they are strongly attracted to one another and the solid maintains its shape as a result of the strength of this attraction. The force or energy of their attraction to one another is greater than the energy of their movement.
- In the liquid state, particles pack closely together. However, unlike in the solid state, they move readily around one another. This is why a liquid fills the space at the bottom of a container. In other words, the particles have enough energy to move, but not enough energy to separate from one another.
- In the gaseous state, particles have more energy than in the liquid phase. The attractions that held them together as a liquid are overcome by the energy involved with their movement. They travel in straight lines until they reach an object that they bounce off and change direction. In other words, they have too much energy to stay together.

For water to change from a liquid to a gas, the particles must increase their movement to the point where the kinetic energy involved with their movement is greater than the energy that attracts – through dispersion forces and hydrogen bonding – one water molecule to another.

As temperature is a measure of the average kinetic energy of particles, boiling point signifies the temperature at which particles have enough energy to overcome intermolecular forces and transition between a liquid and a gas. For water, this temperature is 100°C or 373 K at atmospheric pressure.

Density

Density is a measure of the mass of a substance per unit of volume and can be demonstrated by the equation:

$$\rho = \frac{m}{V}$$

where ρ is density (g mL^{-1}), m is mass (g) and V is volume (mL).

The density of water is $1.00\,\text{g mL}^{-1}$ at 4°C, meaning that 1.00 g of water occupies 1 mL of volume. At room temperature, the density is slightly less (around $0.99\,\text{g mL}^{-1}$). Volume can be measured in millilitres (mL), centimetres cubed (cm^3), litres (L), decimetres cubed (dm^3) or metres cubed (m^3). You can use Figure 4 to help you convert between units of volume.

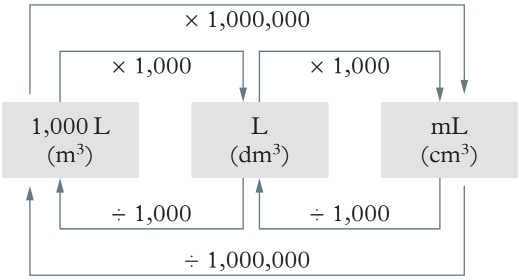


FIGURE 4 Unit conversions of volume

The greater the mass in a fixed volume, the greater the density. When water freezes to become solid ice, the water particles align to leave spaces within its structure (Figure 5). This means that 1 L of ice has a lower mass than 1 L of liquid water and liquid water has a greater density than solid water. This is why ice floats on water. Water is exceptional because most compounds are more dense as solids than as liquids.

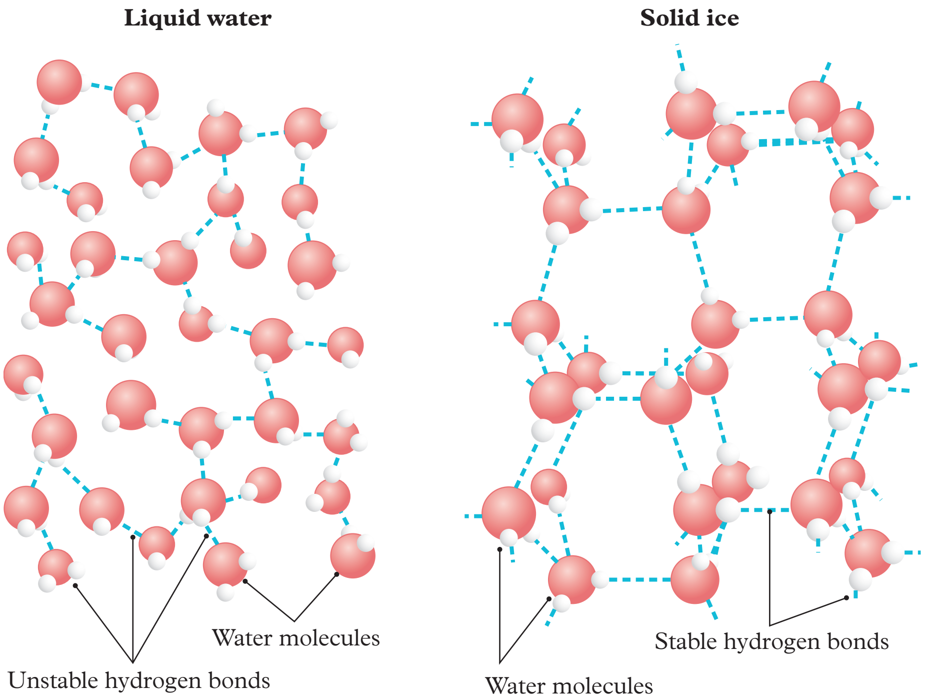


FIGURE 5 There are fewer and less stable hydrogen bonds between water molecules in liquid water compared to solid ice.

Study tip

Remember that the properties of molecules depend on the atoms they are made of, the shape of the molecule and its intermolecular forces.

surface tension the ability of the surface of a liquid to resist an external force

Surface tension

Surface tension is the ability of the surface of a liquid to resist an external force. As water molecules experience hydrogen bonding – the strongest of all intermolecular forces – the molecules are held together strongly.

Surface tension has various effects.

- **Beading**: Figure 7A shows water beading on a leaf's surface. The outer surface of a leaf consists of a waxy material, made of lipids (fats). Because water is polar and the surface of the leaf is non-polar, the water molecules do not interact with it strongly; rather, they tend to interact with each other. Instead, the water rolls off the leaf in a droplet or bead.
- **Flotation**: Water molecules are strongly attracted to one another because of the strength of intermolecular hydrogen bonding. Objects that have a greater density than water can float on its surface (Figure 7B). This only occurs if the material that the object is made from does not interact with water. The object must also be able to distribute its surface area across the surface of the water.
- **Dripping**: Water forms droplets. For example, water adheres to the surface of a tap. As the mass of the water increases, it overcomes the strength of the hydrogen bonding that is holding the drop to the tap. Therefore, the mass of the water drop under the force of gravity becomes greater than the forces holding the water together and to the tap, so it falls (Figure 7C).

FIGURE 6 Ice floats on water because ice is less dense than liquid water.

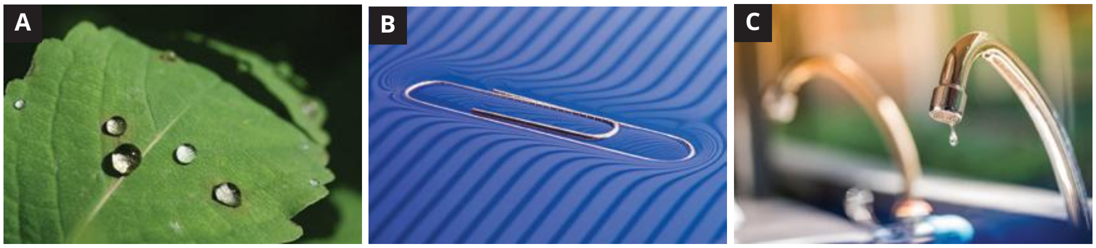


FIGURE 7 The surface tension of water produces various effects. (A) Water on a leaf demonstrates beading. (B) A paperclip on water demonstrates flotation. (C) Droplets falling from a tap demonstrate dripping.

What is a solution?

A solution is the name given to the mixture resulting when a solute dissolves in a solvent. To prepare a solution, a sample of solute is measured (either by weighing on a balance or measuring a volume), and then added to a volume of solvent. The solution can be stirred or heated to enable the solute to dissolve faster. A solution is homogeneously mixed and the solute has broken into individual atoms, molecules or ions.

Ionic solids dissociate into their component ions. For example, an aqueous solution of sodium nitrate, $NaNO_3$ consists of Na^+ and NO_3^- ions surrounded by water molecules. An aqueous solution of sucrose molecules (table sugar) has molecules of sucrose surrounded by and interacting with water molecules.

A solution in which water acts as the solvent is called an **aqueous solution** (from the word *aqua*, meaning "water"). It is given its own state in chemical equations, represented as (aq). Water is often referred to as the universal solvent as it can dissolve more substances (solutes) than any other solvent. If a chemical is referred to as a solution of something, it can be assumed that it has been dissolved in water and it has an aqueous (aq) state. However, it is important to note that this is not the case all of the time.

aqueous solution
a mixture of a solute dissolved in water

Non-polar substances, such as fats will not dissolve in water and therefore no fat can form an aqueous solution.

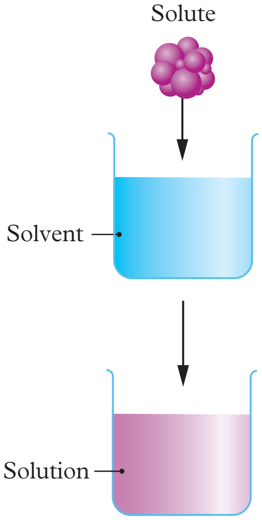


FIGURE 8 A solute dissolves in a solvent to form a solution.

Worked example 12.1A

Identifying the components of a solution

Worked example 12.1A: Watch a video that shows how to solve this problem.

In seawater, **identify** the:

a main solute (1 mark) **b** solvent (1 mark) **c** solution. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Identify" means to locate, recognise and name. We need to recognise the components of a solution. The questions are worth 1 mark each, so we must name the components correctly.
Step 2: For part **a**, recognise that the solute is the substance that is dissolved.	Main solute: sodium chloride, NaCl (1 mark)
Step 3: For part **b**, recognise that the solvent is what the solute is dissolved in.	Solvent: water (1 mark)
Step 4: For part **c**, recognise that the solution is the combination of the solute and solvent.	Solution: salt water or seawater (1 mark)

Your turn

You make a cup of coffee using instant coffee powder with no milk or sugar. **Identify** the:

a solute (1 mark) **b** solvent (1 mark) **c** solution. (1 mark)

Real-world chemistry

Formation of snowflakes in the atmosphere

Snowflakes are incredibly beautiful and unique. No two snowflakes are identical, and the various formations look amazing under a microscope.

All air contains water vapour. This is gaseous water that has evaporated from the land or bodies of water and into the atmosphere. Air can hold a finite amount of water vapour; however, hotter air can hold more water vapour. As air cools, the water vapour condenses into droplets and precipitates as rain.

Humidity is a measure of the concentration of water vapour in the atmosphere and depends on temperature. At 100% humidity, air is saturated. However, air humidity can exceed 100%, in which case it is said to be supersaturated. Supersaturation only happens when the temperature of the atmosphere drops and air still contains the water that was vaporised at the higher temperature.

In order to form snowflakes, the atmosphere must be supersaturated with water and super cool (0°C or below). The air must also contain pollen or dust. The crystals begin to form on the pollen or dust particles, which starts the snowflake forming in a process called "seeding".

FIGURE 9 Snowflake

Apply your understanding

1 **Justify** the observation that hotter air can hold more water vapour than cooler air. (4 marks)

2 **Explain**, using kinetic molecular theory, why water condenses and precipitates when cooled. (4 marks)

3 Using the molecular shape of water and the intermolecular forces it forms, **explain** why water has a high boiling point. (3 marks)

Check your learning 12.1

Check your learning 12.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the shape of the water molecule. (1 mark)

2 **Explain** why a water molecule is polar and forms hydrogen bonds. (3 marks)

3 **Use** your understanding of the intermolecular bonding of water to **explain** why water expands when it freezes. (3 marks)

Analytical processes

4 A small pellet (a sphere of radius 1.4 cm) of osmium has the same mass as a beryllium sphere the size of a tennis ball. **Justify** how this is possible. (2 marks)

Knowledge utilisation

5 **Hypothesise** two other substances which would have high surface tensions. **Justify** your answer using the intermolecular forces present in these substances. (4 marks)

Lesson 12.2

Concentration

Learning intentions and success criteria

Key ideas

→ Concentration is a measure of the number of particles in a specific volume of solution.

→ Concentration can be measured in many different units depending on the nature of the solution. Molar concentration (molarity) can be calculated if the amount (in moles) and volume of the substance are known.

→ Dilution involves the addition of solvent to reduce the concentration of a solution.

What is concentration?

Concentration is a measure of the amount of substance in an accurately known volume of solution. A **concentrated solution** contains a higher amount of substance, whereas a **dilute solution** contains a lower amount of substance in an equivalent volume. Concentration is measured in many different ways, depending on the nature of the solute and the solvent.

concentration
a measure of the number of particles in a specific volume of solution; units of concentration include mol L^{-1} (M), g L^{-1} and mg L^{-1} (ppm)

concentrated solution
a solution containing many particles per unit volume of solution

dilute solution
a solution containing few particles per unit of volume

unsaturated
a solution that can dissolve more solute

saturated
a solution in which no more solute can dissolve

supersaturated
a solution that has been heated to dissolve more solute than can be dissolved under equilibrium conditions; on cooling, the solute remains in solution in a metastable state, until a disturbance (seed crystal, speck of dust, tapping the container) starts the crystallisation process

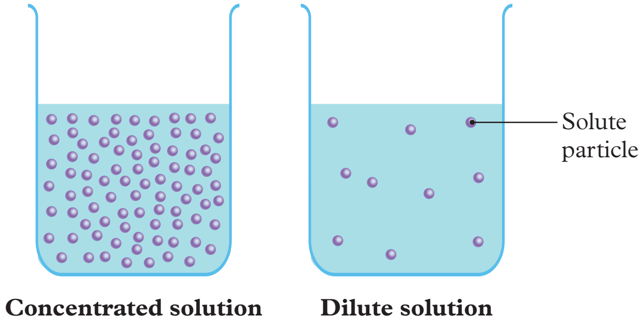


FIGURE 1 A concentrated solution contains a higher amount of solute than a dilute solution.

Saturation

Sometimes, when dissolving a solute in a solvent, you get to the point where no more will dissolve. If you dissolve sugar into water, one teaspoon at a time, you will get to a point where it will stop dissolving and will settle on the bottom of the container. In these cases, we use different terms to do with saturation to describe the ability of the solute to continue to dissolve or not.

- If all the solute dissolves and it is possible to dissolve more, the solution is **unsaturated**.
- If no more solute will dissolve, the solution is **saturated**.
- If a saturated solution is heated, more solute often dissolves. If the solute remains in solution as it cools, this is a **supersaturated** solution and is often used to produce crystals of the solvent.

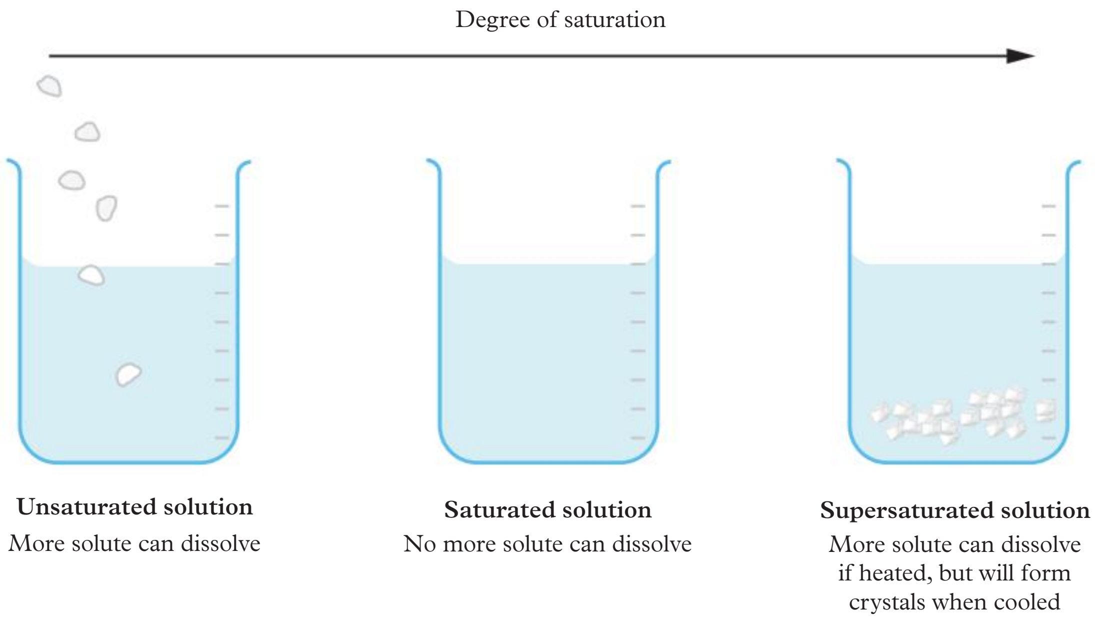


FIGURE 2 The degree of saturation increases as concentration increases.

Skill drill

Analysing and evaluating solubility data

Science inquiry skill: Processing and analysing data (Lesson 1.7) and Evaluating evidence (Lesson 1.8)

A student is required to determine the maximum amount of copper(II) sulfate that will dissolve in 200 g of water at 20°C, 30°C, 40°C and 60°C. To perform the experiment, the student places the beaker with the 200 g of water onto a balance and tares the balance. They then continually add copper(II) sulfate, one gram at a time, with stirring, until no more will dissolve. The mass recorded is the mass that dissolved before no more would, and copper(II) sulfate remained at the bottom of the beaker.

Practise your skills

1 **Identify** the independent and dependent variables. (2 marks)

2 **Sketch** a scatterplot of the data from Table 1. (2 marks)

3 **Explain** the trend in the data, with reference to the variables you identified in question **1**. (3 marks)

4 **Evaluate** whether the experiment is valid. (2 marks)

TABLE 1 Mass of copper(II) sulfate dissolved in water at different temperatures

Temperature (°C)	Mass of $CuSO_4$ in 200 g of water (g)
20	24
30	34
40	42
60	70

What is molarity (mol L^{-1} or M)?

A number of units are used to measure concentration. The most common unit of concentration used by chemists is the number of moles of a substance per litre of solution. This is called **molarity**, or molar concentration, and is represented in units of moles per litre, mol L^{-1} or M. Any of these units can be used to calculate concentration.

molarity
molar concentration, units are moles per litre, mol L^{-1} or M

Molar concentration, or molarity, is calculated as:

$$c = \frac{n}{V}$$

or

$$n = c \times V$$

where c is concentration (mol L^{-1} or M), n is number of moles (mol) and V is volume of the final solution (L). If volume is given in mL, then it must be converted to L first.

The concentration of a substance can also be represented by placing its formula in square brackets. A solution of 0.8 M KI would be represented as [KI] = 0.8 M.

In a laboratory situation, it is usual to weigh a mass of a chemical and dissolve it in a smaller volume of water, make the volume of solution up to the required final volume, then calculate its concentration. This calculation is a multistep process, where moles are calculated first, then concentration.

Often, scientists want to determine the concentration of ions in a solution. This is essential when calculating the pH of a solution (covered in Module 14 and later in Unit 3). To determine the concentration of ions, an extra step must be added in which the number of moles of the ion is determined from the number of moles of the solute. This requires a mole ratio statement (Step 4 in Worked example 12.2D).

Study tip

Do not confuse M (mol L^{-1}) (the unit of concentration) with M (the symbol for molar mass). You can use the symbol M_r for molar mass if this helps to distinguish the terms.

Study tip

Multistep calculations in chemistry can get quite messy, so it is essential that you develop good techniques early. Highlight the important information (numbers, units, chemicals or key words), and then add the letters that represent them (e.g. put the letter n above mole values or m above mass values). Last, put a question mark next to the value you are asked to calculate. Then look at all of the formulas and chemical equations, and find a pathway that will allow you to use the values given to calculate the answer.

Worked example 12.2A

Calculating molar concentration from moles and volume

Worked example 12.2A: Watch a video that shows how to solve this problem.

A 200 mL solution contains 1.2 mol of calcium chloride ($CaCl_2$). **Calculate** its concentration in mol L^{-1}. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine molar concentration. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$c = \frac{n}{V}$ c = ?, n = 1.2 mol, V = 200 mL
Step 3: Substitute the known values into the formula and solve for the unknown value.	$[CaCl_2] = \frac{1.2\,\text{mol}}{\frac{200\,\text{mL}}{1{,}000}}$ (1 mark) $= 6\,\text{M}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	6.0 M (1 mark)

Your turn

A 150 mL solution contains 0.54 mol of sodium hydroxide (NaOH). **Calculate** its molar concentration. (2 marks)

Worked example 12.2B

Calculating molar concentration from mass and volume

Worked example 12.2B: Watch a video that shows how to solve this problem.

A 2.5 g sample of sodium chloride is completely dissolved in water and the solution is then made up to a volume of 56 mL. **Calculate** the concentration (M) of the sodium chloride solution. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine molar concentration. The question is worth 4 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ $n = ?, m = 2.5\,g, M = 58.44\,g\,mol^{-1}$ $c = \frac{n}{V}$ $c = ?, V = 56\,mL$
Step 3: Substitute the known values into the formula and solve for the number of moles.	$n(NaCl) = \frac{2.5\,g}{58.44\,g\,mol^{-1}}$ (1 mark) $= 4.278 \times 10^{-2}$ mol (1 mark)
Step 4: Substitute the known values into the formula and solve for the concentration.	$[NaCl] = \frac{4.278 \times 10^{-2}\,mol}{\frac{56\,mL}{1,000}}$ (1 mark) $= 0.7639\,M$
Step 5: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	0.76 M (1 mark)

Your turn

A 126 g sample of potassium nitrate (KNO_3) is dissolved in water and the solution is made up to 500 mL. **Calculate** the concentration (M) of the solution. (4 marks)

Worked example 12.2C

Calculating volume from mass and molar concentration

Worked example 12.2C: Watch a video that shows how to solve this problem.

A student is required to make a solution of 0.0500 M sodium hydroxide. They find 0.920 g in the chemical store. **Calculate** the final volume of the solution, in mL. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the volume. The question is worth 4 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ $n = ?, m = 0.920\,g, M = 40.0\,g\,mol^{-1}$ $c = \frac{n}{V}$ $c = 0.0500\,M, V = ?$

Think	Do
Step 3: Substitute the known values into the formula and solve for the number of moles.	$n(\text{NaOH}) = \frac{0.92\,\text{m}}{40.00\,\text{g}\,\text{mol}^{-1}}$ (1 mark) $= 2.3 \times 10^{-2}$ mol (1 mark)
Step 4: Substitute the known values into the formula and solve for the concentration.	$V(\text{NaOH}) = \frac{2.3 \times 10^{-2}\,\text{mol}}{0.0500\,\text{M}}$ (1 mark) $= 0.46\,\text{L}$
Step 5: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	460 mL (1 mark)

Your turn

An experiment requires a 0.15 M solution of sodium bromide. Only 3.09 g of NaBr is available. **Calculate** the volume of solution, in mL, that could be made from this mass. (4 marks)

Worked example 12.2D

Calculating particle concentration from mass and volume

Worked example 12.2D: Watch a video that shows how to solve this problem.

1.32 g of sodium phosphate (Na_3PO_4) is completely dissolved in water to make 137 mL of solution. **Calculate** the molar concentration of the sodium ions (Na^+) in the solution. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine molar concentration. The question is worth 4 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ $n = ?$, $m = 1.32\,\text{g}$, $M = 163.94\,\text{g}\,\text{mol}^{-1}$ $c = \frac{n}{V}$ $c = ?$, $V = 137\,\text{mL}$
Step 3: Substitute the known values into the formula and solve for the number of moles.	$n(\text{Na}_3\text{PO}_4) = \frac{1.32\,\text{g}}{163.94\,\text{g}\,\text{mol}^{-1}}$ (1 mark) $= 8.05 \times 10^{-3}$ mol (1 mark)
Step 4: Use a mole ratio statement to determine the number of moles of the ion. Make sure that you keep the previous answer in the calculator; otherwise, rounding may result in an incorrect answer.	One unit of Na_3PO_4 contains 3 Na^+ ions, so the number of moles of Na_3PO_4 must be multiplied by 3. $n(\text{Na}^+) = 3 \times n(\text{Na}_3\text{PO}_4)$ $= 2.4155 \times 10^{-2}$ mol
Step 5: Substitute the known values into the formula and solve for the concentration.	$[\text{Na}^+] = \frac{2.4155 \times 10^{-2}\,\text{mol}}{\frac{137\,\text{mL}}{1{,}000}}$ (1 mark) $= 0.1763\,\text{M}$
Step 6: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	0.176 M (1 mark)

Your turn

A 5.0 g sample of copper nitrate ($Cu(NO_3)_2$) is added to water, making a 250 mL solution. **Calculate** the molar concentration of the nitrate ions (NO_3^-) in the solution. (4 marks)

What is grams per litre ($g\ L^{-1}$)?

Another common unit of concentration is grams of substance per litre ($g\,L^{-1}$). There are several reasons for using $g\,L^{-1}$ instead of $mol\,L^{-1}$. The molar mass of a compound may not be known if it is a new compound and still being investigated or if the substance is impure. It is also used for everyday solutions outside of the chemistry laboratory.

Concentration is calculated as:

$$c = \frac{m}{V} \text{ or } m = c \times V$$

where c is concentration ($g\,L^{-1}$), m is mass (g) and V is volume (L).

This is the same calculation as previously seen, but moles is replaced with mass (in grams).

Study tip

Always convert volume into litres when completing concentration calculations. Remember that the volume is the total volume of the solution after the solute and solvent have been mixed. This is not the same as the volume of water used.

Worked example 12.2E

Calculating concentration in grams per litre

Worked example 12.2E: Watch a video that shows how to solve this problem.

Calculate the concentration, in $g\,L^{-1}$, of a solution of 0.042 g protein dissolved in water to make 241 mL of solution. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine concentration in $g\,L^{-1}$. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$c = \frac{m}{V}$ $c = ?,\ m = 0.042\,g,\ V = 241\,mL$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$[\text{protein}] = \frac{0.042\,g}{\frac{241\,L}{1{,}000}}$ $= 0.1743\,g\,L^{-1}$ (1 mark)
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	$0.17\,g\,L^{-1}$ (1 mark)

Your turn

A 3.46 g sample of magnesium sulfate is dissolved in water, making 200 mL of solution. **Calculate** the concentration of the solution in $g\,L^{-1}$. (2 marks)

Worked example 12.2F

Calculating mass from concentration in grams per litre

Worked example 12.2F: Watch a video that shows how to solve this problem.

A chemist wishes to make 52 mL of a solution of sugar of concentration 0.94 g L^{-1}. **Calculate** the mass of sugar required. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the mass. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$c = \frac{m}{V}$ $c = 0.94\,\text{g L}^{-1}$, $m = ?$, $V = 52\,\text{mL}$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$m(\text{sugar}) = 0.94\,\text{g L}^{-1} \times \frac{52\,\text{mL}}{1{,}000}$ $= 0.04888\,\text{g}$ (1 mark)
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	49 mg or 0.049 g (1 mark)

Your turn

To make a solution with a concentration of 0.5 g L^{-1}, a student only has access to a 250 mL volumetric flask. **Calculate** the mass of solute required to make the solution. (2 marks)

Real-world chemistry

Nutrition information panels

Concentrations are presented as mass per unit of volume on the nutritional panels on the back of food packaging. These concentrations are expressed in grams per 100 grams (for solid foods) or grams per 100 mL (for liquid foods). Fats, proteins and carbohydrates are grouped, and the information refers to the overall mass of these within the product.

Nutrition information

NUTRITION INFORMATION (Average)
SERVINGS PER PACKAGE: 10
SERVING SIZE: 25g(6 Biscuits)

	QUANTITY PER SERVING	% DAILY INTAKE* (PER SERVING)	QUANTITY PER 100 g
ENERGY	493 kJ	5.6%	1,970 kJ
PROTEIN	2.0 g	3.9%	7.9 g
FAT, TOTAL	5.0 g	7.2%	20.0 g
- SATURATED	1.0 g	4.1%	3.9 g
CARBOHYDRATE	15.6 g	5.0%	62.6 g
- SUGARS	0.3 g	0.4%	1.3 g
SODIUM	212 mg	9.2%	848 g

*BASED ON AN AVERAGE ADULT DIET OF 8,700 KJ. ALL VALUES CONSIDERED AVERAGES UNLESS OTHERWISE INDICATED

Nutrition information

NUTRITION INFORMATION (Average)
SERVINGS PER PACK: 10
SERVING SIZE: 200 mL

	QUANTITY PER SERVING	% DAILY INTAKE* (PER SERVING)	QUANTITY PER 100 mL
ENERGY	306 kJ	4%	153 kJ
PROTEIN	1.2 g	2%	0.6 g
FAT, TOTAL	<1 g	<1%	<1 g
- SATURATED	0 g	0%	0 g
CARBOHYDRATE	18.0 g	6%	9.0 g
- SUGARS	16.6 g	18%	8.3 g
DIETARY FIBRE	0.4 g	1%	0.2 g
SODIUM	16 mg	0.7%	8 mg
VITAMIN C	40 mg	100% RDI#	20 mg

*BASED ON AN AVERAGE ADULT DIET OF 8,700 KJ. ALL VALUES CONSIDERED AVERAGES UNLESS OTHERWISE INDICATED

FIGURE 3 The nutritional information of a solid food product (left) and a liquid food product (right)

Apply your understanding

1 **Compare** the units of concentration used for solid and liquid foods. **Explain** why they are expressed differently. (3 marks)

2 **Explain** why the concentrations of substances in foods are expressed as mass per 100 g of mass or mass per 100 mL but NOT moles per 100 g and moles per 100 mL. (2 marks)

What is parts per million (ppm or mg L^{-1})?

The unit parts per million (ppm) is used to measure very low concentrations. To fully understand ppm, consider the density of water, which is 1.00 g mL^{-1}. This means that a millilitre of water has a mass of 1.00 g. As there are 1000 mL of water in a litre, 1 L of water has a mass of 1.00 kg.

One-millionth of 1 kg is 1 mg. Therefore, one-millionth of a litre is also 1 mg. So in water, 1 ppm is equivalent to 1 mg per litre of water (mg L^{-1}). Concentration is calculated as:

$$c = \frac{m}{V} \text{ or } m = c \times V$$

where c is concentration (mg L^{-1} or ppm), m is mass (mg) and V is volume (L).

Worked example 12.2G

Calculating concentration in ppm

Worked example 12.2G: Watch a video that shows how to solve this problem.

125 mL of a sodium chloride solution was made from 0.0024 g of the salt. **Calculate** its concentration in ppm. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine concentration in ppm. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$c = \frac{m}{V}$ c = ?, m = 0.0024 g, V = 125 mL
Step 3: Substitute the known values into the formula and solve for the unknown value.	$[\text{NaCl}] = \frac{0.0024\text{ g} \times 1{,}000}{\frac{125\text{ mL}}{1{,}000}}$ = 19.2 mg L^{-1} (1 mark)
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	19 ppm (1 mark)

Your turn

A 2,500 mL flask contains a solution made with 0.050 g of potassium permanganate ($KMnO_4$). **Calculate** its concentration in ppm. (2 marks)

Worked example 12.2H

Calculating mass from concentration in ppm

Worked example 12.2H: Watch a video that shows how to solve this problem.

Calculate the mass of potassium hydroxide (KOH) present in 250 L of a 3.26 ppm solution. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. In this question, we need to determine the mass. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$c = \frac{m}{V}$ $c = 3.26$ ppm, $m = ?$, $V = 250$ L
Step 3: Substitute the known values into the formula and solve for the unknown value. Recall that ppm is equivalent to mg L^{-1}.	3.26 ppm = 3.26 mg L^{-1} 3.26 mg $L^{-1} = \frac{m}{250\text{ L}}$ (1 mark) $m = 815$ mg
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	0.815 g or 815 mg (1 mark)

Your turn

Calculate the mass of sodium sulfate (Na_2SO_4) is required to make a 10 ppm solution using a 25 mL flask. (2 marks)

How do you convert between units of concentration?

When converting between units of concentration, consider the numerator and denominator of the equations. All of the concentration equations discussed so far express volume in litres (L). Therefore, when converting between units, volume can be disregarded.

Figure 4 shows a shortcut to convert between units. To calculate mass from a mole amount, multiply the mole amount by molar mass. This is demonstrated through the mole equation. Therefore, to convert from mol L^{-1} to g L^{-1}, multiply the concentration in mol L^{-1} by molar mass. To convert g L^{-1} to mol L^{-1}, divide the concentration in g L^{-1} by molar mass.

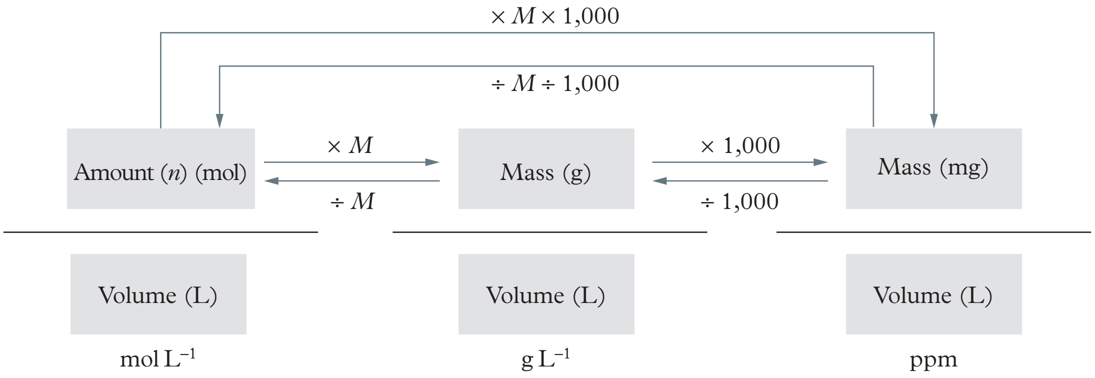


FIGURE 4 Conversion tool for concentration units mol L^{-1}, g L^{-1} and ppm, where M is molar mass.

To convert from g L^{-1} to mg L^{-1} (ppm), multiply g L^{-1} by 1,000. To convert from mg L^{-1} (ppm) to g L^{-1}, divide by 1,000.

Sometimes, it is easier to calculate a concentration in a different unit and convert between units to determine an answer. This is seen in Worked example 12.2J. Either method can be used and both result in a correct answer.

Worked example 12.2I

Converting between units of concentration

Worked example 12.2I: Watch a video that shows how to solve this problem.

Use the correct conversions to express the following values in the units indicated.

a 2.6 ppm LiOH to M (1 mark)
b 5.7 g L^{-1} to ppm (1 mark)
c 0.76 g L^{-1} of HCl to M (1 mark)
d 2.00 M HNO_3 to ppm. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the questions are asking you to do.	"Use" means to apply knowledge or rules to put theory into practice. We need to apply our understanding of concentration unit conversions. The questions are worth 1 mark each, so we must recall the theory and apply it to each quantity.
Step 2: Recall the conversion formulas and complete the calculations. See Figure 4.	**a** $\text{ppm} \xrightarrow{\div M \div 1,000} \text{M}$ **b** $\text{g L}^{-1} \xrightarrow{\times 1,000} \text{ppm}$ **c** $\text{g L}^{-1} \xrightarrow{\div M} \text{M}$ **d** $\text{M} \xrightarrow{\times M \times 1,000} \text{ppm}$
Step 3: Complete the calculations.	**a** $2.6\ \text{ppm} \div 23.95\ \text{g mol}^{-1} \div 1,000 = 1.0855 \times 10^{-4\text{M}}$ **b** $5.7\ \text{g L}^{-1} \times 1,000 = 5,700\ \text{mg L}^{-1}$ **c** $0.76\ \text{g L}^{-1} \div 36.5 = 0.02084\ \text{M}$ **d** $2.00\ \text{M} \times 63.02\ \text{g mol}^{-1} \times 1,000 = 1.26 \times 10^{5}\ \text{mg L}^{-1}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	**a** 1.1×10^{-4} M (1 mark) **b** 5.7×10^{3} ppm (1 mark) **c** 0.021 M (1 mark) **d** 1.26×10^{5} ppm (1 mark)

Your turn

Use the correct conversions to express the following values in the units indicated.

a 0.0200 M KOH to ppm (1 mark)
b 3.2 ppm of NaF to g L^{-1} (1 mark)
c 1.5 M H_2SO_4 to g L^{-1} (1 mark)
d 25 ppm Br^- to M. (1 mark)

Worked example 12.2J

Calculating molar concentration from mass and volume

Worked example 12.2J: Watch a video that shows how to solve this problem.

5.2 mg of NaCl is dissolved in water to make 3.24 L of solution. **Calculate** its concentration in mol L^{-1}. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine molar concentration. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formulas and gather any data required.	$c = \frac{m}{V}$ $c = ?, m = 5.2\,\text{mg}, V = 3.24\,\text{L}$
Step 3: Substitute the known values into the formula and solve for the number of moles. In this case, the values are provided in mg and L, so concentration is mg L^{-1} or ppm.	$[\text{NaCl}] = \frac{5.2\,\text{mg}}{3.24\,\text{L}}$ (1 mark) $= 1.6049\,\text{mg L}^{-1}$
Step 3: Convert to the required units (mol L^{-1})·	$\text{ppm} \xrightarrow{\div M \div 1{,}000} \text{mol L}^{-1}$ $[\text{NaCl}] = 1.6049\,\text{ppm} \div 58.44\,\text{g mol}^{-1} \div 1{,}000$ $= 2.7463 \times 10^{-5}\,\text{M}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	$2.7 \times 10^{-5}\,\text{M}$ (1 mark)

Your turn

5.25 mg of $K_2Cr_2O_7$ is dissolved in water to make 150 mL of solution. **Calculate** its concentration in mol L^{-1}. (2 marks)

Study tip

You can also use the method outlined in Worked example 12.2B and arrive at the same correct answer. Marks cannot be deducted as long as you show your working to demonstrate how you calculated your answer.

What is dilution?

Often, solutions have too great a concentration and must be diluted before a chemist can analyse them. To perform a **dilution**, a small volume is transferred to a new flask and water or another solvent is added (Table 2).

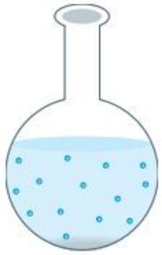
In Table 2, the original solution contains 100 molecules in 250 mL of water. From this solution, a 37.5 mL sample, containing 15 molecules, is measured, then poured into a new flask, where it still contains 15 molecules in 37.5 mL. The sample is then made up to 250 mL with water to dilute it. The new solution now contains 15 molecules in 250 mL of solution and is therefore more dilute.

Assuming that the number of particles is proportional to the number of moles, several statements can be made:

- The concentration of the original solution is the same as the concentration of the sample taken from it.
- The sample and the new diluted solution contain the same number of moles; they both contain 15 molecules.

dilution
a method for lowering the concentration of a solution

TABLE 2 Changes in concentration and number of particles during the process of dilution

	Original solution	Sample transferred to new flask	New sample made up to 250 mL with water
			
Number of particles	100	15	15
Volume (mL)	250	37.5	250

Recall that to calculate the number of moles from a solution, the following formula is used:

$$n = c \times V$$

As the sample and the diluted solution have the same number of moles, it can be deduced that the concentration multiplied by the volume of the sample will result in the same value as the concentration multiplied by the volume of the dilute solution. Therefore, the dilution equation can be used:

$$c_1 V_1 = c_2 V_2$$

where c_1 and V_1 are the concentration and volume of the sample, and c_2 and V_2 are the concentration and volume of the diluted solution.

Worked example 12.2K

Calculating the concentration of a diluted solution

Worked example 12.2K: Watch a video that shows how to solve this problem.

A 5 mL sample of a 50 mL solution of 0.2 M HCl is diluted to 250 mL. **Calculate** the concentration of the new solution. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the diluted concentration. The question is worth 2 marks, so we must gather the correct information, correctly apply the formula and complete the calculation.
Step 2: Select the correct formula and gather any data required.	$c_1 V_1 = c_2 V_2$ $c_1 = 0.2\,\text{M}$, $V_1 = 5\,\text{mL}$ $c_2 = ?$, $V_2 = 250\,\text{mL}$
Step 3: Substitute the known values into the formula and solve for the unknown value. Although it is not necessary here, it is good practice to convert volumes to litres.	$0.2\,\text{M} \times \frac{5\,\text{mL}}{1{,}000} = c_2 \times \frac{250\,\text{mL}}{1{,}000}$ $c_2 = 0.004\,\text{M}$ (1 mark)
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	0.004 M (1 mark)

Your turn

A 20 mL sample is taken from a 1.00 L solution of 1.75 M NaOH. The sample is diluted by adding water to make a final volume of 200 mL. **Calculate** the concentration of the new solution. (2 marks)

Study tip

You should only ever use $c_1 V_1 = c_2 V_2$ for dilutions. Sometimes, students doing stoichiometry questions involving solutions (which you will encounter later) tend to use this rule instead of using the mole ratio, as a shortcut... but this is not a valid procedure!

Challenge

Lead paint

Lead paint has been banned since 1978 because lead is a toxic substance. However, some older homes still contain lead-based paint on walls or woodwork. A chemist performed an analysis to determine whether the concentration of lead in a sample of wall paint is safe. The recommended maximum concentration of Pb^{2+} ions is 6.52×10^{-4} M.

A 20 mL sample of paint was reacted with dilute nitric acid to convert the insoluble $PbCO_3$ present in the paint into soluble $Pb(NO_3)_2$ and this was then diluted with water to a total volume of 500 mL. The diluted lead(II) nitrate solution was then reacted with potassium iodide to form a precipitate of lead(II) iodide. The mass of the precipitate was 0.56 g.

1 **Calculate** the concentration of lead(II) ions in the 500 mL of dilute solution. (6 marks)

2 **Determine** whether the lead(II) in the paint is at a safe concentration. (3 marks)

Real-world chemistry

Atmospheric chemistry

Earth's atmosphere is a complex, interacting system. An understanding of the concentrations of the several gases that make up the atmosphere is essential to understanding phenomena such as climate change and pollution. Sensitive instruments are used to measure the concentration of gas in air in parts per million (ppm) and parts per billion (ppb). Data has been recorded from before 1970 to 2016. A summary of this data from the United States EPA (Environment Protection Authority) website states that, in the atmosphere, the global average concentration of:

- carbon dioxide (CO_2) has increased from 280 to 399.5 ppm
- methane (CH_4) has increased from 722 to 1,834 ppb
- ozone (O_3) has increased from 237 to 337 ppb
- nitrous oxide (N_2O) has increased from 270 to 328 ppb.

Concentrations of greenhouse gases have increased significantly and are now higher than ever recorded. For example, the concentration of CO_2 is approximately 100 ppm higher than at any time in the last 800,000 years.

Water vapour (H_2O) makes up 95% of all greenhouse gases. It has a concentration of approximately 4% in the atmosphere. As temperatures increase due to human activity, the amount of water vapour in the atmosphere also increases. This amplifies the warming effect of CO_2, with recent studies suggesting that it as much as doubles the effects of CO_2. Essentially, the presence of extra water vapour in the atmosphere makes temperature changes caused by CO_2 more substantial.

Apply your understanding

1 **Calculate** the increase in the concentration of carbon dioxide in mol L^{-1}. (2 marks)

2 **Use** the data to **justify** which of the four gases has increased the least. (4 marks)

3 **Propose** a reason why the gas identified in question 2 has increased in concentration to a lesser extent than the others. (1 mark)

Check your learning 12.2

Check your learning 12.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** why the concentration units g L^{-1} and ppm might be used instead of mol L^{-1}. (1 mark)

2 **Convert** between the following units.

- a 0.056 M of $K_2Cr_2O_7$ to ppm (1 mark)
- b 25 ppm KOH to g L^{-1} (1 mark)
- c 82 g L^{-1} of H_3PO_4 to M. (1 mark)

3 **Calculate** the concentration of the following solutions in mol L$^{-1.}$

- a 0.30 mol of NaCl in 50.0 mL of solution (2 marks)
- b Iodide ions when 0.9 g of KI is dissolved in 20 mL solution (4 marks)
- c 4.93 × 10^{19} I_2 molecules in 125 mL of solution. (4 marks)

4 **Calculate** the concentration of the following solutions in ppm.

- a 0.0032 mol of HNO_3 in 345 mL of solution (4 marks)
- b 0.093 g of NaOH in 100 mL of solution (2 marks)
- c 4.32 × 10^{15} molecules of H_2SO_4 in 150 mL solution. (6 marks)

Knowledge utilisation

5 A chemist wishes to grow a copper sulfate crystal. Use your knowledge of saturation to research and **develop** a method for this. **Justify** your steps based on your knowledge of saturation. (4 marks)

6 A student analyses a solution and generates the calibration curve of concentration shown below. The blue dots represent a set of standards that have been used to determine the concentration of the sample (red dot).

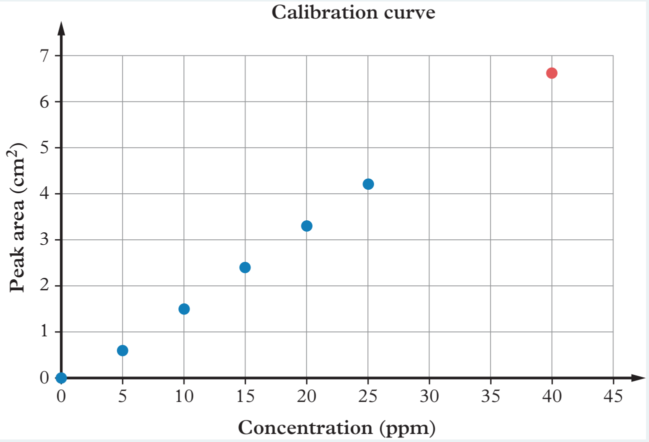


FIGURE 5 Absorbance calibration curve

- a The graph is measuring 40 ppm for the sample. **Critique** the reliability of the data. **Discuss** two things could be done to make the data more valid. (4 marks)
- b **Demonstrate,** showing calculations, how to dilute the sample so that it is within the range of the standards. (3 marks)

7 A chemist has accidentally spilt 10 g of sodium acetate (CH_3COONa) into water and now has 125 mL of solution. **Propose**, using **calculations,** how they could make a 10 ppm solution from this solution. (6 marks)

MODULE 12

Lesson 12.3

Review: Aqueous solutions and molarity

Summary

12.1

- Water has a bent shape, making it a polar molecule with a dipole or a dipole moment. This allows it to form hydrogen bonds with other water molecules.
- Water's ability to form strong hydrogen bonds with other water molecules gives rise to its unique properties.

12.2

- Concentration is a measure of the number of particles in a specific volume of solution.
- Concentration can be measured in many different units depending on the nature of the solution. Molar concentration (molarity) can be calculated if the amount (in moles) and volume of the substance are known.
- Dilution involves the addition of solvent to reduce the concentration of a solution.

Key formulas

Molar concentration	$c = \frac{n}{V}$
Concentration	$c = \frac{m}{V}$
Dilution	$c_1 V_1 = c_2 V_2$
Density	$\rho = \frac{m}{V}$

Review questions 12.3A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 The substance in which a solute dissolves to form a solution, is called a

A solvent.
B saturate.
C solute.
D substance.

2 A solution in which no more solute can dissolve is referred to as

A a solution.
B supersaturated.
C saturated.
D unsaturated.

3 Which of the following is not a unit of concentration?

A M
B mol L^{-1}
C ppm
D kJ L^{-1}

4 What is the increase in volume of solution when 40 mL of 2 M HCl is diluted to make a 0.4 M solution?

A 200 mL
B 2 L
C 160 mL
D 240 mL

5 0.056 g of LiOH is dissolved to make 9.6 L of solution. Which answer is not a correct concentration of this solution?

A 2.4×10^{-4} M

B 5.8×10^{-3} g L^{-1}

C 5.8 ppm

D 5.8×10^{-3} M

6 What mathematical operation must you apply to convert a value from mL to L?

A × 1,000

B ÷ 1,000

C × 100

D ÷ 100

7 A solution that forms crystals is referred to as

A supersaturated.

B saturated.

C unsaturated.

D desaturated.

8 Which is the correct formula for the dilution equation?

A $cV_1 = V_2$

B $c_1V_2 = c_2V_1$

C $c_1c_2 = V_1V_2$

D $c_1V_1 = c_2V_2$

9 The type of bonding between atoms in a water molecule is

A hydrogen bonding.

B ionic bonding.

C non-polar covalent bonding.

D polar covalent bonding.

10 Ice is less dense than liquid water because as water molecules reach low temperatures and freeze,

A the molecules form a rigid open lattice structure held together by hydrogen bonding.

B the molecules begin to repel each other and move further apart.

C the molecules have less heat energy and cannot move closer together.

D the dipole on the molecules is stronger at lower temperatures.

11 Water has a high melting and boiling point for a molecule with a relatively low molar mass because

A the dispersion forces between water molecules require a large amount of energy to overcome.

B the strong covalent bonds within each water molecule require a large amount of energy to break.

C the hydrogen bonding within each molecule requires a large amount of energy to break.

D the hydrogen bonding between water molecules requires a large amount of energy to overcome.

12 100 mL of a 2 M solution contains

A 2 mol of solute in 100 mL of water.

B 2 mol of solute in 100 mL of solution.

C 0.2 mol of solute in 100 mL of water.

D 0.2 mol of solute in 100 mL of solution.

Review questions 12.3B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

13 **Explain** why density is often referred to as a unit of concentration. (2 marks)

14 A stream of water running from a tap curves towards a charged rod placed near it. **Explain**, using the properties of water, why this phenomenon occurs. (3 marks)

15 Water is referred to as a “universal solvent”. **Explain** why. (2 marks)

16 **Calculate** the volume of a solution required to make a 0.2.0 M KNO_3 solution from 0.5 g of KNO_3. (4 marks)

17 **Calculate** the molarity of a solution made from 0.2 g of KOH in 45 mL of solution. (4 marks)

18 **Calculate** the mass of NaCl that must be added to water to generate 450 mL of a 0.50 M solution. (4 marks)

19 **Calculate** the molar concentration of nitrate ions in a 100.0 mL solution containing 2.00 g of lead(II) nitrate. (4 marks)

20 **Calculate** the concentration of a 50.0 mL solution prepared from 4.22×10^{21} units of $CaCl_2$ in water. Express your answer in g L^{-1}. (6 marks)

21 **Calculate** the concentration of dichromate ions in a solution containing 2.6 mg $K_2Cr_2O_7$ in 105 mL of solution. Express your answer in g L^{-1}. (6 marks)

22 **Calculate** the mass of potassium permanganate, in grams, in a 246 mL solution of 0.600 $g\,L^{-1}$ $KMnO_4$. (2 marks)

23 A 0.045 g sample of $FeCl_3$ is used to make a 2.70 L solution. **Calculate** the concentration in ppm. (3 marks)

24 **Calculate** the mass of NaOH required to make 80.0 mL of a 2.7 ppm aqueous solution. (2 marks)

25 **Calculate** the volume of the solution if 0.01 g of $NaNO_3$ is dissolved in water to make a 5.0 ppm solution. (2 marks)

26 **Explain** what a dilution is. (1 mark)

27 Dilution can be used to prepare solutions from more concentrated chemical stocks.

- a **Calculate** the volume, in millilitres, of a 2.2 M $CaCO_3$ solution needed to prepare 120 mL of a 0.25 M solution. (2 marks)
- b 50.0 mL of a $BaCl_2$ solution is diluted to 425.0 mL and has a new concentration of 2.0 M. **Calculate** its original concentration. (2 marks)
- c 25 mL of a 1.0 M solution is used to make a diluted 0.25 M solution. **Calculate** the volume of the diluted solution. (2 marks)

Analytical processes

28 Aqueous solutions have slightly higher boiling points than pure water. The increase in the boiling point is proportional to the number of particles of solute. Three 1.0 M solutions were made and their boiling points were measured. The solutions were sucrose (sugar), sodium chloride and copper(II) nitrate. **Compare** the solutions in terms of the number of solute particles in each solution to:

- a **determine** why they will not have identical boiling points. (3 marks)
- b **determine** which of the three solutions would have the highest boiling point. (1 mark)

29 **Use** examples to **discriminate** between the terms “solute”, “solvent” and “solution”. (3 marks)

Compare the structures of water and trichloroethene to **explain** why water is unable to dissolve grease but trichloroethene is a suitable solvent. (3 marks)

30 100 g of water is heated to 60°C in order to dissolve 106 g of potassium nitrate. When the solution is cooled to 25°C, crystals form at the bottom of the beaker. After filtering, the crystals are allowed to dry and are weighed. The mass is 66 g. **Deduce** the solubility of potassium nitrate, in grams per 100 g of water at 25°C. **Clarify** the assumptions that have been made in order to deduce the solubility. (3 marks)

Knowledge utilisation

31 A scientist performs an experiment and determines that a saturated sodium chloride solution had a boiling point of 108.7°C, whereas water has a boiling point of 100°C. **Evaluate** the scientist’s findings and **explain** why the salt water has a higher boiling point. (3 marks)

32 **Predict** and **justify** whether pure water or salt water has a higher surface tension. (3 marks)

Data drill

Surface tension

An experiment was conducted to determine the trends in surface tension of water (H_2O), methanol (CH_3OH) and ethanol (CH_3CH_2OH) at various temperatures (Figure 1).

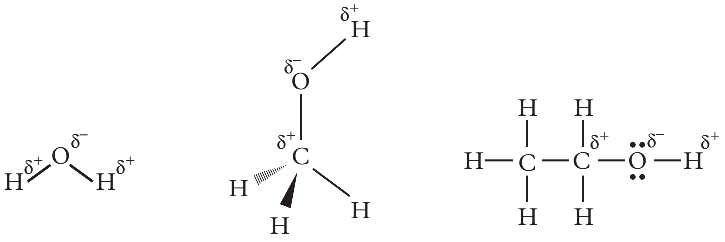


FIGURE 1 The structural formulas of (A) water, (B) methanol and (C) ethanol

Four different temperatures were tested for each substance. The results are shown in Table 1.

TABLE 1 Results from surface tension experiment

Substance	Temperature (K)	Surface tension ($mN\,m^{-1}$)
Water	274.00	75.9857
Water	282.20	74.6379
Water	290.40	73.2832
Water	306.80	70.5455
Methanol	274.90	24.0569
Methanol	281.56	23.4820
Methanol	294.88	22.3546
Methanol	301.54	21.7984
Ethanol	273.24	23.8716
Ethanol	286.56	22.7496
Ethanol	293.22	22.2023
Ethanol	306.54	21.1228

Apply understanding

1 **Identify** the independent and dependent variables. (1 mark)

2 **Identify** the lowest temperatures tested for each substance. **Calculate** the average minimum temperature and percentage uncertainty. Repeat this for the maximum temperatures. (4 marks)

Analyse data

3 **Organise** the data from Table 1 into a graph. (4 marks)

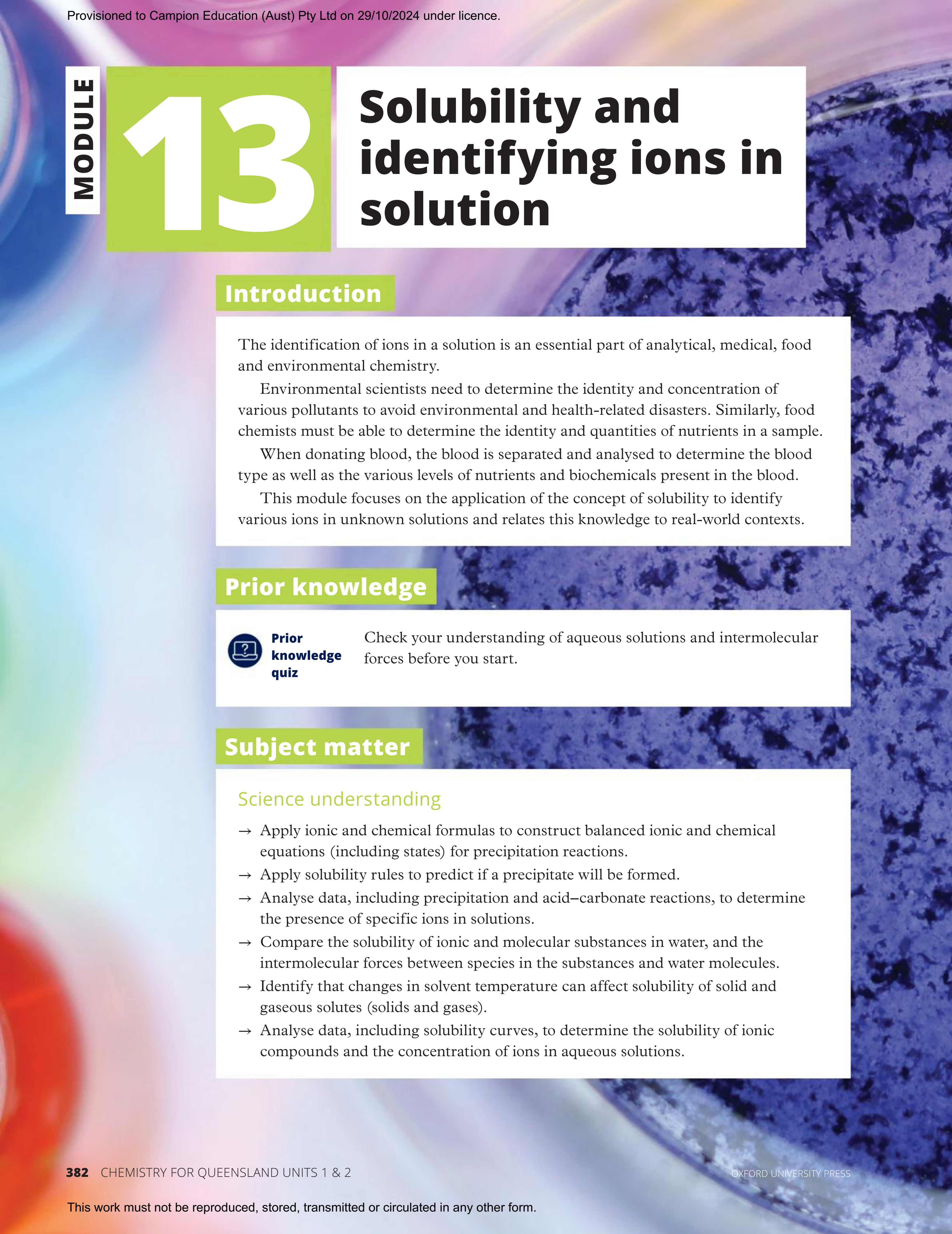
Evaluate evidence

4 **Draw a conclusion** from the data. (2 marks)

5 **Extrapolate** the data to **predict** the surface tension of water at 100°C. (1 mark)

Module 12 checklist: Aqueous solutions and molarity

Quizlet: Revise key terms online to test your understanding



MODULE 13

Solubility and identifying ions in solution

Introduction

The identification of ions in a solution is an essential part of analytical, medical, food and environmental chemistry.

Environmental scientists need to determine the identity and concentration of various pollutants to avoid environmental and health-related disasters. Similarly, food chemists must be able to determine the identity and quantities of nutrients in a sample.

When donating blood, the blood is separated and analysed to determine the blood type as well as the various levels of nutrients and biochemicals present in the blood.

This module focuses on the application of the concept of solubility to identify various ions in unknown solutions and relates this knowledge to real-world contexts.

Prior knowledge

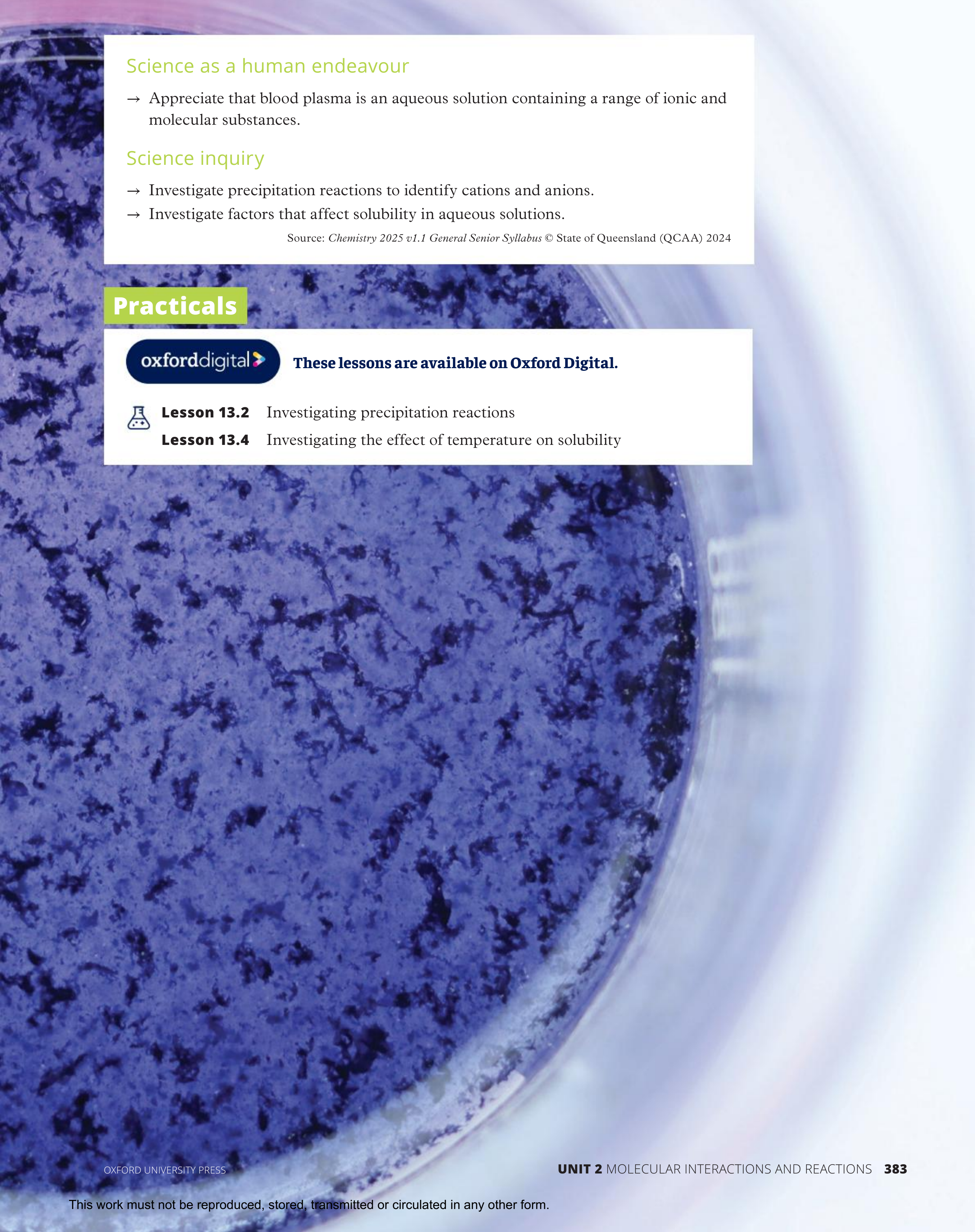
Prior knowledge quiz

Check your understanding of aqueous solutions and intermolecular forces before you start.

Subject matter

Science understanding

→ Apply ionic and chemical formulas to construct balanced ionic and chemical equations (including states) for precipitation reactions.
→ Apply solubility rules to predict if a precipitate will be formed.
→ Analyse data, including precipitation and acid–carbonate reactions, to determine the presence of specific ions in solutions.
→ Compare the solubility of ionic and molecular substances in water, and the intermolecular forces between species in the substances and water molecules.
→ Identify that changes in solvent temperature can affect solubility of solid and gaseous solutes (solids and gases).
→ Analyse data, including solubility curves, to determine the solubility of ionic compounds and the concentration of ions in aqueous solutions.



Science as a human endeavour

→ Appreciate that blood plasma is an aqueous solution containing a range of ionic and molecular substances.

Science inquiry

→ Investigate precipitation reactions to identify cations and anions.

→ Investigate factors that affect solubility in aqueous solutions.

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Practicals

These lessons are available on Oxford Digital.

Lesson 13.2 Investigating precipitation reactions

Lesson 13.4 Investigating the effect of temperature on solubility

Lesson 13.1

Identifying ions in solution

Key ideas

- → A precipitate is an insoluble solid that forms from reactions between some ionic compounds.
- → Precipitation reactions can be represented using balanced full and ionic chemical equations with state symbols (s), (l), (g) and (aq) included.
- → A solubility table can be used to predict and identify whether a precipitate will form from chemical reactions in aqueous solutions.

Learning intentions and success criteria

What is a precipitation reaction?

precipitation reaction
a reaction between two soluble ionic substances that forms a solid product

precipitate
the solid product formed as a result of a precipitation reaction

In a **precipitation reaction**, a **precipitate** (solid) is formed when two aqueous ionic solutions are mixed together. Because each ionic substance consists of cations and ions, the cation and anion swap partners during a reaction to become different substances.

In the reaction between lead(II) nitrate and potassium iodide solutions, lead(II) and potassium ions are cations. Cations are always written first in the name of an ionic substance. Nitrate and iodide are the anions, and they are written after the cations in each compound. When these two substances react, they swap partners to form solid lead(II) iodide precipitate and aqueous potassium nitrate, according to the equation:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow Pbl_2(s) + 2KNO_3(aq)$$

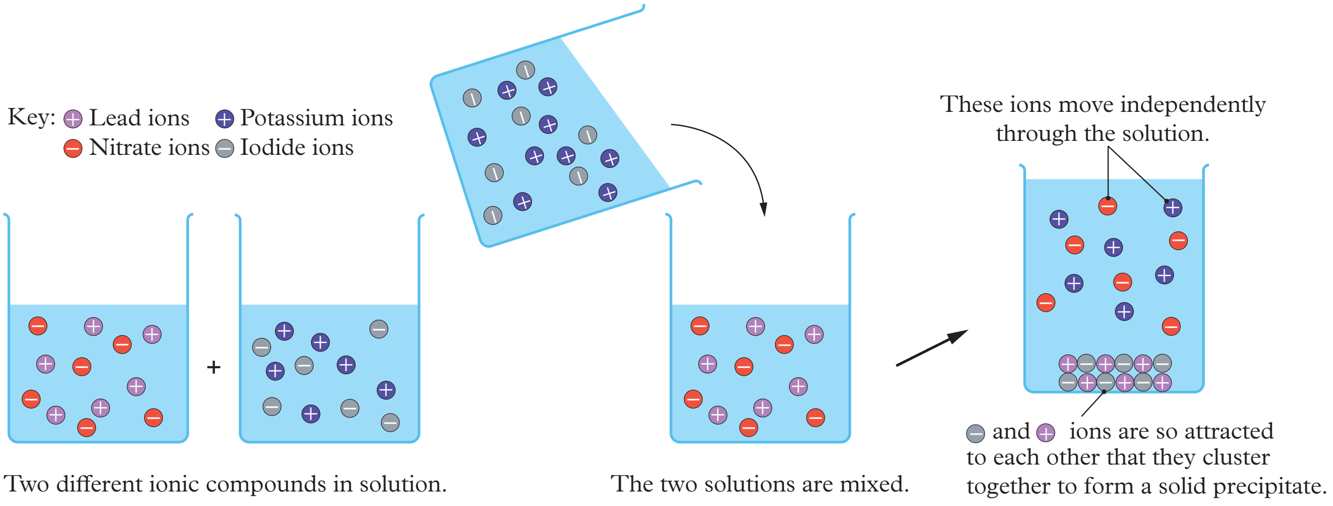


FIGURE 1 In a precipitation reaction, ions in aqueous solution combine to form a solid called a precipitate.

Study tip
Review the swap-and-drop method from Module 3 to write ionic formulas. This same method is used to write the formulas for the products formed from a reaction between two ionic substances.

However, many mixtures of ionic reactants do not result in the formation of a precipitate. To determine whether a reaction occurs to produce a precipitate, you can use a set of solubility rules.

What are the solubility rules?

The solubility rules provide a guideline for the behaviour of ionic substances at 25°C and 1 bar. These rules are provided in the QCAA Formula and data book and in Table 1. For substances indicated as soluble in water, more than 10 g will dissolve in 1 L of water. For substances considered insoluble, less than 1 g will dissolve in 1 L of water.

TABLE 1 Solubility rules of salts in water from the QCAA Formula and data book

Group	Soluble	Exceptions
Nitrates (NO_3^-)	all	none
Group 1 (Li^+, Na^+ and K^+)	all	none
Ammonium (NH_4^+)	all	none
Ethanoates (CH_3COO^-)	all	none
Chlorides (Cl^-)	all	Ag^+ and Hg^{2+} (p)
Bromides (Br^-)	all	Ag^+, Pb^{2+} (p) and Hg^{2+} (p)
Iodides (I^-)	all	Ag^+, Pb^{2+}, Hg^{2+} and Cu^{2+}
Sulfates (SO_4^{2-})	all	Ag^{+*}, Pb^{2+}, Hg^{2+}, Ba^{2+} and Ca^{2+} (p)
Group	**Insoluble**	**Exceptions**
Silver (Ag^+)	all	NO_3^- CH_3COO^- and SO_4^{2-} (p)
Lead(II) (Pb^{2+})	all	NO_3^-, CH_3COO^-, Cl^- and Br^- (p)
Mercury (Hg^{2+})	all	NO_3^- and CH_3COO^-
Hydroxides (OH^-)	all	NH_4^+, Ba^{2+}, Ca^{2+} (p) and Group 1 metals
Oxides (O^{2-})	all	Ba^{2+}, Ca^{2+} (p) and Group 1 metals (no data for NH_4^+)
Carbonates (CO_3^{2-})	all	NH_4^+ and Group 1 metals (no data for Cu^{2+}, Al^{3+} and Fe^{3+})
Phosphates (PO_4^{3-})	all	NH_4^+, Mg^{2+} (p), Mn^{2+} (p), Group 1 metals (no data for Li^+)

(p) partially soluble in water (between 1 and 10 g will dissolve in 1 L of water)

Consider the following reaction occurring in water:

$$Pb(NO_3)_2 + 2NaOH \rightarrow Pb(OH)_2 + 2NaNO_3$$

To determine the solubility of the substances in the equation, the identity of the anions and cations are used with the solubility rules.

- Lead(II) nitrate: The anion is nitrate (NO_3^-). Table 1 indicates that all nitrates are soluble. A soluble substance is one that dissolves in water. Therefore, the **state** of $Pb(NO3)_2$ is aqueous or (aq).
- Sodium hydroxide: The anion is hydroxide (OH^-). Table 1 indicates that most hydroxides are only slightly soluble, except NaOH, KOH and $Ca(OH)_2$. Also, it indicates that most salts of Na^+ are soluble. Therefore, NaOH is soluble and is assigned the aqueous (aq) state.
- Lead(II) hydroxide: The anion is hydroxide (OH^-). Table 1 indicates that most hydroxides are only slightly soluble, except NaOH, KOH and $Ca(OH)_2$. Therefore, $Pb(OH)_2$ is **insoluble**. An insoluble substance forms a solid precipitate and is assigned the solid (s) state.
- Sodium nitrate: The anion is nitrate (NO_3^-). Table 1 indicates that all nitrates are soluble. Therefore, $NaNO_3$ is soluble and is assigned the (aq) state.

The chemical equation can now be rewritten to include the states:

$$Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$$

If this reaction does not result in the formation of a solid, then it is not a precipitation reaction.

Study tip

Think about t-shirt sizes, **S**mall–**M**edium–**L**arge, to remember the exceptions to the halide salts (Cl^-, Br^- and I^-). They are all soluble except for **S**ilver, **M**ercury and **L**ead.

state
the state of matter assigned to each chemical reactant or product – solid (s), liquid (l), gas (g) or aqueous (aq)

insoluble
not able to dissolve in a solvent

Slightly soluble substances

partially soluble only partially dissolves in a solvent, forming a solution of low concentration

In Table 1, you will notice that some ionic compounds are classified as **partially soluble** (indicated with a (p)). This means that some of the chemical dissolves in the solvent, but most remains as a solid. Specifically, only between 1 and 10 grams of the solute will dissolve in 1 L of water. The ions are so strongly attracted to one another (i.e. attractions between the anions and cations are very large) that they are less likely to dissociate or dissolve.

Worked example 13.1A

Determining balanced equations for precipitation reactions

Worked example 13.1A: Watch a video that shows how to solve this problem.

Determine the products of a reaction between silver nitrate and potassium chloride occurring in water and **determine** a balanced chemical equation (including states) for the reaction. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to use the solubility rules to determine the states of the reactants and products, and write them as a balanced equation. The question is worth 1 mark, so we must recall the theory and apply it to the reaction.
Step 2: Identify the ions involved in the reaction.	Silver nitrate: Ag^+, NO_3^- Potassium chloride: K^+, Cl^-
Step 3: Swap the cations and anion pairs, then apply the swap-and-drop method to find the ionic formulas of the products.	Silver chloride: $AgCl$ Potassium nitrate: KNO_3
Step 4: Write the full chemical equation. Add coefficients as required to balance the equation.	$AgNO_3 + KCl \rightarrow AgCl + KNO_3$ This equation is already balanced, so no coefficients need to be added.
Step 5: Determine the states of each species using Table 1.	$AgNO_3$ and KNO_3 are soluble since all nitrates are soluble. KCl is also soluble since all group 1 salts are soluble. These three species have the state symbol (aq). AgCl is insoluble since it is an exception to the rule that all chlorides are soluble, so it has the state symbol (s).
Step 6: Finalise your answer.	$AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$ (1 mark)

Your turn

Determine the products of a reaction between manganese(II) phosphate and barium ethanoate occurring in water and **determine** a balanced chemical equation (including states) for the reaction. (1 mark)

How do you write ionic equations?

spectator ion an ion that does not participate in a chemical reaction

In addition to full chemical equations, there is another way in which precipitation reactions are represented. In this format, the **spectator ions** are often excluded. A spectator ion does not participate in a chemical reaction and its state does not change throughout the reaction.

In the precipitation reaction between sodium chloride and silver nitrate, the following balanced chemical equation is generated:

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

In this reaction, sodium (Na^+) and nitrate (NO_3^-) ions do not participate in the reaction and do not change state – they are spectator ions. The chemical equation can be rewritten to demonstrate the ions in solution before and after reacting:

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + AgCl(s)$$

Because $Na^+(aq)$ and nitrate $NO_3^-(aq)$ are unchanged as reactants and products, they are cancelled out of the equation:

$$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)} + AgCl(s)$$

The equation can be rewritten as an **ionic equation**, excluding spectator ions:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Study tip

Reactants and products can only be cancelled out of an equation if they have an equal number on each side.

ionic equation
a chemical equation that shows only the ions that react during a chemical reaction, along with any neutral chemical species

Worked example 13.1B

Predicting products of a reaction in aqueous solution

Worked example 13.1B: Watch a video that shows how to solve this problem.

Consider the reaction between aqueous solutions of barium chloride and sodium sulfate.

a **Determine** a balanced precipitation reaction equation. (1 mark)

b **Determine** an ionic equation for the reaction. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. In this question, we need to write balanced full and ionic equations. The questions are worth 1 mark each, so we must recall the theory and apply it to the reaction.
Step 2: For part **a**, determine the products and their solubilities, then write a full balanced equation for the reaction, including states. For a refresher, go to Worked example 13.1A.	All compounds are soluble apart from $BaSO_4$, as the barium salt is an exception to the rule that all sulfates are soluble. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$ (1 mark)
Step 3: For part **b**, identify the spectator ions. These do not change state in the reaction.	Na^+ and Cl^- are the spectator ions. They are not included in the final ionic equation.
Step 4: Write a full ionic equation for the reaction after removing the spectator ions.	$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ (1 mark)

Your turn

Consider the reaction between aqueous solutions of copper(II) nitrate and lithium hydroxide.

a **Determine** a balanced precipitation reaction equation. (1 mark)

b **Determine** an ionic equation for the reaction. (1 mark)

Real-world chemistry

Removing mercury pollutants from the environment

Pollutants in the atmosphere have caused problems such as acid rain, respiratory issues, poisoning of waterways and the bioaccumulation of toxic chemicals in marine life. One example is mercury(II) ($Hg^{2+}(aq)$), which (even at low concentrations) has many negative health effects.

When present as the salt of any soluble anion (anion–), $Hg^{2+}(aq)$ can be precipitated out of a water sample by adding sodium iodide. The reaction forms mercury(II) iodide ($HgI_2(s)$), a red–orange precipitate:

$$Hg^{2+}(aq) + anion^{-}(aq) + Na^{+}(aq) + 2I^{-}(aq) \rightarrow HgI_2(s) + Na^{+}(aq) + anion^{-}(aq)$$

This reaction was important for identifying dissolved mercury and removing it from the water, particularly during the mercury pollution of Minamata Bay, Japan. In 1908, a factory was built close to the bay to produce chemicals required for fertilisers. Around 1932, the factory expanded to synthesise other, more dangerous chemicals, such as ethanal (CH_3CHO). The workers did not realise that a side reaction was taking place, forming methyl mercury chloride (CH_3HgCl). This highly toxic compound was dumped into the bay between 1951 and 1968.

In 1956, many residents of Minamata Bay were taken to hospital with symptoms of convulsions and difficulty in speaking and walking. Many people living in the area were found to suffer the same symptoms. The doctors named it Minamata disease. Later, the cause was found to be heavy metal poisoning due to consumption of tuna fished from the contaminated water.

Apply your understanding

1 **Construct** an ionic equation for the formation of mercury(II) iodide. (1 mark)
2 **Explain** why sodium iodide is used to remove mercury from water. (2 marks)
3 **Identify** one other ionic compound that could be used to remove mercury from water. (1 mark)

How are carbonates identified using acids and lime water?

acid–carbonate reaction
a chemical reaction between an acid and a carbonate to form a salt, carbon dioxide gas and water

Chemists are often required to determine the identity of ions in an "unknown" solution. Precipitation and **acid–carbonate reactions** can be used for this purpose. The solubility rules can be used to identify various anions and cations in a solution. Alternatively, acid–carbonate reactions produce carbon dioxide gas, which can be identified in various ways.

Most carbonates (CO_3^{2-}) are only slightly soluble. However, when the carbonate anion is paired with a sodium (Na^+), potassium (K^+) or ammonium (NH_4^+) cation, they are soluble. Because other salts are slightly soluble, it can be difficult to determine whether a precipitate has formed.

Carbonate ions can be identified in a solution by first reacting them with an acid, in an acid–carbonate reaction, to form a salt, carbon dioxide and water. For example, sodium carbonate (Na_2CO_3) reacts with hydrochloric acid (HCl) to form sodium chloride salt (NaCl), carbon dioxide gas (CO_2) and water (H_2O) as a liquid (l):

$$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(aq) + H_2O(l)$$

The sodium and chloride ions are aqueous as both reactants and products. Therefore, the ionic equation for this process is:

$$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$

The formation of bubbles is a key indicator that a gas has been formed in a reaction. Because carbon dioxide is a clear odourless gas, it is impossible to determine the identity of the gas without further testing.

Carbon dioxide extinguishes fire. Therefore, a simple test is to place a lit match in the gas formed by the reaction and observe the behaviour of the flame. If the flame is extinguished, there may be carbon dioxide, but some other gases also extinguish flames. To confirm the presence of carbon dioxide, the gas produced in the reaction can be collected and used in a second reaction.

FIGURE 2 Bubbling carbon dioxide through lime water causes a precipitate of calcium carbonate to form, seen as formation of a cloudy suspension.

Lime water ($Ca(OH)_2(aq)$) is a useful substance for determining whether a gas is carbon dioxide. When carbon dioxide is bubbled through lime water (a clear solution), a calcium carbonate precipitate is formed (Figure 2), according to the equation:

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

Challenge

Limestone

Geologists use diluted hydrochloric acid (HCl) to test whether a rock is a limestone (calcium carbonate). If the rock bubbles when doused with HCl, it is identified as a limestone. Based on what you have learnt, **predict** the type of gas the limestone produces when dissolved with HCl. **Construct** an equation to reflect this reaction. (1 mark)

FIGURE 3 Geologist at a rock outcrop

How are halogens identified using silver?

Chloride (Cl^-), bromide (Br^-) and iodide (I^-) ions all form solid ionic substances with a silver cation (Ag^+). The solubility table verifies that chloride ions precipitate when reacted with silver ions. Although it does not mention bromide and iodide, the rule can be extended to group 7 elements, the halogens. Therefore, to determine the presence of halogen ions in solution, the solution is reacted with silver nitrate ($AgNO_3$). For example, a solution of silver nitrate reacts with sodium chloride to form a white precipitate of silver chloride:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

According to the solubility rules, the sodium cation and the nitrate anion will not precipitate and therefore remain in an aqueous state as reactants and products. Any solid formed is a result of a reaction between silver and chloride:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s) \text{ (white precipitate)}$$

Silver nitrate forms a precipitate with all halogens, which can be identified by their different colours. Silver bromide is a cream solid, and silver iodide is yellow (Figure 4).

$$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s) \text{ (cream precipitate)}$$
$$I^-(aq) + Ag^+(aq) \rightarrow AgI(s) \text{ (yellow precipitate)}$$

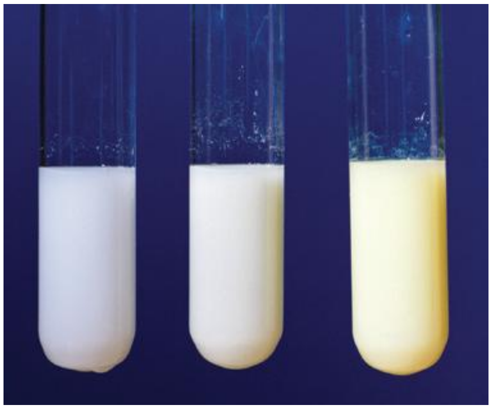


FIGURE 4 Silver chloride (AgCl) precipitate is white (left), silver bromide (AgBr) is cream (centre), and silver iodide (AgI) is yellow (right).

How are metal cations identified using sodium hydroxide and ammonia?

Sodium hydroxide is a useful ionic compound for identifying metal cations. Most sodium compounds are soluble, and most hydroxides are slightly soluble. This means that any precipitate formed must be a result of the hydroxide reacting with a different metal cation.

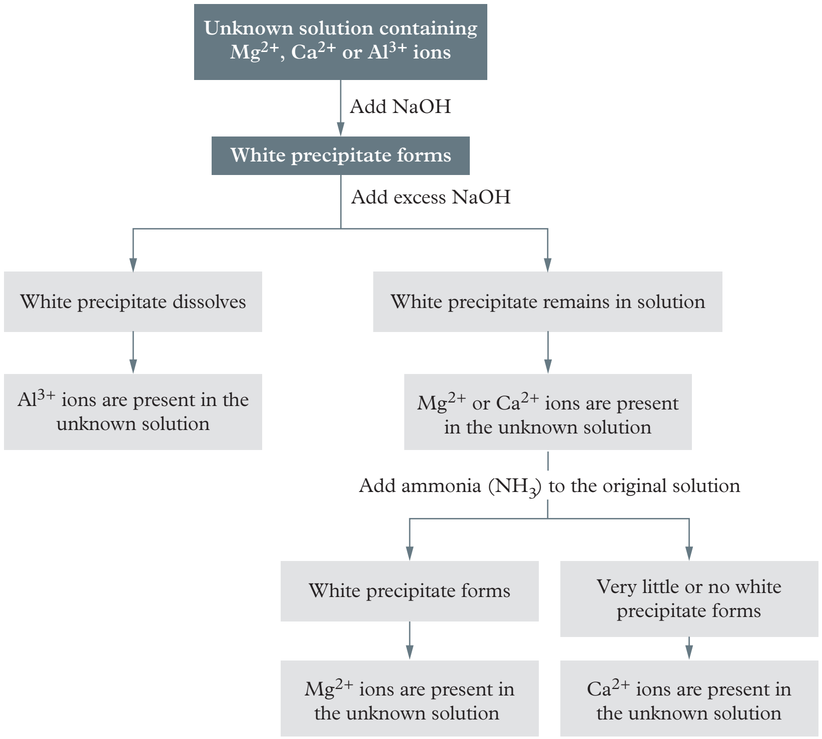


FIGURE 5 Flow chart for the identification of calcium, magnesium and aluminium ions using sodium hydroxide and ammonia

If a metal is not a transition metal, it will precipitate as a white solid. This makes it very difficult to identify the metal and further testing is required to distinguish one metal from another.

Calcium, magnesium and aluminium form white precipitates when reacted with sodium hydroxide. To identify the metal, excess NaOH is added. If the precipitate re-dissolves, then the metal is aluminium (Figure 5).

Calcium and magnesium do not re-dissolve on addition of excess NaOH. A further test must be performed to distinguish between them. Ammonia (NH_3) is added to the original solution, and if a white precipitate forms, magnesium is present. Calcium will also produce a precipitate when ammonia is added, but to a far lesser extent than magnesium.

FIGURE 6 Left to right: precipitates of iron(II), iron(III), copper(II) and nickel(II) when reacted with sodium hydroxide

Transition metals

The solubility rules do not explain the solubilities of all metal ions. The ions of transition metals form coloured solutions, and precipitate with unique colours on reaction with sodium hydroxide (Table 2).

TABLE 2 Colours of hydroxide precipitates of some common transition metals

Name	Formula	Colour
Iron(II)	Fe^{2+}	Green–brown
Iron(III)	Fe^{3+}	Brown–dark red
Copper(II)	Cu^{2+}	Blue
Nickel(II)	Ni^{2+}	Green

The hydroxide reacts with the metal ions to form an insoluble coloured precipitate (Figure 6). If sodium hydroxide is added in excess, there will be none of the original metal ions left in the solution, and it will result in a clear solution.

Study tip

In ionic substances, always remember to write the cation first and then the anion.

Challenge

Identifying ionic solutions

A chemist has four unknown solutions, labelled 1–4, and knows that they contain either iron(II), copper, aluminium or magnesium. **Develop** an overall method to determine which solution would contain each metal cation. (4 marks)

Skill drill

Designing an experiment to determine the solubility of lead compounds

Science inquiry skill: Planning investigations (Lesson 1.4)

A student wants to develop a method to experimentally determine the solubility of lead(II) ions in solutions containing ethanoate, nitrate, chloride, iodide, bromide and sulfate ions.

Practise your skills

1 **Construct** a hypothesis for the experiment. (3 marks)
2 **Construct** an aim for the experiment. (1 mark)
3 **Explain** how they could collect and record their results. Include an explanation of whether the data would be qualitative or quantitative. (3 marks)

Check your learning 13.1

Check your learning 13.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** the behaviour of substances that are slightly soluble. (2 marks)
2 **Explain** why a cation must always occur with an anion. (1 mark)

Analytical processes

3 **Determine** whether these substances are soluble in water. Give reasons to support your answers.

a $NaNO_3$ (2 marks)	**b** $Pb(OH)_2$ (2 marks)	**c** $HgCl_2$ (2 marks)
d KCl (2 marks)	**e** $BaSO_4$ (2 marks)	**f** KOH (2 marks)
g NH_4Cl (2 marks)	**h** Na_2SO_4 (2 marks)	**i** $Mg(OH)_2$ (2 marks)
j $Cu(OH)_2$ (2 marks)	**k** $(NH_4)_3PO_4$ (2 marks)	**l** KI. (2 marks)

4 **Determine** whether the following reactions are precipitation reactions. **Explain** why.
   a $Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2 + 2NaNO_3$ (2 marks)
   b $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$ (2 marks)
   c $FeSO_4 + Mg(NO_3)_2 \rightarrow Fe(NO_3)_2 + MgSO_4$ (2 marks)

5 **Determine** balanced precipitation reaction and ionic equations for the reactions that occur when the following solutions are mixed:
   a cobalt(II) nitrate and ammonium carbonate (2 marks)
   b barium oxide and aluminium nitrate (2 marks)
   c iron(II) nitrate and sodium phosphate (2 marks)
   d lead(II) nitrate and aluminium sulfate (2 marks)
   e silver nitrate and magnesium iodide. (2 marks)

Knowledge utilisation

6 A student wants to perform qualitative and quantitative tests on a precipitation reaction.
   a **Discuss** why a precipitation reaction is considered qualitative. (2 marks)
   b **Devise** a way to make a precipitation reaction quantitative. (1 mark)
   c **Propose** a method to determine the amount of chloride ions in a sample of seawater (containing sodium chloride), using a precipitation reaction as the first step. (4 marks)

7 A student in your class claims that all ions participate in a chemical reaction. **Evaluate** this statement and communicate a response, either affirming or negating their statement. (2 marks)

Practical

Lesson 13.2

Investigating precipitation reactions

Learning intentions and success criteria

Video demonstration

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 13.3

Factors affecting solubility

Learning intentions and success criteria

Key ideas

→ Solutes dissolve in solvents that have similar properties.

→ In general, an increase in temperature (an increase in kinetic energy) results in an increase in solubility for ionic solids.

→ In general, gases become less soluble with increasing temperature.

What factors affect the solubility of substances in water?

Solubility is the ability of a substance (whether it be a solid, liquid or gas) to dissolve in a solvent. If it dissolves easily, the substance is soluble. If it does not dissolve, the substance is insoluble.

Solubility depends on the strength of the intermolecular bonding that holds the molecules or ions within the substance together, the way that the solvent molecules can interact with the solute, and the temperature of the solvent. By altering the intermolecular forces, the properties and temperature of the solvent will change the solubility of a substance.

The ability of substance to dissolve in water depends on the:

- attraction between the substance and the polar water molecules (intermolecular bonding)
- attractive forces within the substance that are holding one particle to another
- the temperature of the solvent, together with whether the solute is a gas or a solid.

For solids to dissolve in water, the hydrogen bonding between water molecules must be disrupted to allow the solute particles to move between them. This requires energy. To compensate, there needs to be sufficient attractive forces between the solute particles and the water molecules.

FIGURE 1 Potassium permanganate ($KMnO_4$) dissolves in water to form a pink solution.

Are ionic substances soluble?

Salts are ionic substances that contain a positive cation (metal) and a negative anion (non-metal). Because salts consist of ions, they contain whole positive and negative charges, which are electrostatically attracted to one another. This is a strong type of bonding (stronger than dipole–dipole attractions) because it involves whole charges (ionic substances), as opposed to partial charges (polar covalent molecules). When a salt (the solute) is dissolved in water (the solvent) to form a salt solution, the ionic substance with whole charges (e.g. NaCl) will interact with a polar substance (e.g. H_2O), which has partial positive and negative charges.

It is a misconception that because ionic substances have whole positive and negative charges, which are stronger than partial charges, the salt will not dissociate. This is true if there is only one water molecule interacting with one salt formula unit, but in reality, this is not the case. Many water molecules will interact with each of the cations and anions in the salt. The sum of all of the partial charges attached to many water molecules provides them with a very strong attractive force to pull the ionic substance apart (Figure 2). This process is called dissolving.

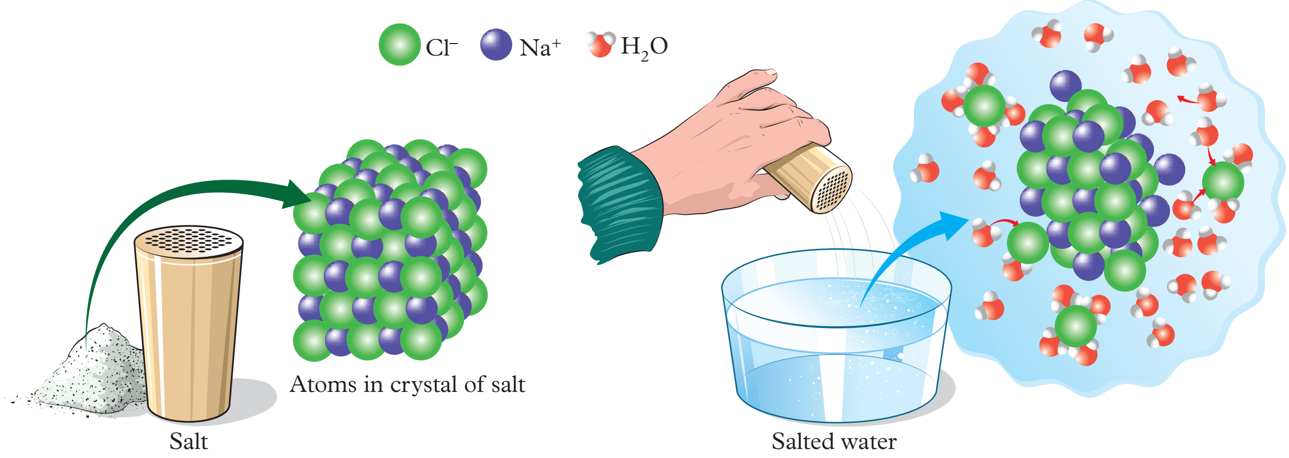
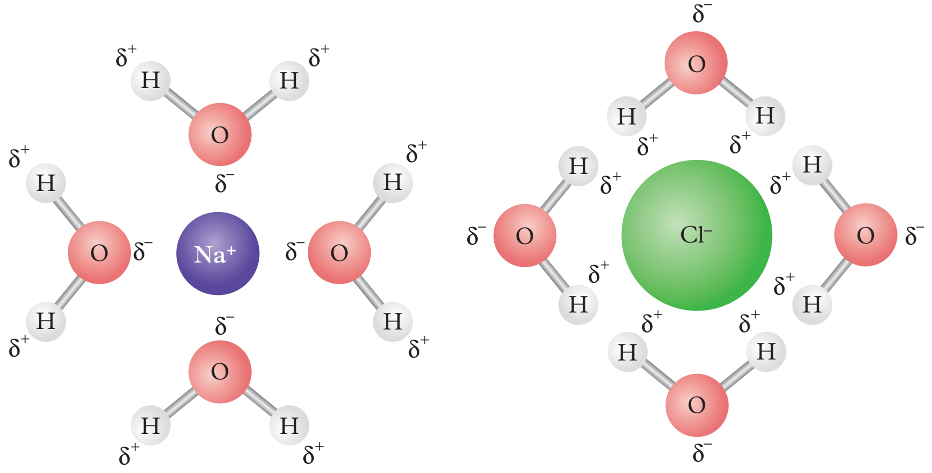


FIGURE 2 Water molecules dissociate anions from cations when they interact.

In water, the salt dissociates (breaks apart) as the sodium cation (Na^+) is attracted to the partially negative part of water (the oxygen atom), and the chloride anion (Cl^-) is attracted to the partially positive part of water (the hydrogen atoms). Once the salt has dissolved in the water, the ions are hydrated and are surrounded by water molecules.



Hydration of sodium ion

Hydration of chloride ion

FIGURE 3 A hydrated Na^+ cation and Cl^- anion and their attraction to water molecules.

For an ionic substance to dissolve (or be soluble) in water, it must form new attractive forces with the water, which replace any existing bonds between cations and anions. In sodium chloride, the existing bonds within the sodium chloride solute involve the ionic bond between groups of neighbouring Na^+ and Cl^- ions. The ionic bond is replaced with an **ion–dipole interaction**, in which the ions are attracted to the dipole of water.

ion–dipole interaction
an ion with a whole charge interacting with a partial charge caused by a polar bond or molecule

Are molecular substances soluble in water?

Molecular substances can dissolve in water if they are polar. The permanent dipoles within the molecules are able to form intermolecular forces with the solvent molecules. If the solute can form hydrogen bonds with water, this intermolecular force is even stronger than dipole–dipole attractions, and solutes in this group are usually soluble. For example, the partially positive hydrogen of water interacts with the partially negative oxygen in ethanol to form a hydrogen bond (Figure 4).

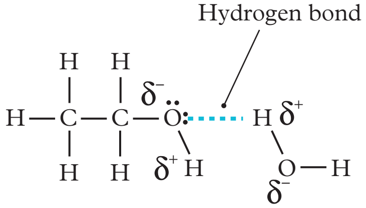


FIGURE 4 Hydrogen bonding between water and ethanol

Study tip

Use spaced practice! Now is a good time to review the types of substances and their bonding (from Module 6) and intermolecular forces (from Module 9). Molecular substances are covalent compounds that contain only non-metals. Ionic substances have ionic bonds which are electrostatic attractions between all the neighbouring positive and negative charges of their anions and cations.

Non-polar molecular substances can only interact with water solvent molecules through weak dispersion forces, so they are much less soluble in water. Some molecular solutes have a polar region and a non-polar region. The solubility of these substances depends on the size of the non-polar region.

The interaction between molecular substances and water is weaker than the interaction between ionic substances and water. This is because ionic substances have whole charges on their ions, whereas molecular substances are only partially positive and partially negative. Therefore, ionic substances generally have a greater solubility than molecular substances.

However, there are exceptions as some ionic substances are insoluble. Ionic substances that are insoluble generally have very strong ionic bonds holding them in their crystal lattice.

How does temperature affect the solubility of solids?

Solubility is defined as the maximum amount of solute that dissolves in a known quantity of solvent at a given temperature. Once the maximum amount of solute has been reached and no more can dissolve, the solution has become saturated.

If the temperature of a solution is increased, the solubility of a substance changes. An increase in temperature causes an increase in kinetic energy. This energy causes substances to become more or less soluble in water, depending on their:

- polarity – polar or non-polar
- state – solid, liquid or gas.

saturation point
the point at which the maximum amount of solute has dissolved in the known quantity of solute at a given temperature

Solubility curves

The relationship between temperature and the solubility of substances can be plotted on a solubility curve (Figure 5). A solubility curve is a graphical representation of the mass of solute that can dissolve in 100 g of water (y-axis) at a certain temperature (x-axis).

Several key pieces of information can be determined from a solubility curve.

- The curve represents the **saturation point** of a solution at a given temperature.
- A solution is unsaturated if it sits below the line.
- A solution is supersaturated if it sits above the line.
- A positive trend indicates that solubility increases as temperature increases.
- A negative trend indicates that solubility decreases as temperature increases.

Study tip

When a graph has a lot of data, use different colours or dashed lines to clearly represent all data.

Study tip

1 mL of water = 1 g of water

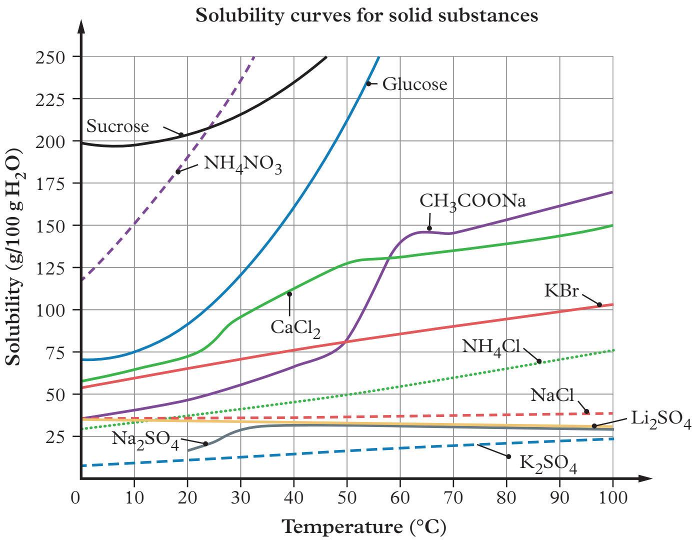


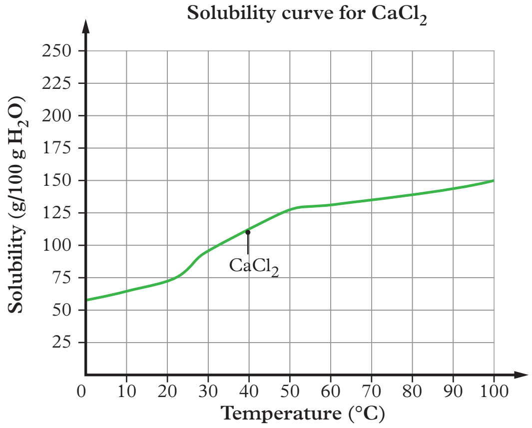
FIGURE 5 Solubility curves of some substances in water

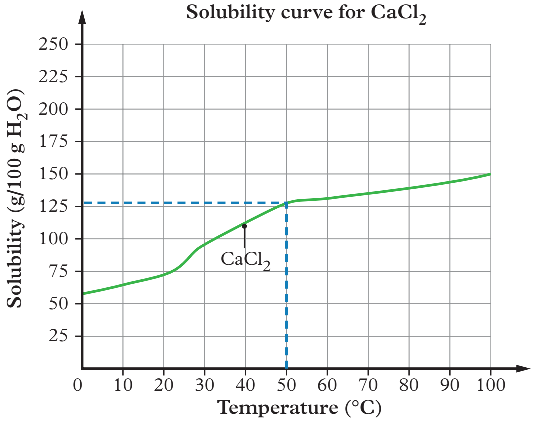
Worked example 13.3A

Interpreting a solubility curve

Worked example 13.3A: Watch a video that shows how to solve this problem.

At 50°C, 140 g of solid $CaCl_2$ is added to 100 mL of water. **Determine** whether all of the solute will dissolve. (2 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to read a solubility curve and make a decision about the solubility of the solute. The question is worth 2 marks, so we must interpret the curve and provide a decision.
Step 2: Look for $CaCl_2$ on the solubility curve in Figure 5.	The $CaCl_2$ line is green. 

Think	Do
Step 3: Find the appropriate temperature on the *x*-axis and read the solubility (g/100 g H_2O). Remember that 1 mL of water = 1 g of water.	At 50°C, a bit more than 125 g of the solute dissolves in 100 g (or 100 mL) of the solvent (water). (1 mark) 
Step 4: Finalise your answer.	Not all of the solute will dissolve. (1 mark)

Your turn

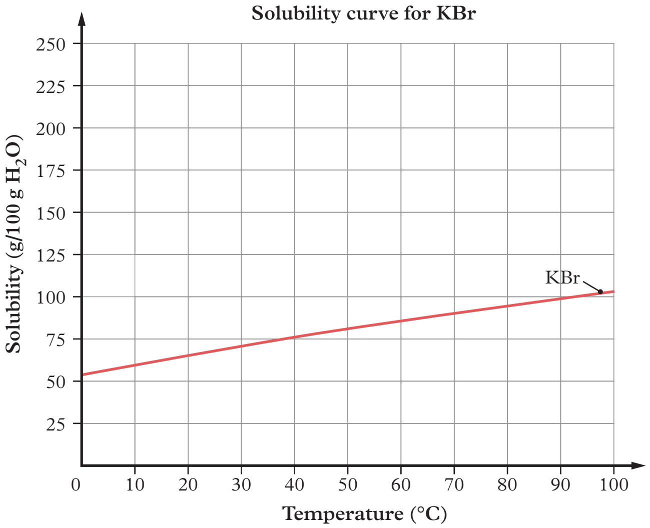
At 80°C, 100 g of KBr is added to 100 mL of water. **Determine** whether all of the solute will dissolve. (2 marks)

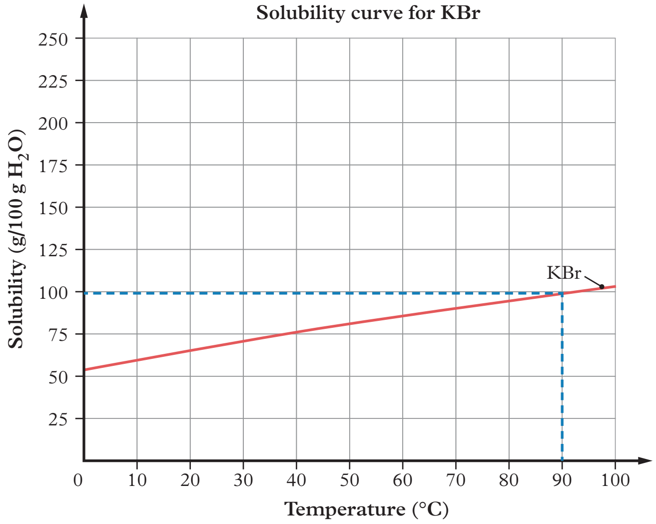
Worked example 13.3B

Determining the mass of dissolved solute

Worked example 13.3B: Watch a video that shows how to solve this problem.

At 90°C, 19.5 g of solid KBr is added to 20 mL of water. **Determine** the mass of solute that will dissolve. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to determine the mass of solute. The question is worth 1 mark, so we must read the graph and provide the correct value.
Step 2: Look for KBr on the solubility curve in Figure 5.	The KBr line is red. 

Think	Do
Step 3: Find the appropriate temperature on the *x*-axis and read the solubility (g/100 g H_2O). Remember that 1 mL of water = 1 g of water.	At 90°C, 100 g of the solute dissolves in 100 g (or 100 mL) of the solvent (water). 
Step 4: This is the mass of solute that dissolves in 100 g or 100 mL. Find the mass that would dissolve in the volume given in the question (20 mL).	$\frac{20\,\text{mL}}{100\,\text{mL}}$ × 100 g of solute = 20.0 g (which is more than 19.5 g) (1 mark)
Step 5: Finalise your answer.	19.5 g will dissolve. (1 mark)

Your turn

At 40°C, 31 g of glucose is mixed with 25 mL of water. **Determine** the mass of solute that will dissolve. (2 marks)

Worked example 13.3C

Calculating solubility

Worked example 13.3C: Watch a video that shows how to solve this problem.

At 21°C, 4.2 g of solute dissolves in 25 mL of water. **Calculate** its solubility in g/100 g of water. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the solubility. The question is worth 1 mark, so we must gather the correct information, use the values provided to complete the calculation.
Step 2: Calculate the mass dissolved in 100 g of water.	$\frac{4.2\,\text{g}}{25\,\text{mL}}$ × 100 mL = 16.8 g of solute
Step 3: Finalise your answer. Solubility in water should be expressed as mass per 100 g H_2O at a specific temperature.	16.8 g/100 g at 21°C (1 mark)

Your turn

At 35°C, 19.5 g of a solute dissolves in 30 mL of water. **Calculate** its solubility in g/100 g of water. (1 mark)

Worked example 13.3D

Calculating solubility

Worked example 13.3D: Watch a video that shows how to solve this problem.

3.6 mg of solute dissolves in 5.2 mL of water at 15°C. **Determine** its solubility in g/100 g of water. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the solubility. The question is worth 1 mark, so we must gather the correct information, use the values provided to complete the calculation.
Step 2: Calculate the mass of solute in 100 g of water.	$\frac{3.6\text{ mg}}{5.2\text{ mL}} \times \frac{1\text{ g}}{1000\text{ mg}} \times 100\text{ mL} = 0.07\text{ g}$ of solute
Step 3: Finalise your answer. Solubility in water should be expressed as mass per 100 g H_2O at a specific temperature.	0.07 g/100 g at 15°C (1 mark)

Your turn

1.45 kg of solute dissolves in 15.0 L of water at 20°C. **Determine** its solubility in g/100 g of water. (1 mark)

Challenge

The solubility of lithium sulfate

The solubility of lithium sulfate (Li_2SO_4) decreases as temperature increases. **Suggest** a reason why. (3 marks)

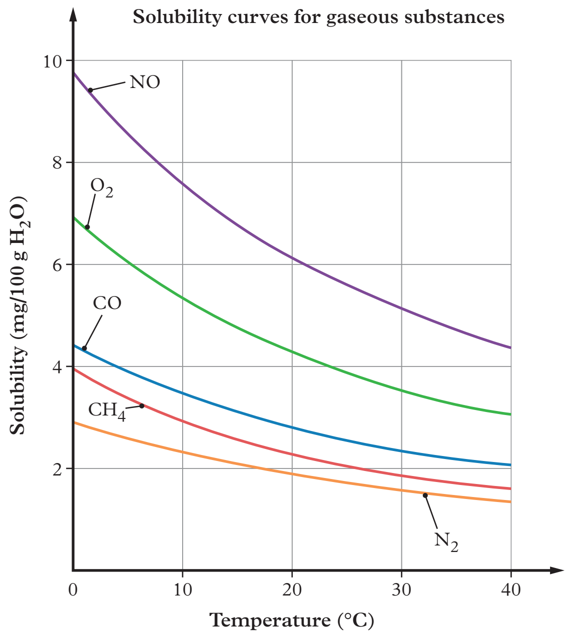


FIGURE 6 The solubility of gases in water decreases as temperature increases.

How does temperature affect the solubility of gases?

Atmospheric gases, such as nitrogen (N_2), oxygen (O_2) and carbon dioxide (CO_2), have very low solubilities in water. As temperature increases and kinetic energy increases, intermolecular forces between the dissolved gas particles and the solvent water molecules (solute–solvent bonds) are broken. Because there are very weak interactions between the water molecules and the gas molecules, these forces break more easily than the interactions between water molecules, and therefore dissolved gases evaporate more easily than water.

The best example of this phenomenon is water boiling. As the temperature of water increases, the solubility of various gases in the water decreases, and these gases form gas bubbles well before the boiling temperature is reached. Contrast this with the boiling point, when water vapour forms bubbles under the surface.

Gas particles are more soluble in solvents under high pressures. A soda stream pumps carbon dioxide gas into a bottle at high pressures, which forces the gas to dissolve in the water solvent to form fizzy water. When the bottle has been stored for a period of time, the pressure inside the bottle increases as a few of the gas particles come out of solution. When the lid of the bottle is unscrewed, the pressure above the drink is released. The lowering of the pressure on the surface allows bubbles of gas to form and rise to the top of the drink, reducing the amount of gas dissolved in the drink.

Real-world chemistry

The chemistry of blood

Blood is a complex mixture of cells and dissolved substances. It can be easily separated using a centrifuge where the blood is spun to force heavier components to the bottom of the vial and lighter ones to the top. This allows scientists to analyse the composition of blood. This information is useful for analysing various bodily processes to determine whether the body is functioning normally or whether medical intervention is required.

Blood separates into three layers (Figure 7).

1 The bottom layer consists of red blood cells. It is the densest component of blood and makes up 45% of its total volume. The primary roles of red blood cells are to carry oxygen around the body and to remove carbon dioxide.

2 The middle layer consists of white blood cells and platelets, occupying less than 1% of the volume. White blood cells are important for the immune function because they fight infection, and attack viruses and bacteria that make their way into the body.

3 The top layer is plasma, which is the least dense component, occupying 55% of the blood's total volume. It contains many solutes that are carried around the body for various functions. Blood plasma is 92% water.

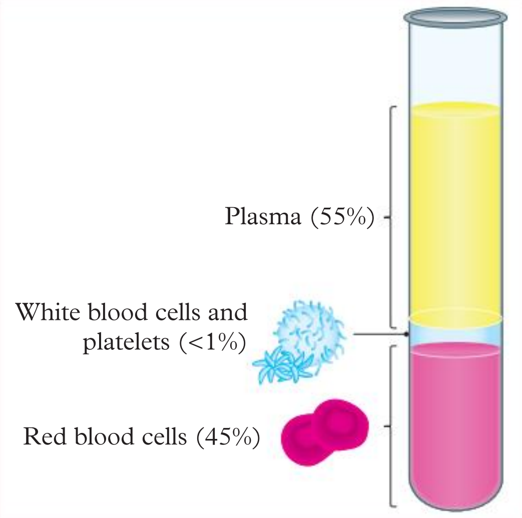


FIGURE 7 A vial of blood after being spun by a centrifuge to separate its components.

Solubility of substances in plasma

Any substance that is carried through the body in the blood's plasma must be highly soluble in water at body temperature (37°C). Transport must happen quickly to provide the body with the nutrition that it requires, keeping it functioning and healthy. Some of the compounds dissolved in blood plasma are summarised in Table 1.

TABLE 1 Compounds dissolved in blood plasma

Compound	Description
Protein	Makes up 8% of blood plasma; have a variety of functions, including the transport of molecules and the function of the immune system
Gas	Dissolved in small amounts in blood plasma for transfer around the body
Electrolytes/minerals	Form ions in solution and can conduct electricity; for example, sodium is the most abundant ion in the body and has many roles, including nerve and muscle cell function
Nutrients	Molecules that the body cannot synthesise on its own and must be obtained from food; some, like glucose, dissolve readily in blood and are transported easily around the body

Apply your understanding

1 **Identify** the component of plasma that can influence its ability to act as a solvent. **Explain** how it dissolves solutes. (2 marks)

2 **Investigate** why gases dissolve more readily in blood than in water. (1 mark)

Check your learning 13.3

Check your learning 13.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** why most substances are more soluble as temperature increases. (4 marks)

2 **Identify** two reasons why a substance may be insoluble in water. (2 marks)

Analytical processes

3 **Sequence** the following substances by their solubility in water, from greatest to least: non-polar molecular substance, ionic substance and polar molecular substance. (1 mark)

4 **Analyse** the solubility curves (Figure 5).

- **a** **Determine** the:
  - **i** solubility of NH_4Cl at 70°C (1 mark)
  - **ii** solubility of $NaNO_3$ at 50°C (1 mark)
  - **iii** mass of CH_3COONa dissolved in 30 mL of water at 40°C (2 marks)
  - **iv** mass of $CaCl_2$ dissolved in 2.5 mL of water at 10°C. (2 marks)
- **b** **Describe** how a student could make a supersaturated solution of KCl. (3 marks)

5 **Consider** what the line on solubility curve represents. **Interpret** what a line in the upward direction indicates. (2 marks)

6 150 g of KBr was added to 100 mL of water, and the mixture was heated to 90°C. **Analyse** the solubility curve (Figure 5) for KBr to answer the following questions.

- **a** **Determine** the mass of solid that would remain undissolved. (2 marks)
- **b** The mixture was then cooled to 65°C and filtered to remove all of the solid that had formed at that temperature. **Determine** the total mass of solid obtained. (2 marks)
- **c** The filtrate was further cooled to 40°C. **Determine** the total mass of undissolved solid present. (3 marks)

Knowledge utilisation

7 Ammonia (NH_3), sulfur dioxide (SO_2) and hydrogen sulfide (H_2S) are polar molecules that are gases at room temperature. **Predict** whether their solubility curves would differ from that of a non-polar molecule. **Justify** your response. (3 marks)

8 **Research** the properties of the three materials: honey, baby oil and milk. **Use** your knowledge of non-polar liquid substances to **propose** the composition and the order of the layers that will form when added to water. (4 marks)

Practical

Lesson 13.4

Investigating the effect of temperature on solubility

Learning intentions and success criteria

Video demonstration

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

MODULE 13

Lesson 13.5

Review: Solubility and identifying ions in solution

Summary

13.1
- A precipitate is an insoluble solid that forms from reactions between some ionic compounds.
- Precipitation reactions can be represented using balanced full and ionic chemical equations with state symbols (s), (l), (g) and (aq) included.
- A solubility table can be used to predict and identify whether a precipitate will form from chemical reactions in aqueous solutions.

13.2
- Practical: Investigating precipitation reactions

13.3
- Solutes dissolve in solvents that have similar properties.
- In general, an increase in temperature (and an increase in kinetic energy) results in an increase in solubility for ionic solids.
- In general, gases become less soluble with increasing temperature.

13.4
- Practical: Investigating the effect of temperature on solubility

Review questions 13.5A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Which one of the following ionic substances is insoluble in water?

A NaOH
B KCl
C AgCl
D $NaNO_3$

2 Non-polar molecules are the least soluble in water because they contain

A dispersion forces, creating a temporary dipole and therefore weaker attraction between water molecules.
B a permanent dipole but a large amount of kinetic energy and therefore weaker attraction to water molecules.
C ionic interactions, but these become weaker as they dissolve because cations and anions move away from one another in the water.
D dispersion forces, creating a permanent dipole and therefore weaker attraction between water molecules.

3 Which one of the following ionic substances is soluble in water?

A $HgCl_2$
B $BaSO_4$
C $Pb(OH)_2$
D K_2SO_4

4 Ionic substances generally have the highest solubility in water because they

A are polar molecules containing partial charges.
B are non-polar molecules.
C are polar molecules containing whole charges.
D contain whole charges on ions.

5 The solubility of most solid substances in water

A increases as temperature increases.
B increases as temperature decreases.
C decreases as temperature increases.
D does not depend on temperature.

6 Which of the following is not a molecular compound?

A H_2O
B NF_3
C H_2S
D Fe_2O_3

7 Which of the following is not a rule for the solubility of salts in water?

A Most chloride (Cl–) salts are soluble.

B Nitrate (NO_3^-) salts generally insoluble.

C Most salts of Na^+ and NH^{4+} are soluble.

D Most sulfate (SO_4^{2-}) salts are soluble.

8 Consider the three molecular substances: propane (C_3H_8), 1-propanol (C_3H_7OH) and propanal (C_3H_6O). In order of increasing solubility in water, they would be

A propane, 1-propanol, propanal.

B propanal, propane, 1-propanol.

C propane, propanal, 1-propanol.

D 1-propanol, propanal, propane.

9 Dissolved ions are surrounded by and attracted to a cluster of water molecules. This is best described as

A hydration of ions.

B solution of ions.

C dissociation of ions.

D dipole–dipole attractions.

Review questions 13.5B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

10 **Explain** why some molecular substances are soluble in water, while others form layers with water. (2 marks)

11 **Explain** why a precipitation reaction only forms a solid and no other product, using an example. (2 marks)

12 **Explain** why the solubility of substances is often converted to or expressed in g/100 g of water. (1 mark)

Analytical processes

13 The structures of glucose and sucrose are shown.

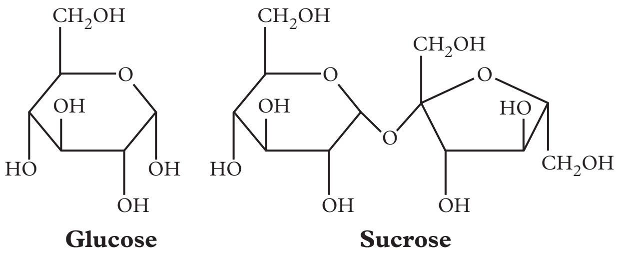


FIGURE 1 The structures of glucose and sucrose

Use the solubility curve and **compare** the structures of these molecules to **explain** the difference in their solubility data. (3 marks)

14 KBr and K_2SO_4 both contain potassium cations but have very different solubility data. **Consider** what differences between these substances are causing this variation. **Deduce** why KBr has a higher solubility than that of K_2SO_4. (3 marks)

15 **Identify** the products of the following precipitation reactions and **determine** whether they are soluble or insoluble. (2 marks)

a sodium sulfate + barium nitrate (2 marks)

b lead(II) nitrate + sodium hydroxide (2 marks)

c aluminium chloride + potassium hydroxide (2 marks)

16 **Analyse** the solubility curve to **determine** the:

a solubility of NH_4NO_3 at 20°C (1 mark)

b solubility of K_2SO_4 at 80°C (1 mark)

c temperature required to produce a saturated solution of glucose if 25 g of glucose is dissolved in 20 mL of water (2 marks)

d solubility of sucrose in 40 mL of water at 30°C (2 marks)

e solubility of $CaCl_2$ in 9 mL of water at 40°C. (2 marks)

17 **Explain** how a student could make a supersaturated solution of NH_4NO_3 at 20°C. (3 marks)

Knowledge utilisation

18 A student claims that a supersaturated solution cannot be made in a laboratory without a hot plate. **Evaluate** this claim and respond using the solubility rules. (2 marks)

19 **Devise** a method to determine whether carbonate, bromide, aluminium and calcium ions are present in an unknown sample. Present your method as numbered steps, detailing the equipment that could be used in a laboratory. (4 marks)

20 A scientist has reached the final step in the determination of unknown ions in a sample of wastewater. The scientist cannot distinguish between chloride, iodide and bromide ions because they are all potentially present in the same solution and their distinctive colouring is unclear.

Evaluate whether the scientist could use colour to draw any conclusions. **Investigate** other techniques for distinguishing between these ions. (5 marks)

21 **Investigate** which sports drinks contain which electrolytes. **Use** this information to **determine** the best sports drink for replacing potassium in the human body. **Discuss** whether there are other alternatives for humans to replace potassium in their bodies. (3 marks)

22 A student finds two solutions on a laboratory bench. The solutions are marked A and B. The student knows, from earlier in the lesson, that the solutions are salt water or sugar water, but does not remember which is which. To identify them, they pour a small amount of A into a beaker and then slowly pour in the same volume of solution B. The student notices that the solutions form two layers, where solution A sits on top of solution B. The salt water was made by mixing 15.25 g of NaCl in 50 mL of water. The sugar solution was made by dissolving 50 g of sugar in 62.5 mL of water. **Determine** the identities of solutions A and B. **Use** calculations to **justify** your answer. (4 marks)

Data drill

Solubility of gases

To determine the solubility of carbon dioxide, a student uses a soda stream. They start by measuring the mass of a 1 L bottle and then filling it with water and reweighing it to determine the mass of the water in the bottle. They then use the soda stream to add carbon dioxide to the water until no more will dissolve into the solution and reweighed the bottle. This experiment was repeated over several days, each of which had a different maximum temperature Table 1.

TABLE 1 Results from the solubility experiment

Temperature of the bottle (°C)	Mass of the bottle (g)	Initial mass of the bottle and water (g)	Final mass of bottle and water (g)
18	145	1,133.7	1,134.7
25	145	1,138.2	1,139.0
30	145	1,135.9	1,136.6

Apply understanding

1 **Calculate** the mass of the carbon dioxide dissolved at 25°C. (1 mark)

2 **Identify** the temperature at which the most carbon dioxide was dissolved. (1 mark)

Analyse data

3 **Organise** the data from Table 1 into a solubility curve. (2 marks)

4 **Identify** the trend in the data. (1 mark)

Evaluate evidence

5 **Extrapolate** the data to **predict** the the mass of carbon dioxide that is soluble in water at 10°C. (2 marks)

6 The experiment did not control the mass of water. **Deduce** the size of the impact this has on the results and **justify**. (2 marks)

Module 13 checklist: Solubility and identifying ions in solution

Quizlet: Revise key terms online to test your understanding

MODULE 14

Acids and bases

Introduction

The word "acid" comes from the Latin *acidus*, meaning sour. Acids may be concentrated or dilute, strong or weak. Weak acids such as the acetic acid in vinegar can be dangerous if they are highly concentrated. A base is often thought of as the opposite of an acid, and many people underestimate the toxicity and corrosive nature of a base. The pH scale is used as a measure of the concentration of hydrogen ions in aqueous solutions.

Prior knowledge

Prior knowledge quiz

Check your understanding of acids, bases and the pH scale before you start.

Subject matter

Science understanding

→ State that pH is dependent on the concentration of hydrogen ions in solution.
→ Identify that the pH scale is a logarithmic scale.
→ Apply the pH scale to compare the levels of acidity or alkalinity of aqueous solutions.
→ Apply the Arrhenius model to explain the behaviour of strong and weak acids and bases in aqueous solutions.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Lesson 14.1

The Arrhenius model of acids and bases

Key ideas

→ Arrhenius acids increase the concentration of hydrogen ions in water and Arrhenius bases increase the concentration of hydroxide ions in water.

→ The strength of an acid or base depends on how easily it ionises in water. Acids and bases that ionise completely are strong, whereas those that undergo partial ionisation are weak.

→ The concentration of an acid or base depends on the number of acid or base molecules, respectively, in a specified volume of an aqueous solution.

Learning intentions and success criteria

What is the Arrhenius model of acids and bases?

acid
a chemical compound that donates a hydrogen ion

base
a chemical compound that accepts a hydrogen ion

Lemons, vinegar and soft drinks all contain **acids**. Acids (and their "partner" bases) were first described by Robert Boyle in 1680 – he described how some substances changed the colour of some natural dyes. In the 1700s, acids were described as having a sour taste and the ability to react with limestone to produce a gas. Today, acids such as citric acid and acetic (ethanoic) acid are often used in lollies to intentionally produce the popular sour flavour (Figure 1).

Often, **bases** are considered the matching partner to acids. Bases, such as sodium hydroxide, are bitter to taste and feel slippery. Like acids, bases are corrosive and can produce chemical burns on skin.

In chemistry, we need better methods to test acids and bases because it is not feasible or wise to taste every chemical to determine whether they are an acid or a base. In 1884, a scientist named Svante Arrhenius examined a group of materials with the properties of acids and proposed the idea that when they were dissolved in water, the solutions had very high concentrations of hydrogen ions (H^+). As a result, Arrhenius defined acids as the group of chemicals that can produce a high concentration of hydrogen ions in an aqueous solution.

FIGURE 1 Sour lollies can contain a variety of different acids.

FIGURE 2 Most detergents and toothpaste are bases that dissolve in water (alkali).

Using the Arrhenius definition of an acid, when hydrochloric acid (HCl) gas is dissolved in water, the HCl molecules are assumed to immediately **ionise** to form H^+ (hydrogen ions) and Cl^- (chloride ions):

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Once Arrhenius had defined an acid as a material that was able to produce a high concentration of hydrogen ions when mixed with water, he started working on the group of chemicals known as bases. He found that when a base was dissolved in water to form an alkali ne solution, it resulted in a high concentration of hydroxide ions (OH^-).

One of the most common examples of this is when sodium hydroxide (NaOH) is mixed with water. Sodium hydroxide is an ionic compound, made up of sodium ions (Na^+) and hydroxide ions (OH^-). Arrhenius tested many different bases that behaved in the same way and concluded that all bases increase the concentration of hydroxide ions in a water solution:

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

A very common misconception in chemistry is that the concentration of a solution in some way correlates to its strength. Specifically, strength refers to acid and base chemistry. However, in the context of aqueous solutions and concentration, it is important to understand the distinction between strength and concentration of an acidic or basic solution. In chemistry, “strong” and “weak” do not have the same meaning as in everyday life.

ionise
form ions

Study tip

A hydrogen atom consists of a single positive proton that has a single negative electron moving in the space around it. If the electron is removed (when a hydrogen ion is formed), only the proton remains. For this reason, a hydrogen ion is also known as a proton.

What is the concentration of acids and bases?

Remember that concentrated solutions contain more solute than the same volume of dilute solutions. Therefore, 1 L of a concentrated salt solution contains more Na^+ and Cl^- ions than 1 L of a dilute salt solution. The same holds true for a concentrated acid or base solution. A concentrated solution of hydrochloric acid has a high number of H^+ and Cl^- ions. A dilute solution of sodium hydroxide contains a low number of Na^+ and OH^- ions. This does not change how readily the acid or base molecules release hydrogen ions or hydroxide ions. Instead, it is a measure of the number of acid or base particles there are in a set volume of solution.

What is the strength of acids and bases?

Acids and bases are often described as weak or strong. For an acid, these terms refer to the extent to which the acid reacts with water to form hydrogen ions. For a base, they refer to the extent to which the base reacts with water to form hydroxide ions.

FIGURE 3 The intensity of blue colour of the liquid detergent indicates its concentration. In the bottle, it is more concentrated and therefore has a darker hue, whereas the diluted volume in the jug is lighter, and therefore less concentrated.

Study tip

Dissociation to form ions is not the same as ionisation. Dissociation occurs when an ionic compound separates into its separate ions. Ionisation occurs when the molecule reacts with water to produce ions.

Strong acids and bases

The strength of an acid refers to its ability to ionise and form its cation (positively charged ion) and anion (negatively charged ion) in water. This process is called ionisation.

strong acid
an acid that completely ionises in water

Hydrochloric acid is a **strong acid**. In water, HCl molecules ionise according to the chemical equation:

$$HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

As HCl ionises completely into ions in water, no HCl remains. This is what makes HCl a strong acid.

strong base
a base that completely dissociates in water

The same principle applies for bases. Sodium hydroxide is a **strong base** because it ionises completely in water according to the chemical equation:

$$NaOH(aq) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

Weak acids and bases

weak acid
an acid that only partially dissociates in water

Ethanoic acid (CH_3COOH), commonly known as acetic acid, is a **weak acid**. It is safe for human consumption as a dilute solution and is the acid in vinegar. It partially dissociates in water according to the equation:

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^{-}(aq) + H^{+}(aq)$$

The main difference between the chemical equations used to represent complete and partial ionisation is the arrow ($\rightleftharpoons$) used. In partial ionisation, there are both forward and back reactions occurring simultaneously. So, CH_3COOH begins to ionise, driving the reaction forward and producing CH_3COO^- and H^+ ions. In the reverse or back reaction, CH_3COO^- and H^+ react together to form CH_3COOH.

As the concentration of H^+ ions formed in this reaction is very small, CH_3COOH is a weak acid. In other words, the reaction goes in the forward direction to only a very small extent.

weak base
a base that only partially ionises in water

Ammonia (NH_3) is a **weak base**. The Arrhenius model cannot adequately explain why a solution of ammonia is a base, as it does not contain OH^- ions. The reason it behaves as a base is that when it dissolves in water, it reacts and partially ionises according to the chemical equation:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^{+}(aq) + OH^{-}(aq)$$

As the concentration of OH^- ions formed is very low, the basicity of the solution is much less than if an equivalent amount of the strong base sodium hydroxide has been added. Therefore, ammonia is considered a weak base.

TABLE 1 Some common weak and strong acids and bases

Strong acids	Weak acids	Strong bases	Weak bases
Hydrochloric acid (HCl)	Ethanoic acid (CH_3COOH)	Sodium hydroxide (NaOH)	Ammonia (NH_3)
Nitric acid (HNO_3)	Carbonic acid (H_2CO_3)	Potassium hydroxide (KOH)	Methanamine (CH_3NH_2)
Sulfuric acid (H_2SO_4)	Hydrofluoric acid (HF)	Lithium hydroxide (LiOH)	Sodium carbonate (Na_2CO_3)

Challenge

Ionisation of a triprotic acid

Phosphoric acid (H_3PO_4) is a weak triprotic acid because it contains three hydrogen ions (H^+). This means that it undergoes ionisation three times to lose each of these ions to eventually form the phosphate ion (PO_4^{3-}). **Determine** three balanced chemical equations for the ionisation of phosphoric acid in water. (3 marks)

What is the difference between strength and concentration?

The confusion between "strong" and "concentrated" can be easily understood. When a drink, such as cordial, is concentrated, it has a stronger taste to it and this may be the basis for the misconception. In chemistry, "strength" and "concentration" are very separate concepts with different meanings and cannot be treated in the same way.

- A strong acid can be concentrated. This means that there are many units of the acid present and that all of these molecules are ionised in water.
- A strong acid can also be dilute. This means that there are only a few units of the acid present and that all of these molecules are ionised in water.
- A weak acid can be concentrated. This means that there are many units of the acid present but only a few of these molecules are ionised in water.
- A weak acid can also be dilute. This means that there are only a few units of the acid present and only very few of these are ionised in water.

This can also be applied to bases.

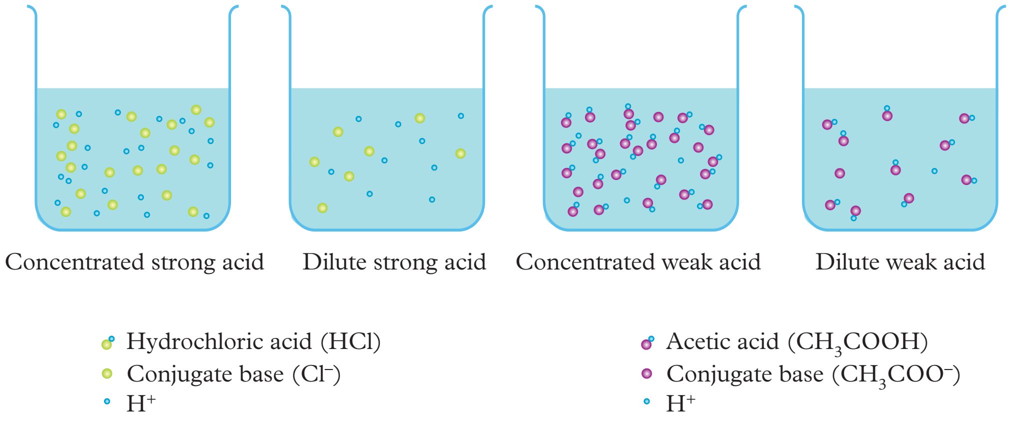


FIGURE 4 Acids can be strong or weak, and concentrated or dilute.

What are the limitations of the Arrhenius model?

Despite the many advantages of the Arrhenius model of acids and bases, some substances do not fit this model. The Arrhenius model requires all acids and bases to be able to produce hydrogen ions or hydroxide ions respectively. However, some substances (e.g. NH_3), can still act as a base despite not containing hydroxide ions. Other compounds (e.g. SO_3) do not contain a hydrogen but react with water to release a hydrogen ion.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$
$$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$$

As a result, two scientists Brønsted and Lowry independently suggested a modification of the Arrhenius model. They suggested that acids release hydrogen ions and bases accept and bond with hydrogen ions (protons). The Brønsted–Lowry model is covered in Unit 3.

Real-world chemistry

Not all hydrogen leads to acid

A molecule with a large number of hydrogen atoms is not necessarily an acid. Methane (CH_4) has four hydrogen atoms covalently attached to the carbon atom. The covalent bond is very strong, meaning the hydrogen atoms are not easily released when mixed with water. Similarly, glycerol (Figure 5) has three OH groups in its molecule, but the OH groups do not form hydroxide ions. This means glycerol does not act as a base.

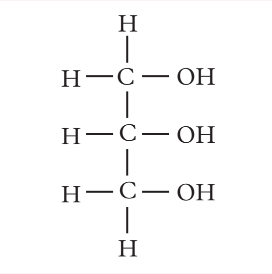


FIGURE 5 Glycerol does not act as a base, even though it has three OH groups.

Apply your understanding

1 **Explain** why most covalent compounds are unlikely to ionise. (2 marks)

2 **Infer** whether all ionic compounds act as bases. **Justify** your answer. (2 marks)

Check your learning 14.1

Check your learning 14.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Identify** the properties of:
  - a an acid (2 marks)
  - b a base. (2 marks)

2 **Explain** how Arrhenius defined:
  - a an acid (1 mark)
  - b a base. (1 mark)

Analytical processes

3 A list of properties of acids and bases is shown. **Use** a Venn diagram to **classify** each property as belonging to an acid, a base, or both. (3 marks)
  - sour taste
  - bitter taste
  - corrosive
  - feels slippery
  - produces a chemical burn on skin
  - produced in the stomach
  - ionises in water
  - dissociates in water.

4 **Use** the Arrhenius model to **differentiate** between a dilute strong acid and a concentrated weak acid. (4 marks)

Knowledge utilisation

5 Ammonia (NH_3) is often described as a base even though it does not contain hydroxide ions. **Discuss** how the Arrhenius model has been modified to account for this observation. (2 marks)

Lesson 14.2

The pH scale

Key ideas

→ Aqueous solutions with a pH below 7 are acidic. A pH of 7 indicates a neutral aqueous solution, and a pH above 7 indicates a basic aqueous solution.

→ Acidic solutions with low pH values have high concentrations of hydrogen ions.

→ Indicators can be used to identify the hydrogen ion concentration of a solution or its pH.

Learning intentions and success criteria

What is the pH scale?

Aqueous solutions of acids and bases can be described based on their **pH**. This provides a measure of the concentration of hydrogen ions present in the solution. The **pH scale**, a numerical system, can then be used to determine whether an aqueous solution is acidic or basic. Most commonly, the pH scale is measured between 0 and 14. Pure water is considered **neutral** (neither acidic nor basic) and is assigned a pH of 7.

Acidic solutions have a pH of less than 7. The lower the pH value, the more acidic the solution and the more concentrated the hydrogen ions in that solution. Most detergents, such as ammonia and caustic soda, are basic and have a pH of greater than 7. Chlorine bleach is a basic solution of sodium hypochlorite (NaOCl) with a pH of 12.

pH
a qualitative measure of hydrogen ion concentration

pH scale
a numerical scale that is used to determine the level of acid or base in an aqueous solution

neutral
an aqueous solution that has a pH of 7

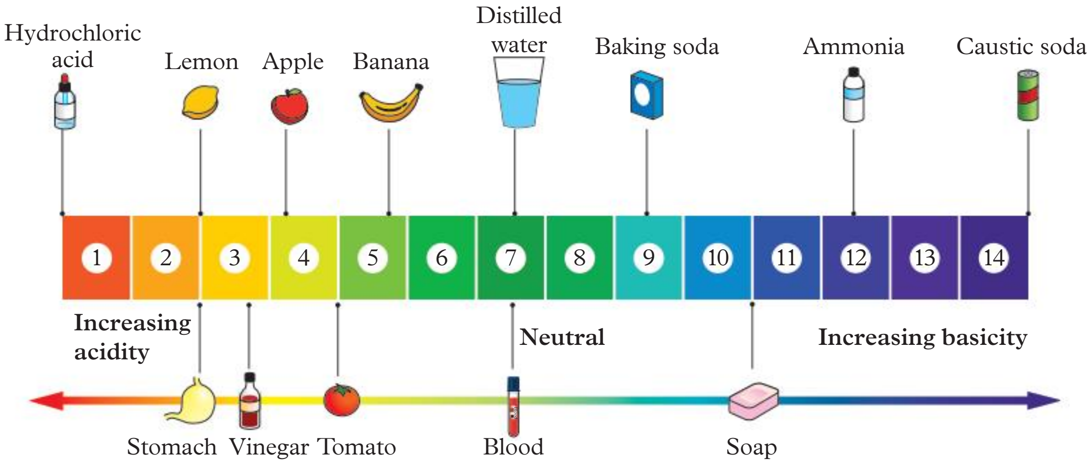


FIGURE 1 The pH scale is used to report the concentration of hydrogen ions in an aqueous solution. Acidic solutions have a pH below 7, basic solutions have a pH above 7 and neutral solutions have a pH of 7.

How is pH calculated?

According to the Arrhenius model, acids are substances that increase the concentration of hydrogen ions in aqueous solutions. Strong acids react completely to form hydrogen ions, while weak acids only partially react. In the Arrhenius model of acids and bases, bases have high concentrations of hydroxide ions. These hydroxide ions can bind to free hydrogen ions to form water. As a result, the concentration of hydrogen ions decreases.

The concentration of free hydrogen ions can therefore be used to indicate the “power” of the acids and bases. Many sources quote pH as meaning “power of hydrogen” or *puissance d’hydrogen* (French for “strength of hydrogen”). Alternatively, the “p” is thought to be a mathematical symbol that refers to the negative logarithmic scale, and the H refers to the concentration of hydrogen. Therefore, pH is the negative logarithmic scale of hydrogen ions:

$$pH = -\log_{10} [H^+]$$

where $[H^+]$ is the molar concentration of hydrogen ions.

TABLE 1 The pH of a solution is an indication of the hydrogen concentration

	Hydrogen concentration (M)	pH
Highly acidic solution	$0.1 = 10^{-1}$	1
Weakly acidic solution	$0.000001 = 10^{-6}$	6
Neutral	$0.0000001 = 10^{-7}$	7
Weakly basic solution	$0.000000001 = 10^{-9}$	9
Highly basic solution	$0.00000000000001 = 10^{-14}$	14

Using a $\log_{10}$ scale to convert hydrogen ion concentration to a pH scale means that each pH value has a hydrogen ion concentration 10 times larger or smaller than the pH value one pH unit away. The equation for pH can be rearranged to determine the concentration of hydrogen ions in a solution:

$$[H^+] = 10^{-pH}\ (M)$$

An aqueous solution with a pH of 2 has ten times more hydrogen ions in the same volume than a solution with a pH of 3. The lower the pH value, the higher the hydrogen ion concentration.

Consider the example below.

$$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$$

In this example, one mole of HNO_3 will ionise to form one mole of hydrogen ions, and therefore they have the same number of moles. As such, they will have the same concentration as the HNO_3 ionises in the same volume that the hydrogen forms.

Study tip

A change in pH from 0 to 14 covers a change in hydrogen ion concentration from 1 M to 1×10^{-14} M, i.e. 14 orders of magnitude. The pH scale is a useful way to compress large changes in concentration, so they are easier to think about.

Worked example 14.2A

Calculating pH from hydrogen ion concentration

Worked example 14.2A: Watch a video that shows how to solve this problem.

Calculate the pH of a solution containing a strong acid if the H^+ concentration is 0.0010 M. (1 mark)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Calculate” means to determine or find a number or answer by using mathematical processes. We need to determine the pH. The question is worth 1 mark, so we must gather the correct information, use the values provided to complete the calculation.
Step 2: Select the correct formula and gather any data required.	$pH = -\log_{10}[H^+]$ $[H^+] = 0.0010\,M, pH = ?$
Step 3: Substitute the known values into the formula and solve for the unknown value.	$pH = -\log_{10}(0.0010)$ $= 3$

Think	Do
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures. pH doesn't have units.	3.0 (1 mark)

Your turn

Calculate the pH of an acidic solution if it has a H^+ concentration of 1×10^{-2} M. (1 mark)

Worked example 14.2B

Calculating pH of an acid from its mass

Worked example 14.2B: Watch a video that shows how to solve this problem.

A 1.50 g sample of HNO_3 is dissolved in water to make 200 mL of solution. **Calculate** the pH of the solution. (5 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the pH. The question is worth 5 marks, so we must gather the correct information, use the values provided to complete the calculation.
Step 2: Select the correct formula and gather any data required.	$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$, so $[HNO_3] = [H^+]$ $n = \frac{m}{M}$ $n = ?$, $m = 1.50\,\text{g}$, $M = 63.02\,\text{g}\,\text{mol}^{-1}$ $c = \frac{n}{V}$ $c = ?$, $V = 200\,\text{mL}$ $\text{pH} = -\log_{10}[H^+]$ $\text{pH} = ?$
Step 3: Substitute the known values into the formula and solve for the number of moles.	$n(HNO_3) = \frac{1.50\,\text{g}}{63.02\,\text{g}\,\text{mol}^{-1}}$ (1 mark) $= 0.0238\,\text{mol}$ (1 mark)
Step 4: Substitute the known values into the formula and solve for the concentrations.	$[HNO_3] = \frac{0.0238\,\text{mol}}{\frac{200\,\text{mL}}{1{,}000}}$ (1 mark) $= 0.119\,\text{M}$ $[H^+] = [HNO_3] = 0.119\,\text{M}$ (1 mark)
Step 5: Substitute the known values into the formula and solve for the pH.	$\text{pH} = -\log_{10}(0.119)$ $= 0.9244$
Step 6: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures. pH doesn't have units.	0.924 (1 mark)

Your turn

A 0.35 g sample of HCl is dissolved in 120 mL of water. **Calculate** the pH of the solution. (5 marks)

Challenge

pH and flower colour

Two colours of hydrangea flowers act as an indicator for the pH of the soil. Hydrangeas have different coloured flowers depending on the presence of aluminium ions and the acidity of the soil.

In a neutral or basic soil ($pH \geq 7$), aluminium ions in the soil form aluminium hydroxide. Aluminium hydroxide cannot move into the roots of the plant. Hydrangeas in this soil type will have blue flowers. In acidic soil ($pH < 7$), there are few hydroxide ions. This leaves the aluminium ions free to form other complex molecules that can move into the plant, reacting with pigments in the petals to produce pink-red flowers. **Deduce** the colour that the flowers of a hydrangea will be if there is no aluminium in acidic soil. **Justify** your response. (2 marks)

FIGURE 2 The colour of hydrangea flowers depends on soil acidity and the presence of aluminium ions.

How is the strength of an acid or base measured?

A simple method for determining the strength of an acid or base is to measure its conductivity and compare to other acidic or basic solutions of the same concentration. Strong acids completely dissociate, so they contain large numbers of charged ions. More charged ions increase the conductivity of the solution. When connected to an electric circuit, this results in a higher current and a brighter light bulb due to the higher number of changed ions (Figure 3). A lower conductivity implies that there are very few ions in solution, indicating a weak acid.

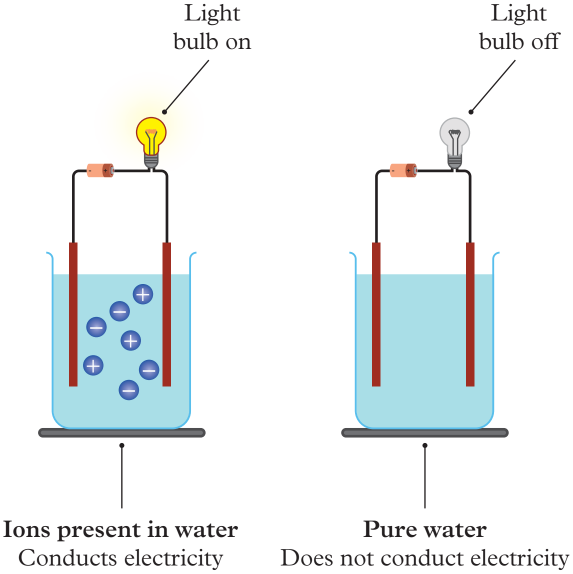


FIGURE 3 The strength of an acid or base can be determined by measuring the electrical conductivity of the aqueous solutions of the acid or base.

The strength of an acid or base can also be determined by measuring its pH with an indicator and comparing to standards.

What are indicators?

indicator
a substance that changes colour in the presence of high and low hydrogen ion concentrations

Indicators are dyes that have different coloured forms when they are protonated (have a H^+) and unprotonated (have lost a H^+). The indicator may be in solution or be infused into paper or a test strip. Indicators are weak acids. Different indicators change colour in different pH ranges (Figure 4).

$$\underset{\text{Colour 1}}{\text{H–indicator}} \rightleftharpoons H^+ + \underset{\text{Colour 2}}{\text{indicator}}$$

These colour changes can be used to indicate the pH of the solution. Many common indicators were originally purified from plants. The juice of a red cabbage, for example, is green in a basic solution, blue in a neutral solution and red in an acidic solution.

Indicator	pH range
Methyl violet	0.15–3.2
Thymol blue	1.2–2.8 and 8.0–9.6
Methyl orange	3.2–4.4
Methyl red	4.8–6.0
Bromothymol blue	6.0–7.6
Phenolphthalein	8.2–10.0
Alizarin yellow	10.2–12.0

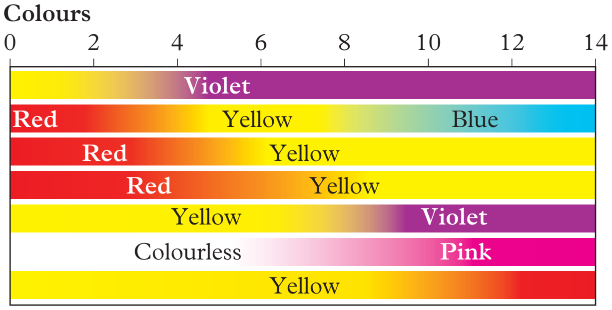


FIGURE 4 Different indicators change colour at different pH values.

Universal indicator

FIGURE 5 pH is indicated by the colours of universal indicator in acidic and basic solutions.

One of the most common indicators is universal indicator. Universal indicator is a mixture of many different indicators, including methyl red, bromothymol blue and phenolphthalein. This means it changes colours several times over a wide pH range. It is red in highly acidic solutions (pH 0 to 3), yellow in weakly acidic or very dilute strongly acidic solutions (3 to 6), green at neutral pH (7), blue in a weakly basic or very dilute strongly basic solutions (8 to 11) and purple in a strongly basic solution (11 to 14).

Universal indicator changes colour in a solution depending on the concentration of hydrogen ions in solution (Figure 4). A highly acidic solution has a pH of less than 2. Neutral solutions, such as water, have a pH of 7. Strongly basic solutions can have a pH of 13 to 14 and less concentrated solutions of bases can have a pH of 8 to 12.

Litmus paper

Litmus paper can be used to determine whether a solution is either acidic or basic. Red litmus paper turns blue in basic solution. Similarly, blue litmus paper turns red in acidic solution (Figure 6). Because litmus paper does not differentiate the strength of an acid, it is not often used in higher levels of chemistry.

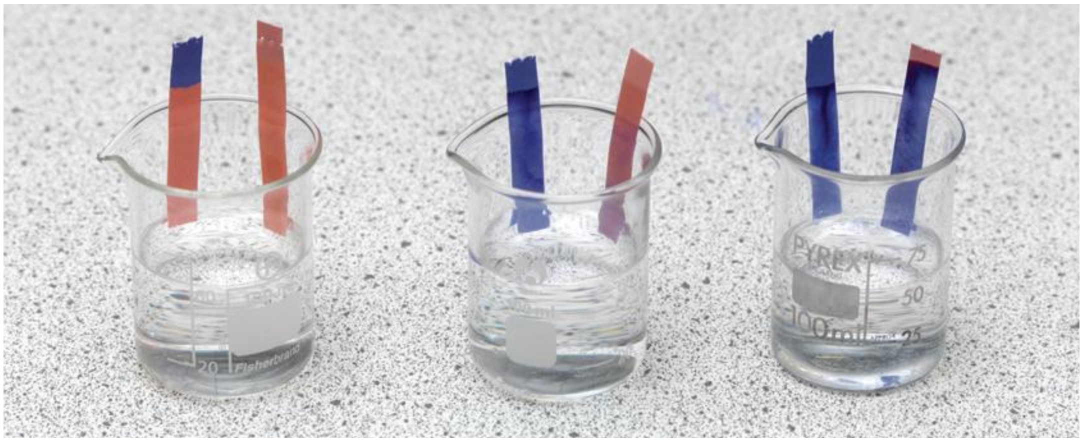


FIGURE 6 Litmus paper is used to determine whether a solution is acidic or basic.

FIGURE 7 A pH probe can measure the strength of a solution very precisely.

pH meters

Although indicators are a useful visual tool to determine the pH of a substance, most chemists use pH meters to more accurately measure the pH of substances. A pH meter measures the voltage of a solution that varies with different hydrogen ion concentrations. pH meters are more precise than indicators because they can measure values to 0.01 of a pH value.

Skill drill

Analysing results from an experiment identify strong and weak acids and bases

Science inquiry skill: Planning investigations (Lesson 1.4) and Communicating scientifically (Lesson 1.9)

Yani is working in the lab with his friend Alec. He notices that Alec has mislabelled four beakers with A, B, C and D, rather than with their chemical names. Yani goes to the front of the lab to get new solutions but there is no more chemical left and he must work with what he has. He performs four tests to identify each solution as a strong or weak acid or base. Yani uses the conductivity test, litmus paper and a reaction with calcium carbonate. He obtains the data shown in Table 2.

TABLE 2 Data from Yani's tests

Sample	Conductivity	Reaction with $CaCO_3$	Blue litmus	Red litmus
A		Bubbles	Red	Red
B	Low		Blue	Blue
C		Occasional bubbles	Red	Red
D	High		Blue	Blue

Practise your skills

1 **Identify** whether the data is qualitative or quantitative. (1 mark)
2 **Identify** the independent variable. (1 mark)
3 **Define** an aim for the experiment. (1 mark)
4 Draw a justified conclusion about the identities of the four solutions based on Yani's data. (4 marks)

Real-world chemistry

Acids and bases in digestion

The stomach is an essential organ in the human digestive system because it is the only place where food is broken down in an acidic environment. It contains hydrochloric acid (HCl), a strong acid, at a concentration of approximately 0.1 M. This acidic environment is essential for the digestion of proteins and absorption of nutrients. Lucky for us, our stomachs are lined with acid-resistant mucus that stops the acid from eating through our stomach lining.

If the acid-rich contents of the stomach are regurgitated, they burn the throat and mouth. This is the main cause of heartburn, which is treated by taking antacids – chemicals that help to neutralise the acid. Sodium bicarbonate ($NaHCO_3$) is an excellent alternative.

The strong acid reacts completely with the weak base $NaHCO_3$ in water. Carbon dioxide gas is lost by exhalation.

$$NaHCO_3(aq) + HCl(aq) \rightleftharpoons NaCl(aq) + CO_2(g) + H_2O(l)$$

Due to the rather bitter taste of $NaHCO_3$, it is also unlikely that a heartburn sufferer would consume too much. It is important to only neutralise the regurgitated acid in the oesophagus and not all of the HCl in the stomach. The stomach acid is necessary for effective digestion of proteins and to kill bacteria and other microorganisms.

Apply your understanding

1 **Calculate** the pH of the stomach fluids. (1 mark)

2 **Describe** the effect that the antacid will have on the pH of the stomach. (2 marks)

FIGURE 8 Antacids neutralise stomach acid.

Check your learning 14.2

Check your learning 14.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Identify** each of the following pH values as that of an acidic, neutral or basic solution.
   - **a** pH 3.1 (1 mark)
   - **b** pH 7 (1 mark)
   - **c** pH 11.3 (1 mark)
   - **d** pH 1.3. (1 mark)

2 **Describe** how an indicator can be used to determine if a substance is an acid or a base. (2 marks)

Analytical processes

3 **Use** Table 3 to **determine** the colour of each indicator at:
   - **a** pH = 2 (1 mark)
   - **b** pH = 10. (1 mark)

TABLE 3 pH indicators for acids and bases

Indicator	pH at which colour changes	Strong acid pH	Strong base pH
Bromophenol blue	4.0	yellow	blue
Methyl red	5.1	red	yellow
Phenolphthalein	9.3	colourless	red

4 **Interpret** Table 4 to **determine** which plants will grow well with low concentrations of hydrogen ions in the soil. (2 marks)

TABLE 4 Vegetable optimum pH levels

Vegetable	Optimal pH
Avocado	6.5–7.5
Broccoli	6.0–7.0
Cabbage	5.6–6.6
Capsicum	6.0–8.0
Cucumber	5.0–6.0
Garlic	5.0–6.0
Mushroom	6.2–6.8
Onions	7.0–8.0
Spinach	5.0–7.0
Tomato	5.5–7.0

Knowledge utilisation

5 If the concentration of hydrogen ions in a solution increases, **predict** what changes there would be in the pH value. (1 mark)

6 The optimum pH for water in a swimming pool is considered to be 7.4. A student claimed that this was a neutral pH. **Evaluate** this claim and comment on whether or not you agree or disagree. (3 marks)

MODULE 14

Lesson 14.3

Review: Acids and bases

Summary

14.1
- Arrhenius acids increase the concentration of hydrogen ions in water and Arrhenius bases increase the concentration of hydroxide ions in water.
- The strength of an acid or a base depends on how easily it ionises in water. Acids and bases that ionise completely are strong, whereas those that undergo partial ionisation are weak.
- The concentration of an acid or base depends on the number of acid or base molecules, respectively, in a specific volume of an aqueous solution.

14.2
- Aqueous solutions with a pH below 7 are acidic. A pH of 7 indicates a neutral aqueous solution, and a pH above 7 indicates a basic aqueous solution.
- Indicators can be used to identify the hydrogen ion concentration of a solution or its pH.

Key formulas

pH	$pH = -\log_{10}[H^+]$ $[H^+] = 10^{-pH}$

Review questions 14.3A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Which of the following is the best definition of an acid?

A An acid is a substance that tastes sour.

B An acid is a substance that produces hydroxide ions.

C An acid is a substance that produces hydrogen ions in water.

D An acid is a substance that reacts with water to produce hydrogen.

2 Equal mole amounts of substance X and substance Y were used to make aqueous solutions of the same volume. The solution made from substance X had a high concentration of hydrogen ions, while the solution containing substance Y had a low concentration of hydrogen ions. Which is the correct statement?

A X has a higher concentration of acid than Y.

B Y has a higher concentration of acid than X.

C Y is a stronger acid than Y.

D Y is an alkali.

3 The best description of an indicator is

A a dye that changes colour.

B a dye that changes colour in a basic solution.

C a dye that changes colour in an acidic solution.

D a dye that changes colour in different hydrogen ion concentrations.

4 A student placed a drop of a liquid on blue litmus paper (which becomes red in acidic solutions) and observed that there was no colour change. This meant the student could assume that the substance was

A a base.

B an acid.

C neutral.

D not an acid.

5 An aqueous solution becomes yellow when universal indicator is added. What does this indicate about the pH of the solution?

A pH = 1

B pH = 5

C pH = 7

D pH = 14

6 A strong base will produce a solution with a pH of

A 1 to 2.

B 4 to 5.

C 7.

D 13 to 14.

7 Which of the following is a weak acid?

A Diluted HCl

B Ammonia

C Carbonic acid

D Potassium hydroxide

8 Which of the following is a strong base?

A Concentrated ammonia

B Water

C Lemon juice

D Diluted sodium hydroxide

9 A 2.00 g sample of HCl is dissolved in 250 mL of water. What is the pH of the resultant solution?

A 3.66

B 0.66

C 0.90

D 2.10

10 5.25 g of barium hydroxide, $Ba(OH)_2$, is dissolved in water to make 375 mL of solution. What is the pH of the resultant solution?

A 13.21

B 10.21

C 12.91

D 0.79

Review questions 14.3B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

11 **Define** the term “ionisation”. (1 mark)

12 **Identify** what all strong acids have in common. (1 mark)

13 **Identify** what all strong bases have in common. (1 mark)

14 **Describe** the Arrhenius model of acids and bases. (2 marks)

15 **Explain** how an aqueous solution of H_2SO_4 can be either concentrated or dilute, but can only be strong (and not weak). (3 marks)

16 **Explain** why an aqueous solution of HCl is able to conduct electricity. (2 marks)

Analytical processes

17 **Contrast** the strong acid HCl and the weak acid HNO_2 and **sketch** a diagram to illustrate the difference. (4 marks)

18 **Compare** a concentrated weak acid and a dilute weak acid. (2 marks)

19 When acids and bases react with each other, it is called a neutralisation reaction. **Consider** why this term is used. **Use** a balanced chemical equation to support your answer. (3 marks)

20 **Determine** the pH you would expect for the following solutions.

a $[H^+] = 10^{-4}$ (1 mark)

b $[H^+] = 10^{-7}$ (1 mark)

c $[H^+] = 10^{-11.}$ (1 mark)

21 A chemist describes a base as being able to completely dissociate (separate into ions) in an aqueous solution. **Consider** what this suggests about the strength of the base. (2 marks)

22 **Reflect on** why more than one indicator should be used when trying to determine if a substance is an acid. (2 marks)

23 **Explain** what type of indicator should be used to determine the pH of a cloudy (opaque) substance. (2 marks)

FIGURE 1 Indicator paper

24 **Infer** whether the addition of water changes the strength or the concentration of an acid. **Use** calculations to support your response. (2 marks)

25 **Apply** the Arrhenius definition of a base to **explain** how $Ca(OH)_2$ qualifies as a base even though it is relatively insoluble in water. (2 marks)

Knowledge utilisation

26 **Predict** if more, less or the same amount of base would be required to neutralise 20 mL of 0.1 M strong acid than would be required to neutralise 20 mL of 0.1 M weak acid. **Justify** your response. (3 marks)

27 **Investigate** why people who suffer from bulimia also experience tooth decay and mouth ulcers. (3 marks)

28 **Discuss** how a difference of 1 pH unit (from pH 3 to pH 4) changes the concentration of hydrogen ions. (2 marks)

29 **Develop** a diagram to represent a:

- **a** dilute aqueous solution of the weak acid, ethanoic acid (CH_3COOH) (2 marks)
- **b** concentrated aqueous solution of the strong base potassium hydroxide (KOH). (2 marks)

Data drill

Analysis of the pH in acid–base reactions

A student performs reactions of two acids with the same base (0.100 M NaOH). They add NaOH to the acid one drop at a time and after each drop, the pH of the solution is measured. Figure 1 shows how pH changes as the volume of NaOH increases for acids A (Figure 1A) and B (Figure 1B). They mark a point on each graph called the equivalence point, which is where the acids are reacting with NaOH in the molar ratio defined by their balanced equations.

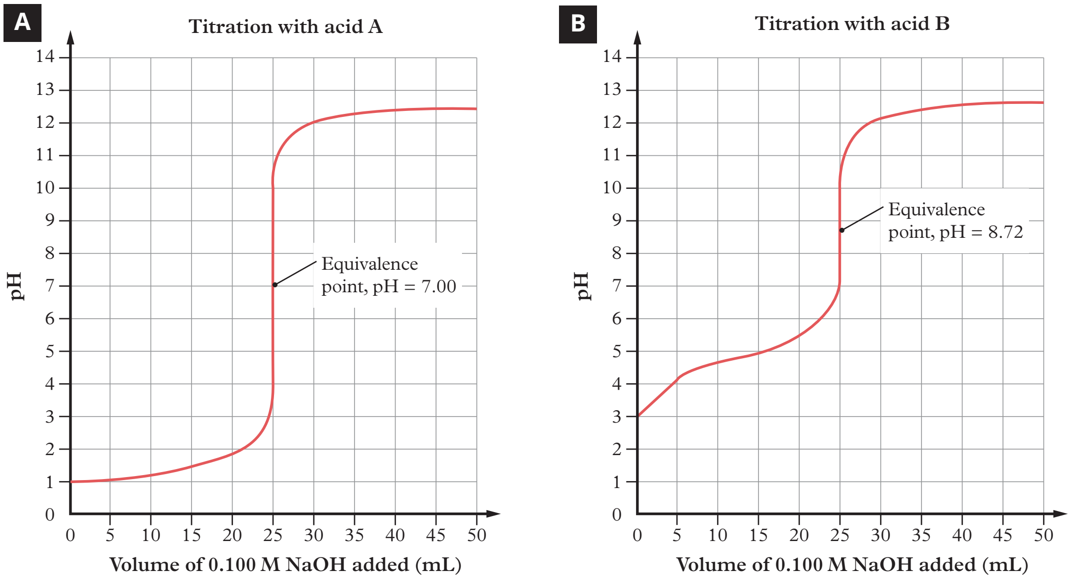


FIGURE 1 Reactions of (A) acid A and (B) acid B with 0.1 M NaOH

Apply understanding

1 **Use** Figure 1 to **calculate** the concentration of hydrogen ions in acids A and B at the equivalence point. (2 marks)

2 **Identify** indicators that would be appropriate to detect the equivalence point for these reactions. (2 marks)

Analyse data

3 **Identify** the trend in pH as equivalence point is approached. (1 mark)

4 **Contrast** the shapes of the two curves in relation to initial pH and the gradients up until when 20 mL of base was added. (2 marks)

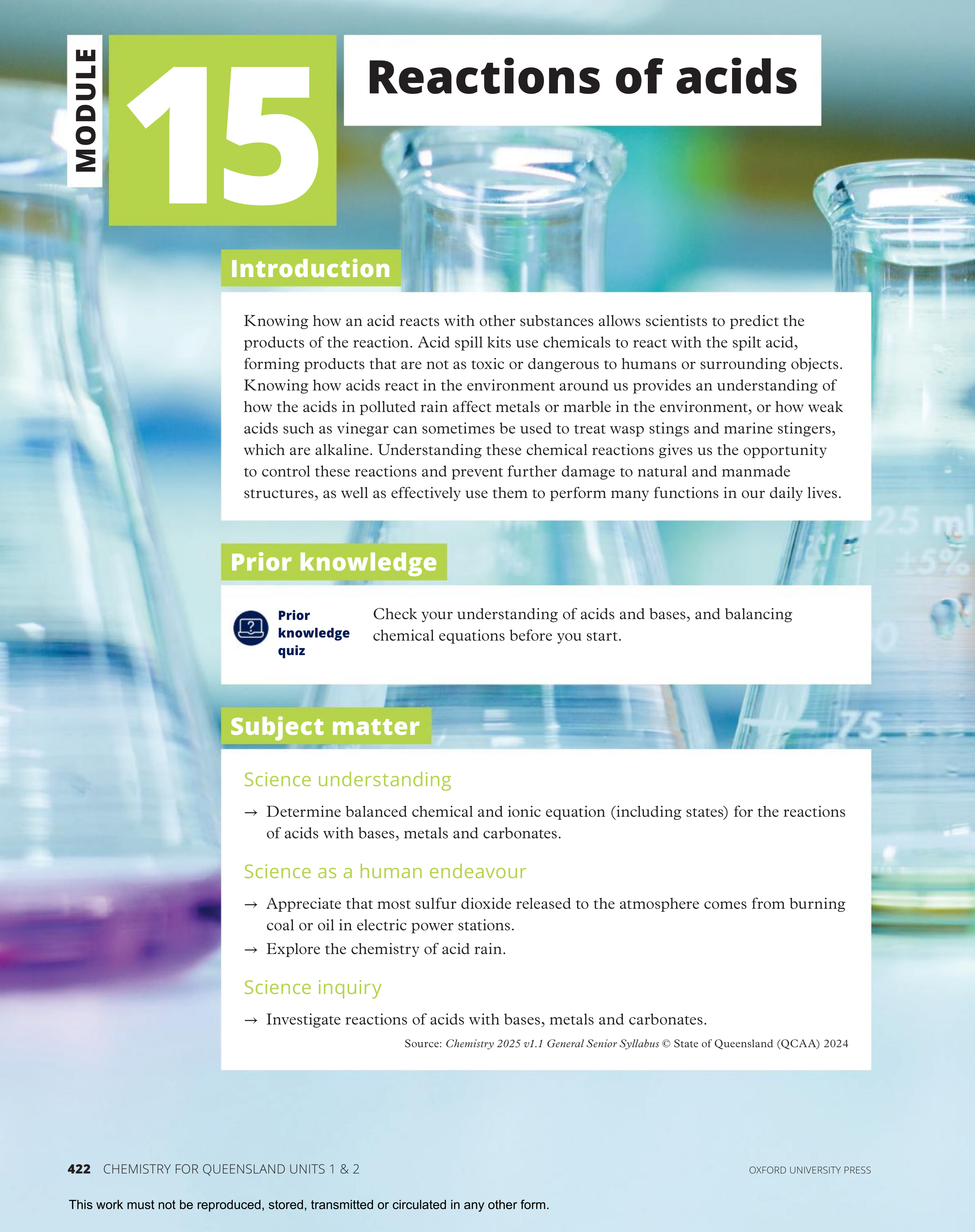
Evaluate evidence

5 **Deduce** whether each acid is strong or weak. (2 marks)

6 **Justify** your answers to question **5**. (2 marks)

Module 14 checklist: Acids and bases

Quizlet: Revise key terms online to test your understanding



MODULE 15

Reactions of acids

Introduction

Knowing how an acid reacts with other substances allows scientists to predict the products of the reaction. Acid spill kits use chemicals to react with the spilt acid, forming products that are not as toxic or dangerous to humans or surrounding objects. Knowing how acids react in the environment around us provides an understanding of how the acids in polluted rain affect metals or marble in the environment, or how weak acids such as vinegar can sometimes be used to treat wasp stings and marine stingers, which are alkaline. Understanding these chemical reactions gives us the opportunity to control these reactions and prevent further damage to natural and manmade structures, as well as effectively use them to perform many functions in our daily lives.

Prior knowledge

Prior knowledge quiz

Check your understanding of acids and bases, and balancing chemical equations before you start.

Subject matter

Science understanding

→ Determine balanced chemical and ionic equation (including states) for the reactions of acids with bases, metals and carbonates.

Science as a human endeavour

→ Appreciate that most sulfur dioxide released to the atmosphere comes from burning coal or oil in electric power stations.

→ Explore the chemistry of acid rain.

Science inquiry

→ Investigate reactions of acids with bases, metals and carbonates.

Source: *Chemistry 2025 v1.1 General Senior Syllabus* © State of Queensland (QCAA) 2024

Practicals

These lessons are available on Oxford Digital.

Lesson 15.3 Investigating reactions of acids

Lesson 15.1

Acid–base neutralisation reactions

Key ideas

- → Neutralisation reactions involve mixing an acid and a base to produce water and a metal salt.
- → If the neutralisation reaction is between a strong acid and a strong base, the final solution will have a pH of 7.
- → The solution resulting from neutralisation of a weak base by a strong acid will have a pH of less than 7. The solution resulting from neutralisation of a weak acid by a strong base will have a pH of greater than 7.
- → A titration is a quantitative analysis technique that can be used to determine the concentration of an acid or base solution. It uses the neutralisation reaction.

Learning intentions and success criteria

What are acid–base neutralisation reactions?

When an acid and a base are mixed together, the hydrogen ion and hydroxide ion react to produce water and a metal salt:

$$\text{acid} + \text{base} \rightarrow \text{water} + \text{metal salt}$$

Figure 1 shows how this is applied to the acid and base examples above. Below the reaction equation is the complete ionic equation, which shows the separate ions.

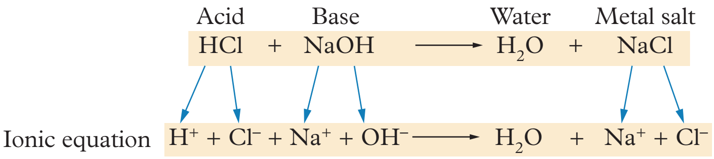


FIGURE 1 In a neutralisation reaction, the hydrogen and hydroxide ions bond to form water.

Study tip

The ability to identify common acids and bases will help you to recognise a neutralisation reaction: acid + base → water + metal salt

This process of an acid reacting with a base is called an acid–base **neutralisation reaction.** In the example above, Na^+ and Cl^- are spectator ions that do not take part in the reaction. The spectator ions combined are called a salt (an ionic compound that dissociates completely in water). If the spectator ions are removed from the equation, neutralisation can be written as an ionic equation of a strong acid–base neutralisation reaction:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

neutralisation reaction
the reaction between an acid and a base to produce water and a metal salt

If a strong acid reacts with a strong base, then the concentration of hydrogen ions and hydroxide ions decreases as they combine to form neutral water molecules. As a result, the pH of the acid solution increases as the base is added, until the moles of acid and base are equal. The pH of the resultant solution is a neutral pH of 7.

When a strong acid reacts with a weak base the product is a weak acid salt that partially ionises to make the pH of the resultant solution slightly less than 7, i.e. slightly acidic.

Likewise, when a strong base reacts with a weak acid in stoichiometric amounts, a weak base salt is produced. It partially ionises in water and causes the pH of the resultant solution to be greater than 7.

Worked example 15.1A

Determining balanced chemical and ionic equations for acid–base reactions

Worked example 15.1A: Watch a video that shows how to solve this problem.

Determine balanced chemical and ionic equations for the reaction between nitric acid and sodium hydroxide. (2 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Determine" means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to recognise the type of reaction occurring and express it using balanced equations. The question is worth 2 marks, so we must recall the theory and apply it to the reaction.
Step 2: Identify the types of reactants involved in the reaction.	Nitric acid (HNO_3) is a strong acid. Sodium hydroxide (NaOH) is a strong base. This is an acid–base neutralisation reaction.
Step 3: Recall the products of the reaction.	acid + base → water + metal salt
Step 4: Write the reactants on the left-hand side of the reaction arrow and the products on the right-hand side of the reaction arrow. This will be the full chemical equation.	$HNO_3(aq) + NaOH(aq) \rightarrow H_2O(l) + NaNO_3(aq)$ (1 mark)
Step 5: Identify the spectator ions. These do not change state in the reaction.	Na^+ and NO_3^- are the spectator ions. They are not included in the final ionic equation.
Step 6: Write a full ionic equation for the reaction after removing the spectator ions.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ (1 mark)

Your turn

Determine balanced chemical and ionic equations for the reaction between hydrochloric acid and potassium hydroxide. (2 marks)

How do you write ionic equations for acid–base reactions?

The ionic reaction equations are a little different when a weak acid or base are involved. Let's consider weak acids as an example. They exist mainly in their molecular form rather than as separate ions. When ethanoic (acetic) acid reacts with sodium hydroxide, the overall reaction could be written as:

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

The weak acid molecule does not ionise to any great extent, so will be written in molecular form in the ionic equation.

$$CH_3COOH(aq) + Na^+(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + Na^+(aq) + H_2O(l)$$

The Na^+ ion is the only spectator ion that can be left out, resulting in the ionic equation:

$$CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O(l)$$

What is a titration?

Neutralisation reactions can be used to determine the concentration of an acid or a base. This makes use of a process called an acid–base **titration**, in which a solution of unknown concentration is slowly neutralised by dripping (or titrating) an acid or a base of known concentration into it. Titrations are commonly used in various ways in industry, such as the monitoring of acids in wine.

The solution of unknown concentration being titrated is called the **analyte**. This analyte is then neutralised by a **titrant** of a known concentration. In a typical acid–base titration, the analyte is the acid and the base is the titrant, but can also involve the reverse, i.e. the base is the analyte being studied.

The **equivalence point** is the point in the titration when the acid and base are present in the exact stoichiometric ratios given by the balanced reaction equation. For example, in the reaction between HCl and NaOH, it would be when there are an equal number of moles of HCl and NaOH, since they have a 1 : 1 ratio in the balanced equation. This is where the reaction reaches completion.

This can be shown visibly by using a suitable indicator, which is added to the analyte at the start of the titration. The point at which the indicator undergoes a permanent physical change in colour is termed the **end point**.

For an accurate analysis, an indicator should be chosen that changes colour very close to the equivalence point. This can be seen using the titration curve in Figure 3. For example, an unknown concentration of acid will become slowly neutralised by slowly adding a base. The concentration of the unknown acid can then be determined by calculating the number of moles of the known base it took for the solution to reach the pH determined by the indicator.

When the titration involves a strong acid being titrated with a strong base, a suitable indicator is bromothymol blue. The change in colour would occur between pH 6.0 and 7.6, corresponding to the approximate pH at the point of neutralisation.

titration
a quantitative analytical technique used to find the unknown concentration of a solution

analyte
the sample being analysed

titrant
the solution being delivered in a titration

equivalence point
the point in a titration at which a reaction is complete; when the acid and base are mixed in exact stoichiometric ratios given by the balanced chemical equation

end point
the point in an acid–base titration at which the indicator changes colour

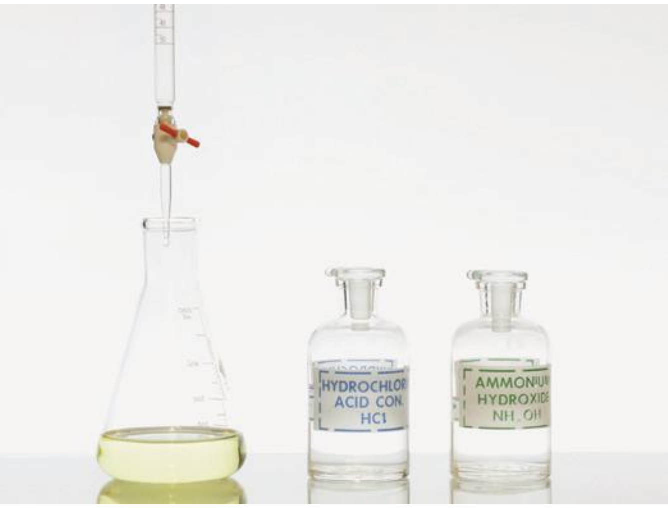


FIGURE 2 Titrating (dripping) a strong base into an acid until it becomes neutralised (as shown by the indicator) is a way of determining the concentration of the acid or base.

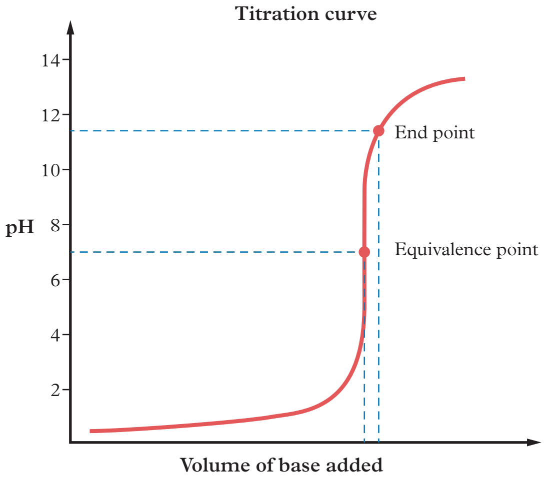
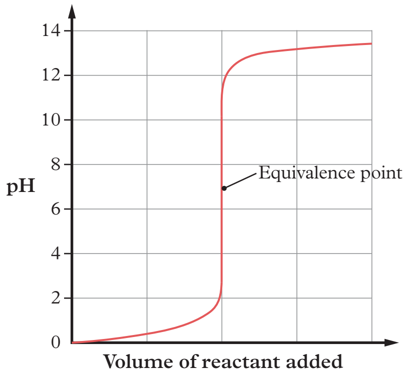
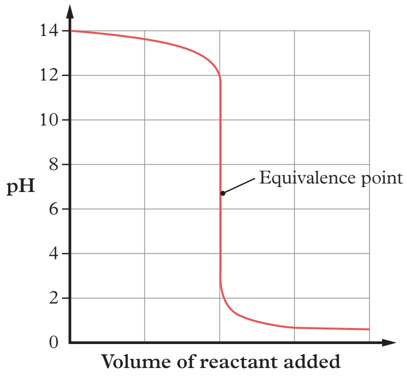
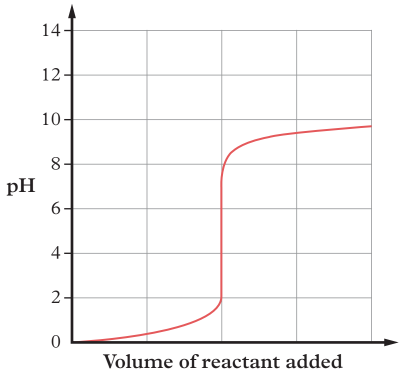
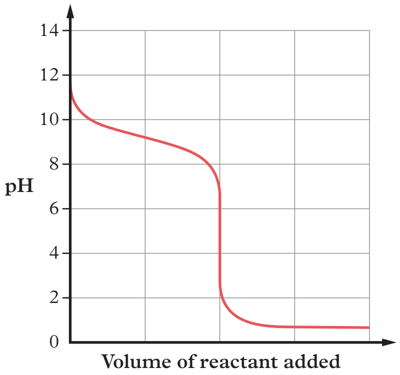
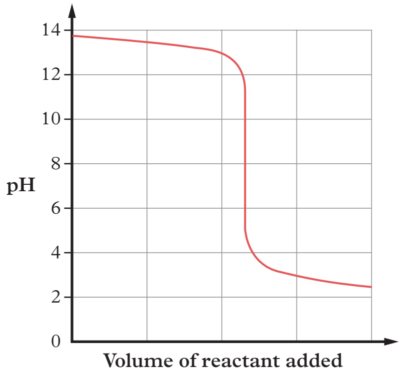
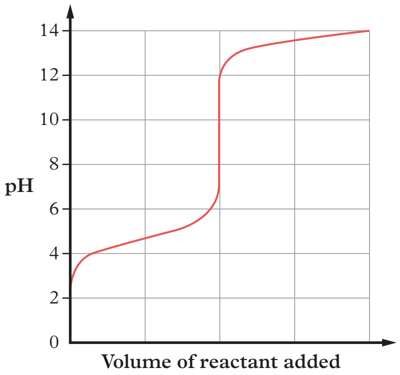


FIGURE 3 A titration curve for a strong acid and a strong base; in this example, the indicator did not change colour close enough to the equivalence point, so the end point is not a precise measure and the calculations of concentration will be inaccurate.

When the titration involves a weak acid being titrated with a strong base, bromothymol blue would have changed colour too early, when the acid was still in excess. Phenolpthalein would therefore be considered a better indicator because it changes colour between pH 8.3 and 10.0.

Titration curves for common titrations are shown in Table 1.

TABLE 1 Common titration curves

Strong acid analyte + strong base titrant	Strong base analyte + strong acid titrant
	
Strong acid analyte + weak base titrant	**Weak base analyte + strong acid titrant**
	
Strong base analyte + weak acid titrant	**Weak acid analyte + strong base titrant**
	

Check your learning 15.1

Check your learning 15.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Define** the terms:
  - **a** reactant (1 mark)
  - **b** product (1 mark)
  - **c** neutralisation reaction (1 mark)
  - **d** neutral pH. (1 mark)

2 **Explain** how a titration is used to determine the concentration of an acid. (3 marks)

3 **Describe** what titrations are used for when investigating acids and bases. (1 mark)

Analytical processes

4 **Determine** the full balanced chemical equation and the ionic equation for the neutralisation reaction between:
  - **a** hydrochloric acid and potassium hydroxide (2 marks)
  - **b** sulfuric acid and sodium hydroxide (2 marks)
  - **c** methanoic acid (CHOOH, a monoprotic weak acid) and barium hydroxide. (2 marks)

5 A neutralisation reaction occurs between barium hydroxide and sulfuric acid.
  - **a** **Determine** the full balanced chemical equation and the ionic equation for the neutralisation reaction between barium hydroxide and sulfuric acid. **Consider** the solubility of the salt formed in this reaction. (3 marks)
  - **b** **Consider** the ionic equation you wrote in part **a**. **Explain** what is happening to the number of ions present, and thus the conductivity of the reaction mixture, as the neutralisation reaction approaches the equivalence point. (4 marks)

6 **Deduce** whether a solution will have a neutral pH of 7 when a strong acid and a weak base are mixed in their stoichiometric ratio so that they react completely. (2 marks)

Knowledge utilisation

7 **Investigate** an example of neutralisation that occurs in your life. **Determine** a full balanced equation and ionic equation for this reaction. (3 marks)

FIGURE 4 An antacid contains a weak base that treats indigestion or acid reflux by neutralising the stomach acid.

Lesson 15.2

Other acid reactions

Key ideas

→ Acids react with most metals to produce a metal salt and hydrogen gas.

→ Aids react with carbonates to produce water, metal salt and carbon dioxide.

→ Hydrogen liberated from a reaction can be detected by a pop test, whereas a lit splint or bubbling the gas through lime water can be used to test for the presence of carbon dioxide.

Learning intentions and success criteria

How do acids react with metals?

Acids react with some metals to produce a metal salt and hydrogen gas:

$$\text{acid} + \text{metal} \rightarrow \text{metal salt} + \text{hydrogen}$$

The type of metal salt produced depends on the metal and acid. For example, hydrochloric acid reacts with magnesium to produce magnesium chloride and hydrogen gas:

$$2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$$

In this reaction, the chloride ion is a spectator ion. Therefore, the ionic equation is:

$$2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + H_2(g)$$

This can be written as a general ionic equation for acids and metals:

$$2H^+(aq) + M(s) \rightarrow M^{2+}(aq) + H_2(g)$$

where M represents the metal.

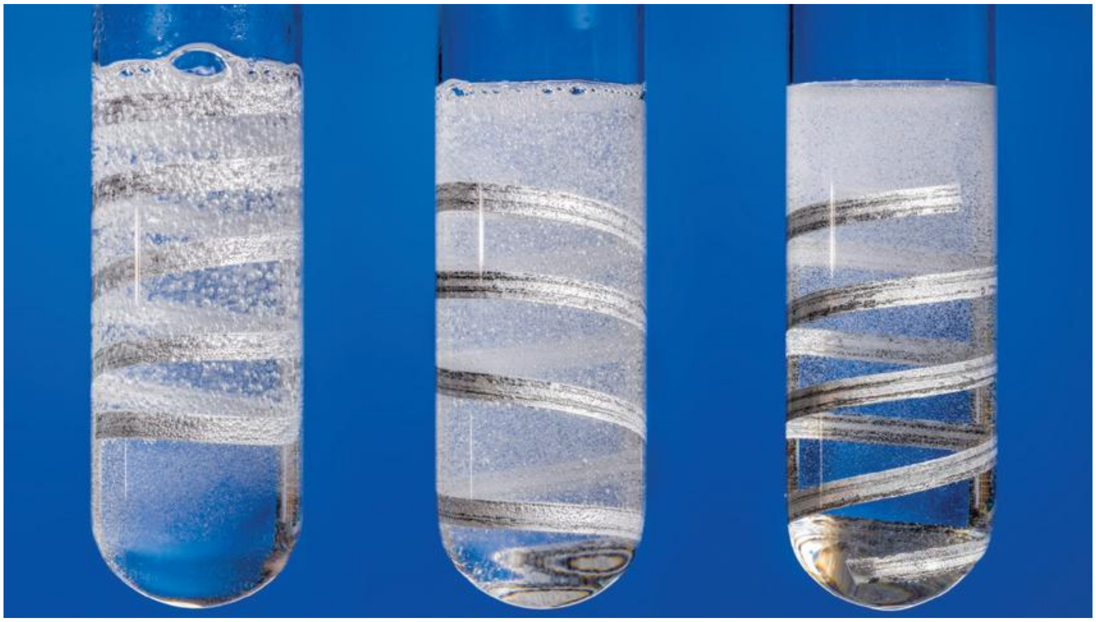


FIGURE 1 Magnesium ribbon reactions with methanoic acid (left), ethanoic acid (middle) and propanoic acid (right) to produce clear hydrogen gas bubbles.

Pop test

pop test
a test to determine the presence of hydrogen gas

The **pop test** is a test to determine if a gas produced is hydrogen (Figure 2), which is reflective of an acid–metal reaction. The hydrogen gas is usually collected in a test tube before being exposed to a lit splint or match. The heat from the flame is enough to initiate a combustion reaction between the collected hydrogen gas and oxygen in the surrounding air, to produce water. The reaction is accompanied by an audible "pop" sound.

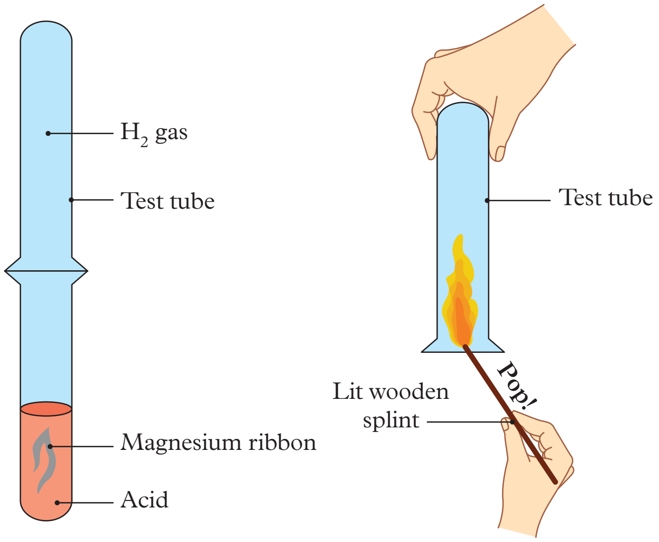


FIGURE 2 The pop test can be used to determine if a gas that is produced in a reaction is hydrogen.

Reactivity of metals

Module 2 discussed the different properties of the elements (or trends) across the periodic table. This is relevant to the way different metals react with acids. Some metals (e.g. magnesium and aluminium) are very reactive with acid (Figure 3). This means the reaction occurs very quickly, with hydrogen gas often seen bubbling rapidly off the metal. Other metals (e.g. lead and tin) are much less reactive with acid. The reaction occurs much slower and can take many weeks to occur to any noticeable extent. There are even a few metals that do not react with acid; for example, copper, mercury and silver.

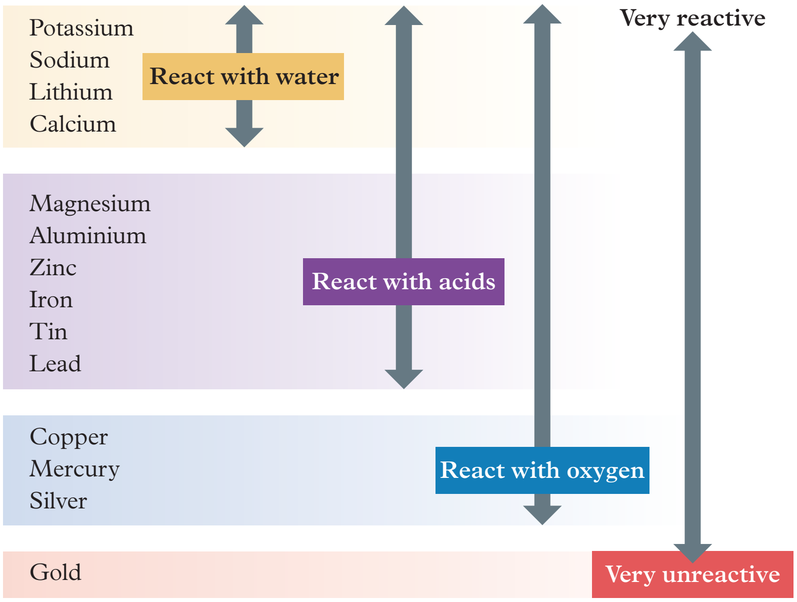


FIGURE 3 Different metals have different levels of reactivity. This is noticeable when the metal is placed in acid.

How do acids react with carbonates?

Acids react with carbonates to produce water, a metal salt and carbon dioxide:

acid + carbonate → water + metal salt + carbon dioxide

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + 2NaCl(aq) + CO_2(g)$$

The chloride ions and sodium ions are spectator ions in this reaction. If they are removed from the equation, the ionic equation becomes:

$$2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$$

Since carbon dioxide gas is formed, detecting its presence is a simple way to determine whether an acid–carbonate reaction has occurred.

lit splint test
a test to determine the presence of carbon dioxide gas

Lit splint test

The **lit splint test** is used to determine if a gas produced is carbon dioxide. The carbon dioxide gas is usually collected in a test tube before being exposed to a lit splint or match (Figure 5). As the flame cannot burn without oxygen being present, when the lit splint is lowered into the carbon dioxide filled test tube, the flame goes out.

FIGURE 4 Sodium carbonate reacts with ethanoic acid to produce visible carbon dioxide gas bubbles.

FIGURE 5 A lit splint requires oxygen to keep burning. Placing a lit splint into a test tube of carbon dioxide makes the flame to go out.

Worked example 15.2A

Determining the products of an acid reaction

Worked example 15.2A: Watch a video that shows how to solve this problem.

Determine the products that form when HCl undergoes a reaction with the following compounds and **determine** balanced chemical and ionic equations for each.

a $CaCO_3(aq)$ (2 marks)

b Zn(s). (2 marks)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	“Determine” means to establish, conclude or ascertain after consideration, observation, investigation or calculation. We need to recognise the types of reaction occurring and express them as balanced chemical equations. Each question is worth 2 marks, so we must recall the theory and apply it to the reaction.
Step 2: Identify the types of reactants involved in the reactions.	HCl is a strong acid. **a** $CaCO_3(aq)$ is a carbonate. This is an acid–carbonate reaction. **b** Zn(s) is a metal. This is an acid–metal reaction.
Step 3: Recall the products of the reactions.	**a** acid + carbonate → water + metal salt + carbon dioxide **b** acid + metal → metal salt + hydrogen
Step 4: Write the reactants on the left-hand side of the reaction arrow and the products on the right-hand side of the reaction arrow. This will be the full chemical equation. Check the coefficients to make sure it is balanced.	**a** $HCl(aq) + CaCO_3(aq) \rightarrow H_2O(l) + CaCl_2(aq) + CO_2(g)$ This is not balanced; we need to add a 2 in front of HCl so that there are 2 × H and 2 × Cl atoms on each side. $2HCl(aq) + CaCO_3(aq) \rightarrow H_2O(l) + CaCl_2(aq) + CO_2(g)$ (1 mark) **b** $HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$ This is not balanced; we need to add a 2 in front of HCl so that there are 2 × H and 2 × Cl atoms on each side. $2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$ (1 mark)
Step 5: Identify the spectator ions. These do not change state in the reaction.	**a** Ca^{2+} and Cl^- are the spectator ions. They are not included in the final ionic equation. **b** Cl^- is the spectator ion. It is not included in the final ionic equation.
Step 6: Write a full ionic equation for the reaction after removing the spectator ions.	**a** $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$ (1 mark) **b** $2H^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + H_2(g)$ (1 mark)

Your turn

Determine the products that form when HNO_3 undergoes a reaction with Na_2CO_3 and **determine** balanced chemical and ionic equations for the reaction. (2 marks)

Challenge

Acid spills

A common compound used to neutralise strong acid spills in the laboratory is sodium carbonate anhydrous or soda ash (Na_2CO_3). When soda ash reacts with sulfuric acid (H_2SO_4) it produces a salt solution. **Explain**, using a chemical equation, how this occurs. (3 marks)

Real-world chemistry

Acid rain

Humans have relied on burning coal and oil for energy for many centuries. Coal often contains sulfur, and combustion of coal can produce sulfur dioxide gas (SO_2). Over time, the gas breaks down in the atmosphere, as it does with the SO_2 gas produced in volcanic eruptions.

The number of combustion reactions has increased because of industrialisation. This means the amount of SO_2 in the atmosphere has increased, allowing the gas time to react with the water in the atmosphere and form H_2SO_3. This then reacts with oxygen (O_2) to produce sulfuric acid (H_2SO_4):

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(g)$$

$$2H_2SO_3(g) + 7O_2(g) \rightarrow 2H_2SO_4(g)$$

As the gas reacts with the water in the atmosphere, it starts to dissolve, and the reaction falls to the ground as acid rain. Acid rain reacts with exposed metal or marble ($CaCO_3$), as seen in the following reactions:

$$H_2SO_4(l) + Fe(s) \rightarrow H_2(g) + FeSO_4(aq)$$

$$H_2SO_4(l) + CaCO_3(s) \rightarrow CO_2(g) + H_2O(l) + CaSO_4(aq)$$

The iron(II) sulfate and calcium sulfate are soluble in water so the metal and marble objects erode over time.

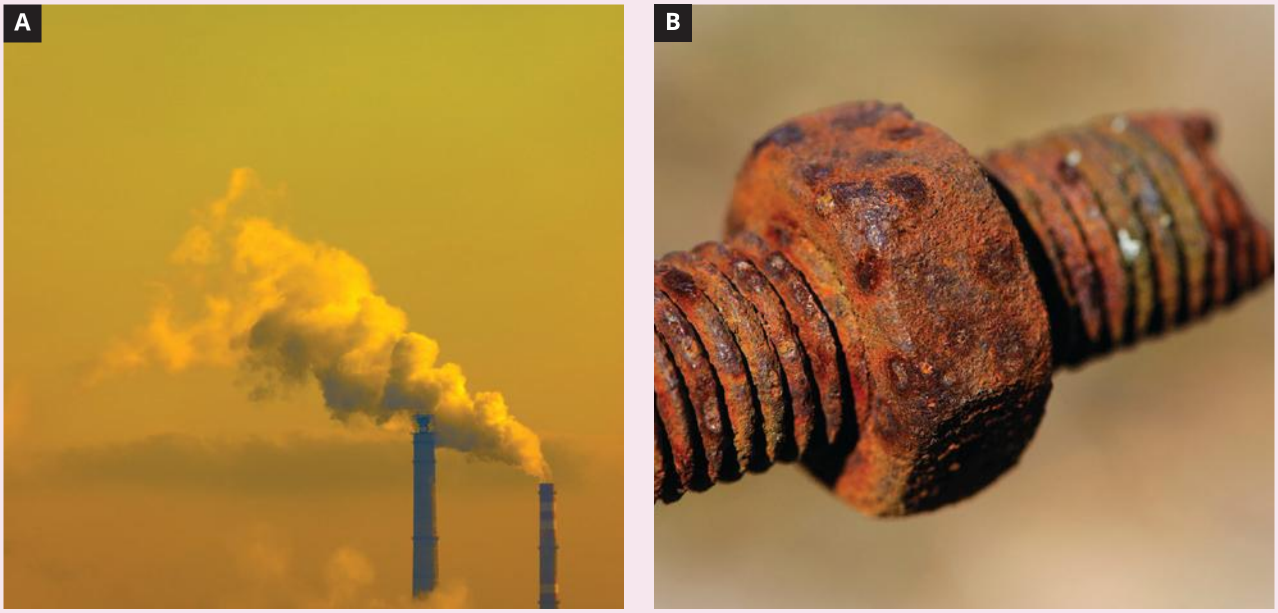


FIGURE 6 The reaction between industrially produced sulfur dioxide (SO_2) and water in the atmosphere produces acid rain. The acid reacts with iron bolts, causing them to break down.

Apply your understanding

1 **Determine** ionic equations for the reaction between sulfuric acid and iron, and sulfuric acid and calcium carbonate. (2 marks)

2 **Determine** balanced chemical and ionic equations that illustrate how nitric acid can also form from a reaction between nitrogen dioxide (a man-made fossil pollutant frequently produced in high temperature combustion) and water. (2 marks)

Skill drill

Predicting the results of acid reactions

Science inquiry skill(s): Planning investigations (Lesson 1.4)

Small samples of various solid compounds were reacted with 1 M hydrochloric acid, as shown in Figure 7.

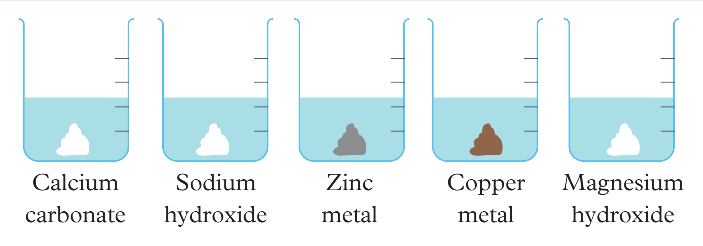


FIGURE 7 Samples of solid compounds or elements reacting with hydrochloric acid

Practise your skills

1 **Identify** the beakers in which you would expect a reaction to occur. **Describe** what you would observe. (4 marks)

2 **Determine** balanced chemical equations for the beakers in which you would expect to observe a reaction. (4 marks)

3 Some of the reactions in Figure 7 will not look like they are occurring. **Propose** a way in which you could observe the reaction occurring. (3 marks)

Check your learning 15.2

Check your learning 15.2: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Recall** the general equation (showing predicted products) for the reaction of:

- **a** acid + base (1 mark)
- **b** acid + metal (1 mark)
- **c** acid + carbonate. (1 mark)

2 **Recall** the general ionic equation for each of the general equations shown in question **1**. (3 marks)

Analytical processes

3 **Determine** the products of the following reactions and **determine** balanced chemical and ionic equations for each:

- **a** H_2SO_4(aq) + $Mg(OH)_2$(s) (2 marks)
- **b** H_2SO_4(aq) + Mg(s) (2 marks)
- **c** H_2SO_4(aq) + $MgCO_3$(s) (2 marks)

4 **Contrast** the pop test and the lit splint test. (2 marks)

Knowledge utilisation

5 The shells of many sea organisms consist of calcium carbonate. **Investigate** the impact of the increasing acidity of the oceans on the survival of these animals. Provide chemical reactions to support your answer. (3 marks)

Learning intentions and success criteria

Video demonstration

Practical

Lesson 15.3

Investigating reactions of acids

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

MODULE 15

Lesson 15.4

Review: Reactions of acids

Summary

15.1
- Neutralisation reactions involve mixing an acid and a base to produce water and a metal salt.
- If the neutralisation reaction is between a strong acid and a strong base, the final solution will have a pH of 7.
- The solution resulting from neutralisation of a weak base by a strong acid will have a pH of less than 7. The solution resulting from neutralisation of a weak acid by a strong base will have a pH of greater than 7.
- A titration is a quantitative analysis technique that can be used to determine the concentration of an acid or base solution. It uses the neutralisation reaction.

15.2
- Acids react with most metals to produce a metal salt and hydrogen gas.
- Different metals react with acid at different rates.
- A pop test can be used to determine if a gas is hydrogen.
- Acids react with carbonates to produce water, metal salt and carbon dioxide.
- Hydrogen liberated from a reaction can be detected by a pop test, whereas a lit splint or bubbling the gas through lime water can be used to test for the presence of carbon dioxide.

15.3
- Practical: Investigating reactions of acids

Review questions 15.4A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Consider the reaction:

$$CH_3CH_2NH_3^+ + NaOH \rightarrow Na^+ + CH_3CH_2NH_2 + H_2O$$

Which of the substances is a strong base?

A $CH_3CH_2NH_3^+$

B NaOH

C $CH_3CH_2NH_2$

D H_2O

2 The end point in an acid–base titration occurs when:

A the pH is 7.

B neutralisation is complete.

C the indicator changes colour.

D stoichiometric amounts of the acid and base are combined.

3 Antacids are used to partially neutralise the acid produced by cells in the stomach. Antacids contain

A an acid.

B a base.

C a metal.

D water.

4 In the reaction between sodium carbonate and hydrochloric acid, what products will be produced?

A Sodium chloride and water

B Only water and carbon dioxide

C Sodium chloride, water and hydrogen

D Sodium chloride, water and carbon dioxide

5 In the equation:

$$NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$$

the chloride ion is the

A acid.

B base.

C spectator ion.

D metal salt.

6 A student compared the rate of bubble formation when hydrochloric acid was reacted with Mg, Zn, Cu and Pb. Which of the following observations would have been made?

- **A** The bubbles formed with Cu would have a reddish tint from the colour of the metal.
- **B** The piece of Zn would have been used up before any of the other metals.
- **C** Cu and Pb would have both formed bubbles very slowly.
- **D** Mg would have produced bubbles at a rapid rate.

7 During the neutralisation of an acid by a base, the pH

- **A** increases rapidly until it is close to neutral, and then increases more slowly.
- **B** increases slowly until it is close to neutral, and then increases more rapidly.
- **C** decreases rapidly until it is close to neutral, and then decreases more slowly.
- **D** decreases slowly until it is close to neutral, and then decreases more rapidly.

8 In regions with hard water, limescale or calcium carbonate frequently builds up in bathrooms. A suitable agent to clean up the limescale would be

- **A** a strong base such as dishwater detergent.
- **B** sodium hydrogen carbonate, also called bicarbonate soda.
- **C** sodium chloride in table salt.
- **D** vinegar, containing ethanoic acid.

FIGURE 1 Limescale on a tap

9 Many geologists carry a small dropper bottle of vinegar, which contains the weak acid, acetic acid. When used to test rocks, geologists would most likely be looking for

- **A** neutralisation, to see if the rock was basic.
- **B** bubbles, to see if there are metals in the rock.
- **C** bubbles, to see if there are carbonate minerals in the rock.
- **D** the rock to react to find valuable ore bodies underneath.

Review questions 15.4B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

10 **Define** a neutralisation reaction by providing the reactants and products. (1 mark)

11 **Identify** the products from the reaction of an acid with a metal. (1 mark)

12 **Identify** the products from the reaction of an acid with a carbonate. (1 mark)

13 Fertilisers often contain metal salts. For each of the following metal salts, **identify** an acid and base that could be used to produce it.

- **a** ammonium nitrate (2 marks)
- **b** potassium phosphate (2 marks)

FIGURE 2 Chemical fertiliser

Analytical processes

14 **Determine** the full balanced equation and the ionic equation for the reaction between sodium hydroxide and sulfuric acid. (2 marks)

15 **Determine** the products of the reaction between reactant 1 and reactant 2 and complete the blank cells in Table 1. (9 marks)

TABLE 1 Products and reactions

Reactant 1	Reactant 2, hydrochloric acid	Reactant 2, sulfuric acid	Reactant 2, nitric acid
Sodium carbonate			
Iron(II) carbonate			
Copper(II) carbonate			

16 **Determine** the full balanced equation and the ionic equations for the following reactions. **Consider** any insoluble salts that are formed.

- **a** Calcium hydroxide + sulfuric acid (2 marks)
- **b** Lithium hydroxide + hydrochloric acid (2 marks)
- **c** Calcium carbonate + nitric acid (2 marks)
- **d** Potassium hydroxide + sulfuric acid (2 marks)
- **e** Calcium carbonate + nitric acid (2 marks)
- **f** Lead(II) + hydroiodic acid (2 marks)
- **g** Zinc + phosphoric acid. (2 marks)

Knowledge utilisation

17 The pH of a swimming pool should be maintained at 7.5. The water in a pool is found to have a pH of 5, so the pH needs to be raised to an appropriate pH. **Propose** a safe solution to this problem. **Justify** your response. (2 marks)

18 Bee stings and ant bites often contain acidic molecules. **Determine** what you might use to neutralise these molecules remaining on the skin surface. **Justify** your response. (2 marks)

19 A lit splint is used to determine if a gas is either hydrogen or carbon dioxide. **Justify** why you would not use a lit splint to test if a gas is oxygen. (2 marks)

20 Plants need to absorb a variety of plant nutrients to survive and grow.

- **a** **Evaluate** the data in Figure 3 to **identify** what nutrients a plant would be able to absorb in a soil of pH 4 to 5. (1 mark)
- **b** **Consider** how a plant's ability to absorb nutrients would be affected if lime (a base) was added to the soil in part **a** so that the pH increased to 8. (2 marks)

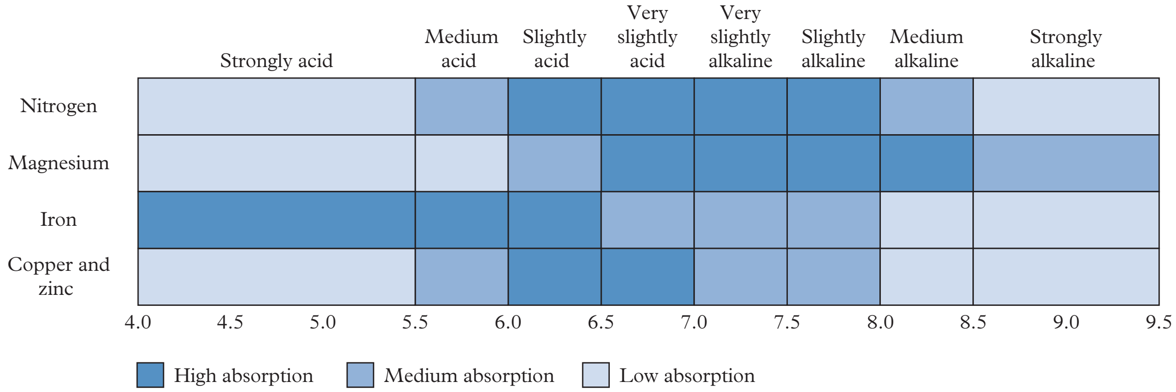


FIGURE 3 Levels of absorption of different nutrients by plants at different pH values

Data drill

Titration curves

A titration was conducted for two monoprotic acids, Acid 1 and Acid 2, using LiOH as the titrant. Monoprotic means that they can only donate one hydrogen ion each. The titration curve is shown in Figure 1, which plots the volume of LiOH(aq) added in mL against the pH of the solution.

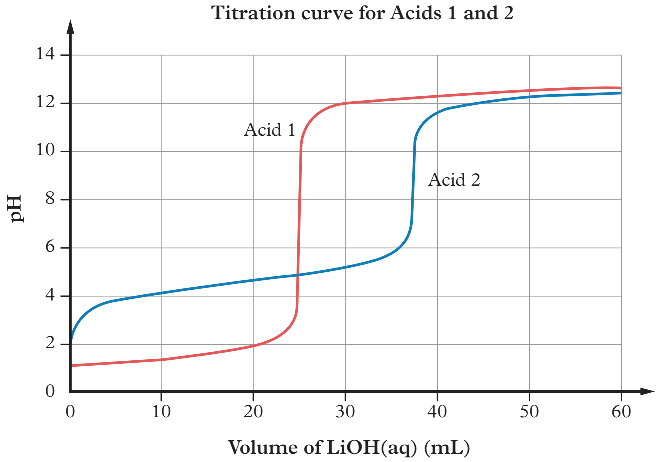


FIGURE 1 Titration curves for Acid 1 and Acid 2

Apply understanding

1 **Identify** the products that can be formed by the reaction of Acids 1 and 2 and LiOH. (2 marks)

2 **Determine** whether phenol red is an appropriate indicator for the titrations. (2 marks)

Analyse data

3 **Compare** the relative strengths of Acid 1 and Acid 2. (2 marks)

Evaluate evidence

4 **Deduce** whether it is possible to tell which acid has a greater concentration. (3 marks)

5 **Predict** what the titration curve for Acid 1 may look like if a weak base was used instead of LiOH. (1 mark)

Module 15 checklist: Reactions of acids

Quizlet: Revise key terms online to test your understanding

UNIT 2 Topic 2 review

Multiple choice

(1 mark each)

1 A student dissolves 20.0 g of potassium dichromate in 150 mL of water. What is the solute, solvent and solution, respectively?

- **A** Potassium dichromate, aqueous potassium dichromate, water
- **B** Aqueous potassium dichromate, potassium dichromate, water
- **C** Potassium dichromate, water, aqueous potassium dichromate
- **D** Water, potassium dichromate, aqueous potassium dichromate

2 A solution is made by dissolving 15.8 g of NaCl in 125 mL of water. What is the concentration of the resultant solution?

- **A** 0.12 M
- **B** 2.16 M
- **C** 120 M
- **D** 2.16×10^{-3} M

3 A chemist needs to make a solution of 1.5 M iron(III) chloride but only has 7.50 g of the solid salt. What volume of water should be used to make the maximum volume of solution?

- **A** 5.00 mL
- **B** 11.25 mL
- **C** 30.83 mL
- **D** 69.36 mL

4 In a precipitation reaction, iron(III) bromide reacts with ammonium phosphate. The correct overall equation to represent this reaction is

- **A** $FeBr_2(aq) + (NH_4)_3PO_4(aq) \rightarrow Fe_3(PO_4)_2(aq) + NH_4Br(s)$
- **B** $3FeBr_2(aq) + 2(NH_4)_3PO_4(aq) \rightarrow Fe_3(PO_4)_2(s) + 6NH_4Br(aq)$
- **C** $FeBr_3(aq) + (NH_4)_3PO_4(aq) \rightarrow FePO_4(s) + 3NH_4Br(aq)$
- **D** $FeBr_3(aq) + (NH_4)_3PO_4(aq) \rightarrow FePO_4(aq) + NH_4Br(s)$

5 A student writes the precipitation reaction between aqueous solutions of cobalt(II) hydroxide and potassium phosphate but forgets to add states and balance it.

$$Co(OH)_2 + K_3PO_4 \rightarrow KOH + Co_3(PO_4)_2$$

Which of the following represents the ionic equation of the precipitation reaction?

- **A** $OH^-(aq) + K^+(aq) \rightarrow KOH(s)$
- **B** $3Co^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Co_3(PO_4)_2(s)$
- **C** $3Co(OH)_2(aq) + 2K_3PO_4(aq) \rightarrow 6KOH(s)$
- **D** $3Co(OH)_2(aq) + 2K_3PO_4(aq) \rightarrow Co_3(PO_4)_2(s)$

6 The products of a reaction between sodium carbonate and sulfuric acid are

- **A** $NaSO_4$ and CO_2
- **B** Na_2SO_4 and CO_2
- **C** $NaSO_4$, CO_2 and H_2O
- **D** Na_2SO_4, CO_2 and H_2O.

Use the following information to answer questions 7 to 9.

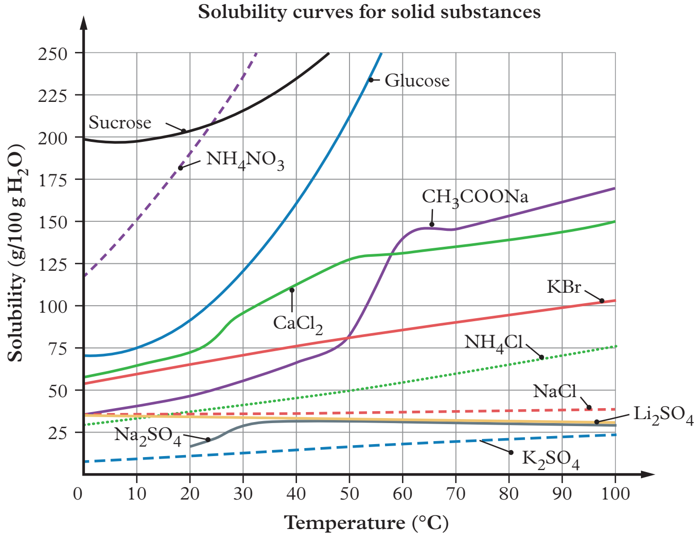


FIGURE 1 Solubility curves for solid substances

7 Sucrose is very soluble in water because

- **A** it has more whole positive and negative charges, which interact with the dipoles in water molecules.
- **B** it has more partial positive and negative charges, resulting in more hydrogen bonding with water molecules.
- **C** it has more partial positive and negative charges, resulting in more dipole–dipole attractions with water molecules.
- **D** it has less whole positive and negative charges, which increases its ability to interact with one of water's dipoles at a time.

8 The concentration of a saturated solution of glucose, $C_6H_{12}O_6$, in water at 20°C is closest to

- **A** 0.5 M.
- **B** 5.0 M.
- **C** 15.0 M.
- **D** 0.15 M.

9 The concentration of a saturated solution of CH_3CO_2Na in water at 56°C is closest to

- **A** 1.50 M.
- **B** 3.00 M.
- **C** 15.0 M.
- **D** 30.0 M.

10 A 10.00 mL sample of a 0.50 M HCl solution was diluted to 150.00 mL. What is the pH of the resultant solution?

- **A** 0.88
- **B** 1.30
- **C** 1.48
- **D** 2.30

11 A 3.80 g sample of calcium hydroxide is dissolved in 250.0 mL of water. What is the pH of the resultant solution?

- **A** 0.38
- **B** 0.69
- **C** 13.31
- **D** 13.61

12 Which of the following is true about the pH of weak acids?

- **A** Weak acids will partially ionise in water, decreasing $[H^+]$ and increasing pH.
- **B** Weak acids will completely ionise in water, increasing $[H^+]$ and decreasing pH.
- **C** Weak acids will partially ionise in water, increasing $[H^+]$ and decreasing pH.
- **D** Weak acids will completely ionise in water, decreasing $[H^+]$ and increasing pH.

13 Which of the following is true?

- **A** Acid and metal forms a salt.
- **B** Acid and water has a high pH.
- **C** Acid and base forms a salt, water and hydrogen.
- **D** Acid and metal carbonate forms a salt, carbon dioxide and water.

14 In a reaction between 0.054 g of magnesium and 20.00 mL of 0.500 M hydrochloric acid at SLC, 50.00 mL of hydrogen gas is formed. What is the mass of magnesium remaining?

- **A** 0.0004 g
- **B** 0.0268 g
- **C** 0.0272 g
- **D** 0.0535 g

15 A 2.00 g sample of sodium carbonate is reacted with 50.00 mL of 2.00 M sulfuric acid. A gas is produced and it is collected at 105 kPa and 18°C. The volume of the gas is

- **A** 16.68 mL.
- **B** 26.88 mL.
- **C** 71.23 mL.
- **D** 434.6 mL.

Short response

16 The solubility curves for various substances are shown.

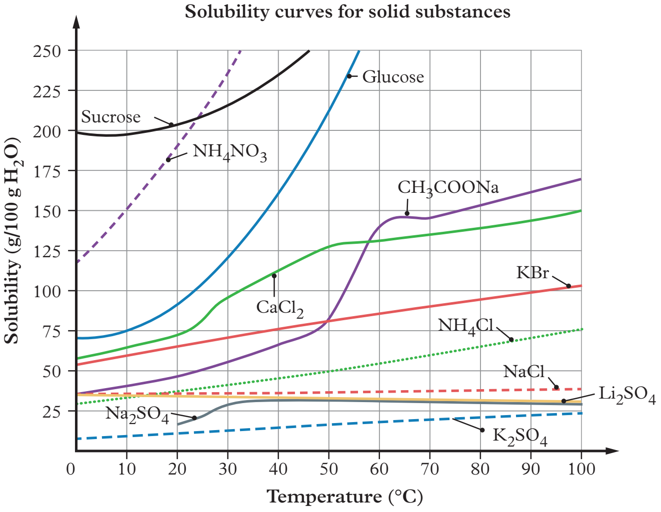


FIGURE 2 Solubility curves for solid substances

Use the solubility curve to **explain** how to make a supersaturated solution of NaCl in 150 mL of water at 70°C. In your response, use the terms "solute", "solvent" and "solution". (3 marks)

17 Acids and bases can be described based on their strength.

a **Use** chemical equations to **explain** the difference in the concentration of H^+ ions in a strong and a weak acid. (3 marks)

b **Contrast** a dilute base and a weak base. (2 marks)

18 **Explain** why calcium chloride has a melting point of 772°C but will readily dissolve in water. (4 marks)

19 **Explain** whether ammonia (NH_3) or ammonium chloride (NH_4Cl) is more soluble in water. (3 marks)

20 The following solubility curve represents the solubility of gases in water.

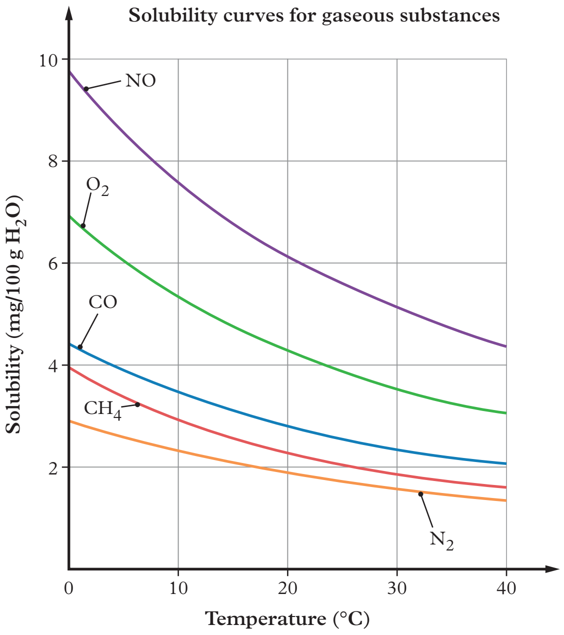


FIGURE 3 Solubility curves for gases

a **Justify** the difference between the solubility of NO and O_2. (3 marks)

b **Justify** the difference between the solubility of CH_4 and N_2. (3 marks)

21 The solubility curve represents the solubility of ionic and covalent substances in water.

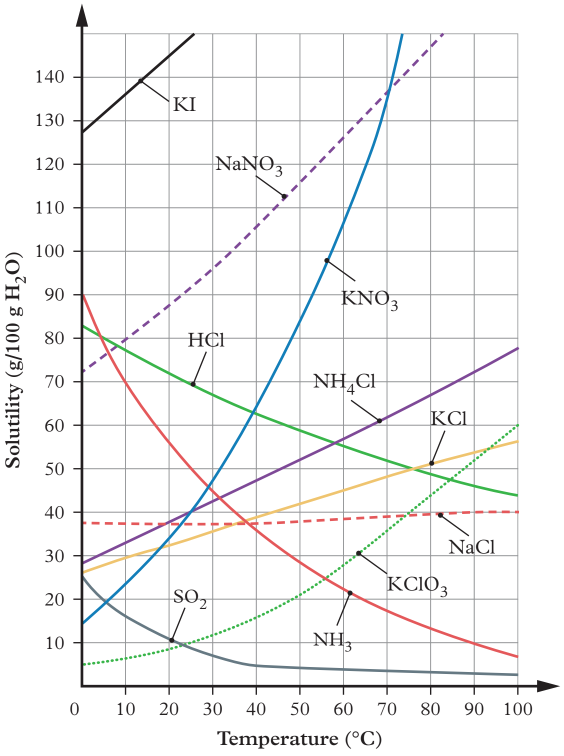


FIGURE 4 Solubility curves for ionic and covalent substances

- **a** **Calculate** the mass of $KClO_3$ required to make a 250.00 mL saturated solution at 60°C. (2 marks)
- **b** **Determine** the temperature required to make a saturated 20.00 mL solution of KNO_3 using 15 g of the salt. (2 marks)

22 **Determine** balanced precipitation reaction equations and ionic equations for the following

- **a** lithium phosphate reacts with zinc iodide (2 marks)
- **b** manganese(II) chloride and sodium hydroxide. (2 marks)

23 A chemist has four unknown solutions, labelled A, B, C and D. They know that the solutions contain either potassium chloride, magnesium carbonate, sodium hydroxide or iron(II) nitrate.

- **a** **Describe** a method to **determine** how each solution could be identified and **describe** the expected results for each test. (4 marks)
- **b** **Evaluate** the ability of each test to identify the solute with absolute certainty. **Describe** any limitations. (2 marks)

24 Calculate the pH of

- **a** a solution of 0.00500 g of H_2SO_4 in 220.00 mL of water (5 marks)
- **b** a solution of 5.6×10^{-5} g of NaOH in 300 mL of water. (5 marks)

25 **Determine** the balanced chemical equations, including states, for reactions between the following aqueous solutions:

- **a** $MgCO_3 + H_2SO_4$ (1 mark)
- **b** $HNO_3 + Ca(OH)_2$ (1 mark)
- **c** Cu + HCl. (2 marks)

26 **Calculate** the molar concentration of a solution made by dissolving 6.52 g of potassium dichromate in 150.00 mL of water. (2 marks)

27 A 5.00 g sample of NaCl is added to 25.00 mL of water and then diluted by adding 60.00 mL of water. **Calculate** the molar concentration of the diluted solution. (6 marks)

28 **Determine** the volume of gas generated by a reaction between 3.600 g of barium carbonate and 10.00 mL of 0.5 M nitric acid (HNO_3) at 21°C and 104 kPa. (7 marks)

29 **Consider** what a solubility curve for a strong and weak base may look like when plotted on the same graph.

- **a** **Describe** the difference in the shape of the graphs. (2 marks)
- **b** **Deduce** which base is more soluble and **explain** why. (2 marks)
- **c** **Describe** the effect of temperature on the solubility of both bases. (2 marks)

30 A scientist wishes to neutralise a 50.00 mL solution of 1.00 M H_3PO_4 using 20.00 mL of KOH solution.

- **a** **Determine** the mass of KOH that must be dissolved in the 20.00 mL for the H_3PO_4 to be completely neutralised. (5 marks)
- **b** **Explain** why it is essential that no reactant is in excess. (2 marks)
- **c** **Predict** the effect on the pH of the solution if too much KOH was added. (1 mark)
- **d** **Explain** how you could test the final solution to **determine** whether a reactant was in excess. (2 marks)

TOTAL MARKS

/95

MODULE 16 Rates of chemical reactions

Introduction

The efficient production of chemicals on an industrial scale depends on the rates of chemical reactions. These determine the amount of product that can be formed in a given time period. A fast reaction ensures that products are formed within a relatively short period of time.

The food industry also relies heavily on knowledge of rates of reaction. Fruit is stored for long periods of time (some apples are stored for up to two years) before it is sold in supermarkets and consumed. They must be stored in a way that slows the breakdown of biomolecules that lead to bruising and rotting.

Rates of reaction can be explained by kinetic molecular theory, which states that particles are in constant random motion. For a reaction to occur, particles must make contact with each other and have sufficient energy to react. This is the foundation that led to the development of collision theory. This theory was first proposed by German chemist Max Trautz in 1916. He also pioneered the understanding of concepts of activation energy and chemical energy.

British chemist William Lewis proposed a very similar theory two years later. Because of the breakdown in communications during World War I, both men were credited with this discovery.

This module investigates rates of reactions, the factors that affect reaction rates, the use of catalysts and the representation of chemical reactions in energy profile diagrams.

Prior knowledge

Prior knowledge quiz

Check your understanding of chemical change and catalysts before you start.

Subject matter

Science understanding

→ Explain how temperature, surface area, pressure (gaseous systems), concentration and the presence of a catalyst, can affect the rate of the reaction.

→ Apply the collision theory to determine the effect of concentration, temperature, pressure and surface area on the rate of chemical reactions.

- → Sketch Maxwell–Boltzmann distribution curves for reactions with and without catalysts.
- → Describe activation energy (E_a).
- → Explain the relationship between the strength and number of the existing chemical bonds, the magnitude of the activation energy and the rate of a chemical reaction.
- → Sketch energy profile diagrams for reactions with and without catalysts.
- → Analyse energy profile diagrams for reactions with and without catalysts determine the enthalpy change and activation energy.
- → Explain how catalysts affect the rate of a chemical reaction.
- → Calculate the rate of chemical reactions by measuring the rate of formation of products or the depletion of reactants. (Formula: rate of reaction $= \frac{\text{increase in product concentration } (\Delta[P])}{\text{time taken}}$ or $\frac{\text{decrease in reactant concentration } (-\Delta[R])}{\text{time taken}}$)
- → Analyse data and graphical representations of relative changes in the concentration, volume and mass against time to determine rate of reaction. (Order of reaction is not required.)

Science as a human endeavour

- → Appreciate that catalysts work in a variety of ways, and knowledge of the structure of the enzyme molecules helps scientists to explain and predict how they are able to lower the activation energy for reactions.
- → Appreciate that collision theory enables chemists to explain and predict the rates of a vast range of chemical reactions in many different contexts.

Science inquiry

- → Investigate factors that affect the rate of chemical reactions.

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Practicals

oxforddigital **These lessons are available on Oxford Digital.**

Lesson 16.2 Investigating the rate of chemical reactions

Lesson 16.1

Factors affecting reaction rates

Key ideas

→ Collision theory states that reactants must collide in the correct orientation and with sufficient energy to result in a successful chemical reaction.

→ Temperature, pressure, concentration and surface area influence the rate of a chemical reaction.

Why is it important to understand reaction rates?

Scientists manipulate chemical reactions by knowing how to speed up or slow down reactions, which is essential in many different areas. Industrial chemists need to speed up chemical reactions to produce their products in the most cost-effective way. Food chemists want to slow down the rate of decay of nutrients and molecules within products to extend shelf life within a supermarket. For this reason, an understanding of **reaction rates** is essential.

Learning intentions and success criteria

reaction rate the change in concentration of a reactant or product per unit of time

collision theory the theory that states that reactants must collide with the correct orientation and sufficient energy for a reaction to occur

What is collision theory?

Collision theory states that reactants must collide in order to react. However, not all collisions result in a reaction. In fact, only a certain percentage will collide successfully to form products. The success of a collision depends on its orientation and energy.

- Reactants must collide with the correct orientation, or in the correct position. Figure 1 demonstrates molecules of ethene (C_2H_4) and hydrogen chloride (HCl) reacting according to the chemical equation:

$$C_2H_4(g) + HCl(g) \rightarrow C_2H_5Cl(g)$$

Unless the molecules collide in the correct orientation, the collision is unsuccessful.

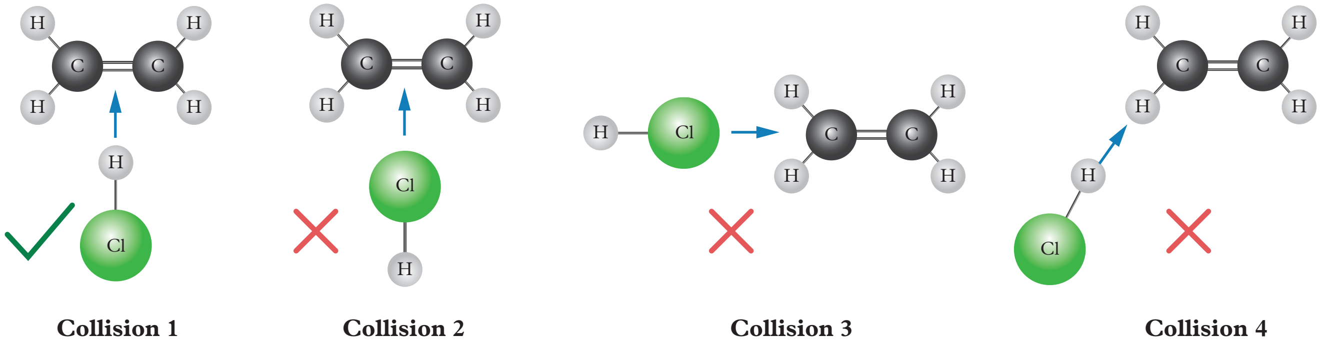


FIGURE 1 The correct and incorrect orientations of ethene (C_2H_4) and hydrogen chloride (HCl) molecules that result in successful and unsuccessful collisions. Collision 2 is not successful because of the repulsion between the slightly negatively charged chlorine atom and the electrons in ethene's double bond.

activation energy (E_a) the minimum amount of energy required in a collision for a reaction to occur

- Reactants must collide with a minimum amount of energy (**activation energy, E_a**) to rearrange and form products. If C_2H_4 and HCl in Figure 1 do not collide with sufficient energy, even a correct orientation does not result in a successful collision.

Activation energy

The reaction rate depends on the size of the activation energy. If the energy required to break the bonds of the reactants is relatively low (i.e. the bonds are weak), lower activation energy is required. Conversely, strong bonds have higher bond energy and are harder to break apart, so this results in a higher activation energy.

bond energy
the amount of energy required to break a chemical bond in the gas phase, measure in kJ mol^{-1}

Bond energy is the amount of energy required to break a bond. The higher the bond energy, the harder it is to break. Therefore, bond energy (strength) and the number of bonds also affect the activation energy and how readily a reaction occurs. Typically, if numerous bonds need to be broken in a reaction, this occurs as a series of steps, each with their own activation energy. Here, we will just consider the overall reaction and activation energy.

What is the Maxwell–Boltzmann energy distribution?

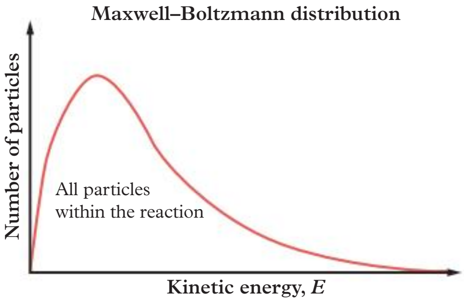


FIGURE 2 A Maxwell–Boltzmann energy distribution curve

The particles within a reaction system do not all contain the same amount of energy. Some have more energy than others. The Maxwell–Boltzmann distribution (Figure 2) represents the number of particles in a reaction system (on the y-axis) and the kinetic energy, E, of each particle (on the x-axis).

The peak of the curve represents the modal kinetic energy of the particles and is relative to the temperature of the reaction. The area underneath the curve represents all particles within the reaction.

A Maxwell–Boltzmann energy distribution is usually plotted with a line representing the E_a of the reaction. Only the particles that have kinetic energy greater than or equal to the activation energy ($E \geq E_a$) will react successfully when they collide.

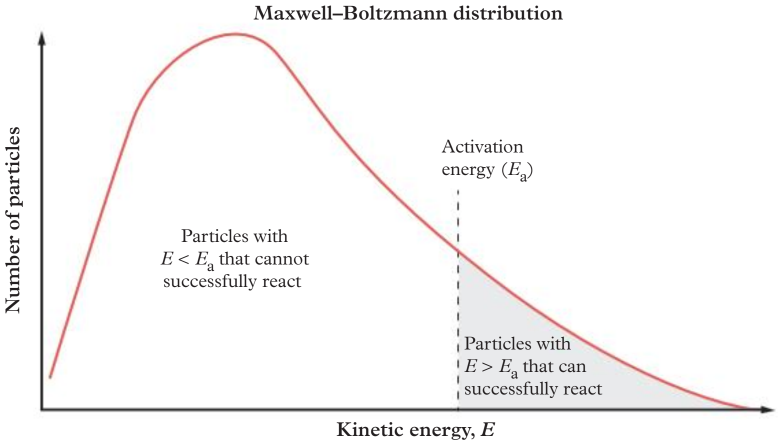


FIGURE 3 The Maxwell–Boltzmann distribution demonstrates the energy of particles in a reaction system.

Study tip
To figure out the best strategy to increase the rate of a reaction, the first thing you should do is identify the states of the reactants and products.

frequency of collisions
how often collisions occur in a reaction vessel in a period of time, i.e. the number of collisions in a unit of time

proportion of successful collisions
the percentage of all collisions which are successful, resulting in the breaking of reactant bonds to form products

What factors influence reaction rate?

Reactions can take anywhere from less than one second to more than a few hundred thousand years, so it is important that scientists can increase the rate of a desired chemical reaction (to make it happen faster).

Reaction rates can be increased by:

- adding energy
- manipulating the activation energy
- increasing the **frequency of collisions** and therefore the **proportion (or number) of successful collisions**.

This can be accomplished by changing five main factors, depending on the state of the reactants and products. They are:

- concentration: liquid, aqueous and gaseous reactants
- pressure: gaseous reactants only
- **surface area**: solids reactants only
- temperature: solids, liquids, aqueous and gaseous reactants
- **catalysts**: solids, liquids, aqueous and gaseous reactants.

surface area
a measure of the total surface area available to react in a chemical reaction

catalyst
a substance that provides an alternative pathway for the reaction to occur that has a lower activation energy and increases the rate of reaction without participating in the chemical reaction

How does concentration affect reaction rate?

Concentration is the measure of the number of particles in a given volume. An increase in concentration means an increase in the number of particles in a given volume. If the number of particles increases, there is an increase in the frequency of reactant collisions (they collide more often) and therefore, there is a higher proportion (percentage) of successful collisions.

If there are more successful collisions, then there will be a faster chemical reaction, as more successful collisions occur in a period of time. Therefore, using a higher concentration of reactants increases the rate of a reaction. Conversely, using dilute, or lower concentration solutions, results in a decrease in the number of successful collisions and a decrease in the rate of reaction.

Study tip

Collisions can occur without resulting in a chemical reaction, so it is essential to explain them in terms of the frequency of collisions and the number or proportion of "successful collisions" when explaining whether a chemical reaction results in the formation of a product.

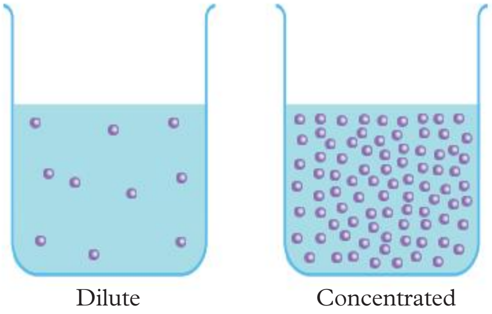


FIGURE 4 A dilute solution has fewer particles than a concentrated solution in a given volume, so it results in fewer collisions.

In aqueous solutions, concentration is the amount of solute in a volume of solvent. In gases, concentration refers to the number of gas molecules in the volume of a container.

Study tip

An increase in concentration will increase the frequency of collisions and therefore, the number of successful collisions and the rate of the reaction.

How does pressure affect reaction rate?

Pressure is the force exerted per unit area by one substance on another substance. For gases, this is the pressure exerted by gas particles when they collide with the walls of a container. An increase in the number of particles in a system results in an increase in pressure.

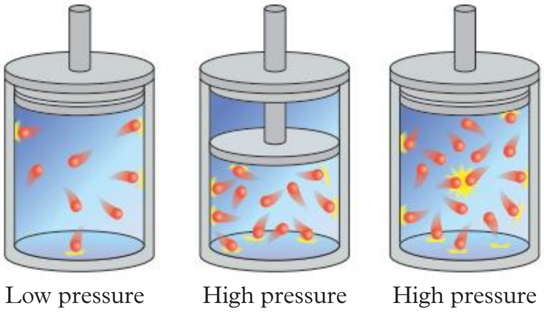


FIGURE 5 Visual representations of low pressure and high pressure gaseous systems

To increase the pressure of a gaseous system, you can:

1 Decrease the volume of the container – this pushes the particles closer together and results in more particles per unit of volume and therefore, a greater concentration of particles (Figure 6). For example, 1,000 oxygen molecules in a 1 L container is less concentrated than 1,000 oxygen molecules in a 100 mL container.

2 Increase the number of particles within the container – in a fixed volume system, the volume cannot be decreased. Therefore, to increase the pressure, more gas must be added to the reaction container, increasing the number of particles and the concentration (Figure 6).

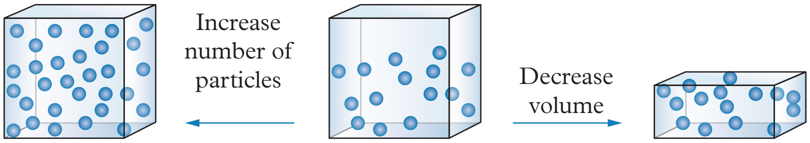


FIGURE 6 Concentration increases when increasing number of particles and decreasing volume.

TABLE 1 The effect on concentration when pressure is increase by increasing the number of particles or decreasing the volume

n	Original gas container with 30 gas molecules	Original gas container with 15 gas molecules	Half of the original container with 15 gas molecules
V	1 L	1 L	500 mL or 0.5 L
c	$c = \frac{n}{V}$ $= \frac{30\,\text{mol}}{1\,\text{L}}$ = 30 particles per litre	$c = \frac{n}{V}$ $= \frac{15\,\text{mol}}{1\,\text{L}}$ = 15 particles per litre	$c = \frac{n}{V}$ $= \frac{15\,\text{mol}}{0.5\,\text{L}}$ = 30 particles per litre

In both methods, an increase in pressure is due to an increase in the concentration of particles. This causes particles to collide more frequently and results in a greater number of successful collisions, which increases the rate of reaction.

How does surface area affect reaction rate?

Reactions occur at the surface of a solid reactant. Therefore, the greater the surface area, the greater the number of collisions. Figure 7 shows a cube with sides of 2 cm and a surface area of 24 cm^2. When the cube is cut into eight smaller cubes with sides of 1 cm, they have a total surface area of 48 cm^2. Therefore, finely divided solids have a larger surface area. This increases the number of successful collisions and the rate of reaction.

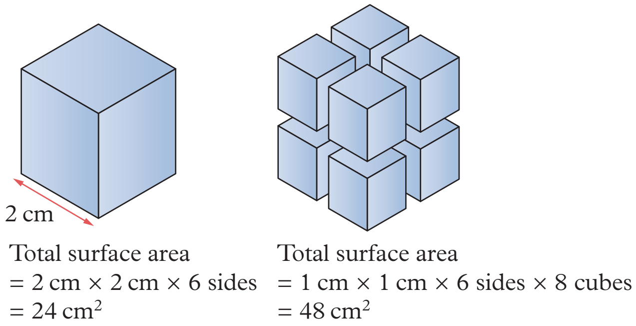


FIGURE 7 The surface area calculation for a bigger 8 cm^3 cube and smaller 1 cm^3 cubes

Study tip

When explaining the effect of pressure, make a link to concentration. Both will increase the frequency of collisions and therefore the proportion of successful collisions, which increases the rate of reaction.

Study tip

An increase in the surface area of the same mass of solid will increase the frequency of collisions and therefore the proportion of successful collisions and the rate of reaction.

Study tip

A common misconception is that the increase in rate of reaction when temperature is increased is due to more collisions occurring, rather than to the increase in energy of the collisions that do occur. The increase in reaction rate is primarily due to the increase in the number of particles that have sufficient energy to react.

How does temperature affect reaction rate?

The activation energy of a reaction is the minimum amount of energy that a collision between particles must have to successfully react. Kinetic molecular theory states that all particles are in constant, random motion. The energy involved with this motion is called kinetic energy. Temperature is a measure of the average kinetic energy in a reaction. Therefore, when a reaction is heated, energy is transferred to the reactants, making them move faster. An increase in temperature results in an increase in energy.

If the temperature of a reaction increases, more particles have energy greater than the activation energy ($E > E_a$). Therefore, many more of the colliding particles have sufficient energy to react.

The Maxwell–Boltzmann curves are drawn for two different temperatures, T_1 and T_2, for the same reaction (Figure 8). The red shaded area contains reactant particles that have sufficient energy to react at the lower temperature, T_1. A small increase in temperature to T_2 causes only a small rise in the average kinetic energy, but a large increase in the shaded area (blue).

Since the average kinetic energy has not increased much, the frequency of collisions increases but only slightly. This plays a very small role in increasing the rate of reaction. Instead, many more of the reactant particles have at least the activation energy and are therefore able to undergo successful collisions to form products.

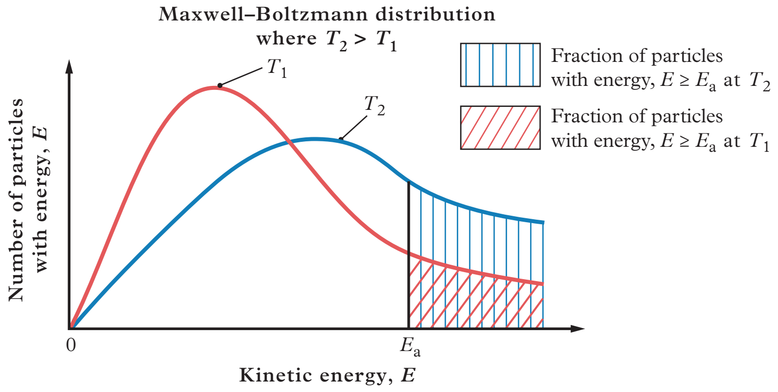


FIGURE 8 The Maxwell–Boltzmann distribution demonstrates the energy of particles in a system. T_1 is a lower temperature than T_2.

Study tip

A "rule of thumb" is that reaction rate will double every 10°C increase in temperature.

Study tip

When drawing reaction curves on a Maxwell–Boltzmann distribution:

- at a higher temperature, the peak of the curve is shifted right and the curve is lower and broader
- at a lower temperature, the peak of the curve is shifted left and the curve is higher and narrower.

Study tip

An increase in temperature will increase the number of particles with energy greater than or equal to the activation energy ($E \geq E_a$). This will increase the proportion of successful collisions and the rate of reaction.

Worked example 16.1A

Recognising the factors that could affect the reaction rate

Worked example 16.1A: Watch a video that shows how to solve this problem.

Consider the following chemical reaction:

$$NO_3(g) + NO(g) \rightarrow 2NO_2(g)$$

Explain, with reference to the frequency and proportion of successful collisions, the factors that could increase the rate of this chemical reaction. (5 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Consider" means to view attentively or scrutinise. "Explain" means to make an idea or situation plain or clear by describing it in more detail or revealing relevant facts. We need to look at the chemical equation, recognise the relevant factors and describe how they can be used to increase reaction rate. The question is worth 5 marks, so we must assess the reaction, correctly identify the factors and describe how they can be manipulated.

Think	Do
Step 2: Consider the states of the reactants and products.	All reactants and products are gases, so this is a gaseous system.
Step 3: Determine which factors are relevant to the states identified.	Temperature, concentration and pressure could be altered to change the rate of this reaction. (1 mark for correctly identifying factors)
Step 4: Explain how manipulating each factor could increase the frequency and proportion (or number) of successful collisions.	The temperature of the reaction could be increased to increase the kinetic energy of gas particles in the system. This results in an increase in the number of particles with energy greater than, or equal to, the activation energy, and thus the proportion of successful collisions. (1 mark) Increasing the concentration of reactants brings the particles closer together. This increases the chance of reactants colliding and therefore increases the frequency of collisions and the proportion of successful collisions. (1 mark) Increasing the pressure of the system would also increase the chance of particles colliding, and therefore the frequency of collisions and the proportion of successful collisions. (1 mark) All three factors increase the frequency of collisions, the proportion of successful collisions and therefore the rate of reaction. (1 mark)

Your turn

Consider the following chemical reaction:

$$2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$$

Explain, with reference to the frequency and proportion (or number) of successful collisions and the surface area of the solid reactant, the factors that could increase the rate of this chemical reaction. (5 marks)

How do you measure the rate of a chemical reaction?

The rate of a chemical reaction is measured as a change over time. This change can be the loss in mass of a solid reactant, the volume of gas generated, the formation of a solid precipitate, a change in colour, an increase in product concentration or a decrease in reactant concentration.

Mathematically, the rate of a chemical reaction can be measured in two ways. The first is to create a graph of the change in a reaction over time. Figure 9 represents the mass of a reactant lost over time. The rate of the chemical reaction starts faster and gradually slows over time.

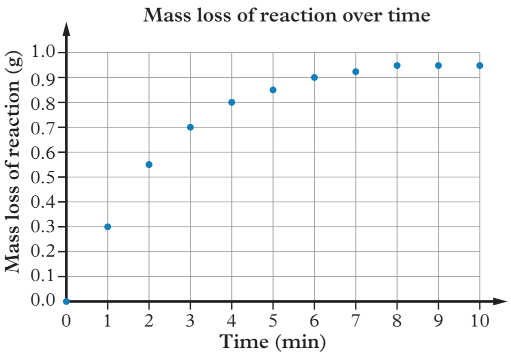


FIGURE 9 The mass of a reactant lost over time

The second measurement involves using a mathematical formula to calculate the rate of reaction at a certain time in the experiment or to calculate the overall (average) rate of reaction at the end:

$$\text{Rate of reaction} = \frac{\text{increase in product concentration } (\Delta[P])}{\text{time taken}}$$

or

$$\text{Rate of reaction} = \frac{\text{decrease in reactant concentration } (-\Delta[R])}{\text{time taken}}$$

More practically, the numerator could be changed to give:

$$\text{Rate of reaction} = \frac{\text{change in mass of reactant}}{\text{time}}$$

If the rate of the reaction was only measured as the final loss in mass over the total time, the rate of the reaction would look very different because it is an average rate calculation. By monitoring the mass loss at various time intervals throughout the experiment, a more detailed picture of the rate of reaction can be observed, particularly how the rate changes over the course of the reaction.

Mass lost over time

Mass can be lost from a solid reactant in a chemical reaction. For example, in the reaction between solid zinc and hydrochloric acid, mass is lost when hydrogen gas is formed as the gas can escape from an open container into the air:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

First, the mass of hydrochloric acid and zinc are measured on an electronic balance (Figure 10). Hydrochloric acid is then poured onto the zinc and the initial mass is recorded. The subsequent loss of mass, which occurs as hydrogen gas escapes, is measured in time intervals until the reaction reaches a constant mass.

A graph of mass (on the *y*-axis) against time (on the *x*-axis) is then plotted (Figure 11). The data in the graph should never reach zero because zinc chloride and the water solvent will remain in the flask when the reaction is complete.

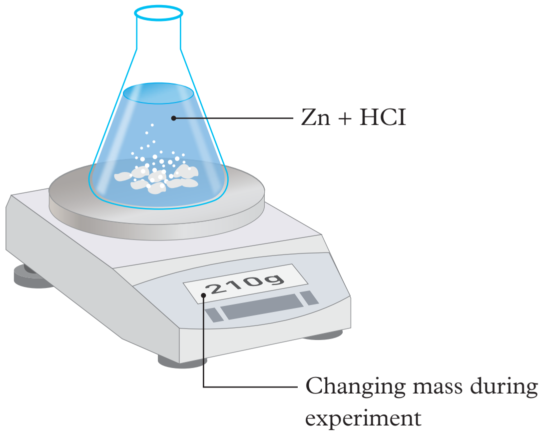


FIGURE 10 The process of measuring mass lost on an electronic balance

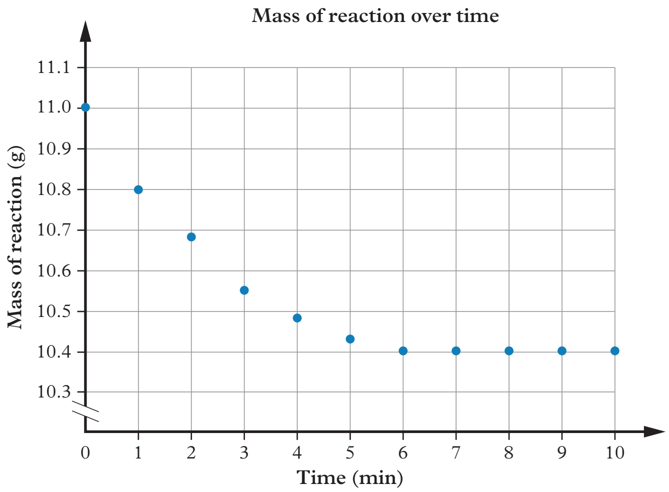


FIGURE 11 Graph of mass lost over time

Volume of gas formed over time

In this method, the gas generated is used to displace a known volume of water. Figure 12 demonstrates a gas being generated in a conical flask. The gas moves through the connecting tube into the trough of water and into a gas jar, which has graduated measurements marked on its side. The water is displaced as the gas enters the jar and the volume of gas generated can be measure.

A graph of volume (on the *y*-axis) against time (on the *x*-axis) is then plotted (Figure 13). The volume of gas is measured using the graduated markings on the beaker that collects the gas.

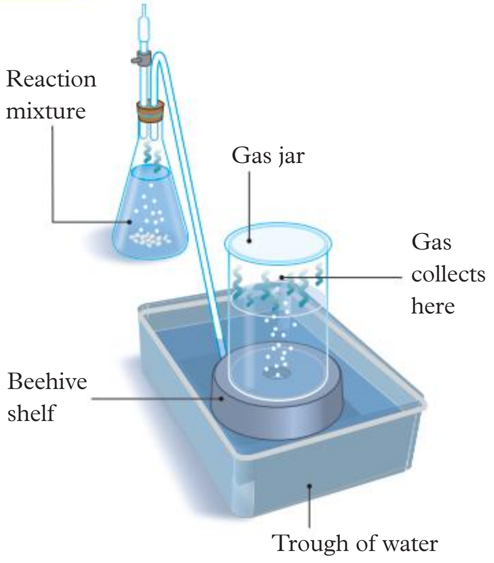


FIGURE 12 The process of measuring gas volume by water displacement

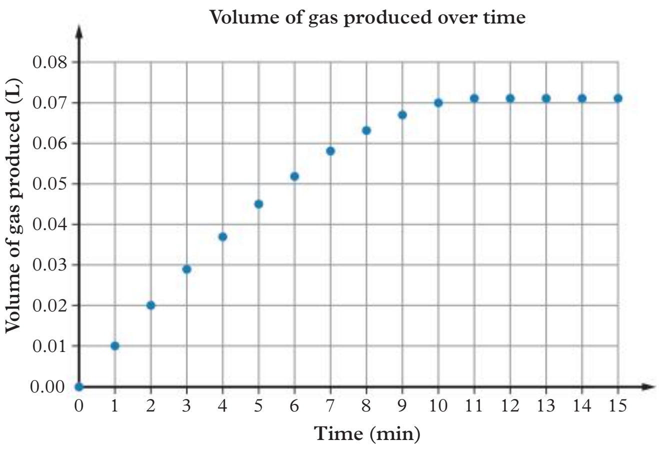


FIGURE 13 Graph of volume of gas produced over time

Worked example 16.1B

Calculating rates of chemical reaction

Worked example 16.1B: Watch a video that shows how to solve this problem.

Calculate the rate of reaction for the reaction shown in Figure 9:

a in the first minute (1 mark)

b for the overall reaction (1 mark)

c between 1 and 2 minutes (1 mark)

d between 3 and 4 minutes. (1 mark)

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the rate of the reaction at various time intervals. The questions are worth 1 mark each, so we must gather the correct information, correctly apply the formula and complete the calculations.
Step 2: Select the correct formula and gather any data required. You will need to read the mass values from Figure 9.	Rate of reaction = $\frac{\text{change in mass of reactant}}{\text{time}}$ **a** Time = 1 min **c** Time = 1 min **b** Time = 8 min **d** Time = 1 min
Step 3: Substitute the known values into the formula and solve for the unknown value.	**a** Rate of reaction = $\frac{0.3\,\text{g} - 0\,\text{g}}{1\,\text{min}}$ $= 0.3\,\text{g}\,\text{mol}^{-1}$ **b** Rate of reaction = $\frac{0.95\,\text{g} - 0\,\text{g}}{8\,\text{min}}$ $= 0.1188\,\text{g}\,\text{mol}^{-1}$ **c** Rate of reaction = $\frac{0.55\,\text{g} - 0.3\,\text{g}}{1\,\text{min}}$ $= 0.25\,\text{g}\,\text{mol}^{-1}$ **d** Rate of reaction = $\frac{0.8\,\text{g} - 0.7\,\text{g}}{1\,\text{min}}$ $= 0.1\,\text{g}\,\text{mol}^{-1}$
Step 4: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures. Note that because the graph is curved, these calculations represent an average rate over that time period.	**a** $0.3\,\text{g}\,\text{min}^{-1}$ (1 mark) **c** $0.25\,\text{g}\,\text{min}^{-1}$ (1 mark) **b** $0.12\,\text{g}\,\text{min}^{-1}$ (1 mark) **d** $0.1\,\text{g}\,\text{min}^{-1}$ (1 mark)

Your turn

Calculate the rate of reaction for the reaction shown in Figure 14:

a in the first minute (1 mark)

b for the overall reaction (1 mark)

c between 1 and 2 minutes (1 mark)

d between 2 and 3 minutes. (1 mark)

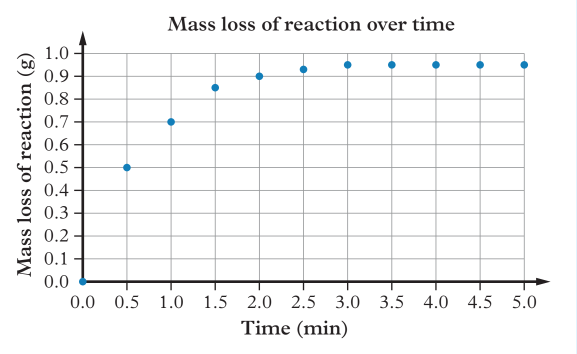


FIGURE 14 Graph of mass gained over time

How do you calculate the end point of a reaction?

The rate of a reaction can also affect the **yield** or amount of product formed from a reaction. If a reaction occurs too slowly, it will not reach its end point. A theoretical end point for a reaction can be calculated using stoichiometry.

yield
the amount of product formed from a chemical reaction

Consider a reaction between solid magnesium metal and hydrochloric acid which produces magnesium chloride and hydrogen gas:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

If the amount (mole) of reactant is known, the theoretical yield can be calculated.

A key part of preparation for an experiment is calculating the theoretical yield, so this can be compared with the experimental value.

Worked example 16.1C

Calculating the theoretical end point of a reaction

Worked example 16.1C: Watch a video that shows how to solve this problem.

A student reacts 2.00 g of magnesium ribbon with an excess of hydrochloric acid at STP. **Determine** the theoretical volume of hydrogen gas that could be generated. (4 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Calculate" means to determine or find a number or answer by using mathematical processes. We need to determine the rate of the reaction at various time intervals. The question is worth 3 marks, so we must gather the correct information, correctly apply the formulas and complete the calculations.
Step 2: Write a balanced chemical equation for the reaction, including states.	$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Step 3: Select the correct formulas and gather any data required.	$n = \frac{m}{M}$ $n(Mg) = ?, m = 2.00\,g, M = 24.31\,g\,mol^{-1}$ $n = \frac{V}{V_m}$ $n(H_2) = ?, V = ?, V_m = 22.7\,L\,mol^{-1}$

Think	Do
Step 4: Substitute the known values into the formula and solve for the moles of magnesium.	$n(\text{Mg}) = \frac{2.00\,\text{g}}{24.31\,\text{g mol}^{-1}}$ (1 mark) $= 0.0823\,\text{mol}$ (1 mark)
Step 5: Use the mole ratio from the balanced equation to calculate the moles of the hydrogen gas.	$n(H_2) = n(\text{Mg})$ (1 mark) $= 0.0823\,\text{mol}$
Step 6: Substitute the known values into the formula and solve for the volume of hydrogen gas at SLC.	$V(H_2) = 0.0823\,\text{mol} \times 22.7\,\text{L mol}^{-1}$ (1 mark) $= 1.868\,\text{L}$
Step 7: Finalise your answer and make sure you have expressed it using the correct units and number of significant figures.	1.87 L (1 mark)

Your turn

In a reaction between calcium carbonate ($CaCO_3$) and nitric acid (HNO_3), a student reacts an excess of calcium carbonate with 20.0 mL of 2.0 M nitric acid at STP. **Calculate** the theoretical volume of carbon dioxide gas that could be generated. (4 marks)

Skill drill

Evaluating an experiment involving an acid–carbonate reaction

Science inquiry skill: Evaluate evidence (Lesson 1.8)

Accuracy is the closeness of an experimental value to its actual or true value. In this case, it is the theoretical mass lost in the reaction (a value that is calculated). In a reaction between hydrochloric acid and sodium hydrogen carbonate, a solution of sodium chloride salt and carbon dioxide gas are produced according to the following equation:

$$HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

In the experiment, 50.0 mL of 0.33 M $NaHCO_3$ was reacted with 100 mL of 0.50 M HCl. On completion of the reaction, it was found to have lost a mass of 0.71 g. When the reaction was repeated a further two times, the reaction had lost a mass of 0.70 g and 0.72 g.

Practise your skills

1 **Explain** why the reaction loses mass. (2 marks)
2 **Calculate** the theoretical mass that can be lost based on the quantities of the reactants. (7 marks)
3 **Comment** on the accuracy of the experimental repeats, i.e. which is more accurate. **Use** the data to **justify** your answer. (2 marks)
4 **Comment** on the precision of the data. (2 marks)

Real-world chemistry

The problem with dust – the sugar refinery explosion

At 7 pm on 7 February 2008, an explosion rang out across the town of Port Wentworth, Georgia, USA. The blast originated in the centre of the Georgia Sugar Refinery, located on the banks of the Savannah River. The refinery, built in 1916, was old. Its construction features were outdated, and many staff members said the machinery was antiquated, some of it being close to 30 years old.

The explosion took place in the centre of the refinery, where the sugar was stored before being bagged and packaged.

Refined sugar is sucrose, with the molecular formula $C_{12}H_{22}O_{11}$. Sucrose undergoes combustion according to the chemical equation:

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(g)$$

Sugar is naturally highly combustible; however, as a fine powder or dust mixed with air, it can spontaneously combust (react with oxygen to produce large amounts of heat). The storage of the sugar led to a build-up of sugar dust within the air. The building was equipped with ventilation fans that extracted dust particles from the air, but these were found to contain enough dust to cause further explosions. The ignition source was probably static electricity or a spark, but spontaneous combustion was not ruled out.

The fire took seven days to extinguish because of the age of the building and the quantity of sugar that continued to act as a fuel source. The task of putting out the fire was made worse because the sugar that did not combust melted at temperatures above 186°C.

The explosion killed 14 people and severely injured 36.

FIGURE 15 The aftermath of the Georgia Sugar Refinery explosion

Apply your understanding

1 **Use** your knowledge of the rate of a reaction to **explain** why sugar dust is more explosive than sugar cubes, with reference to the frequency and proportion of successful collisions. (3 marks)

2 **Explain** two factors affecting the rate of a reaction that could have been put in place to avoid the explosion. (4 marks)

3 **Discuss** the safety precautions that should be put in place when building a new refinery to avoid future disasters. (2 marks)

Check your learning 16.1

Check your learning 16.1: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Identify** and **describe** the two requirements of collision theory. (2 marks)

2 **Recall** the factors that increase the rate of a chemical reaction. (5 marks)

Analytical processes

3 Some students are conducting an experiment by reacting 5.00 g of calcium carbonate ($CaCO_3$) chips with 50.00 mL of 2.00 M hydrochloric acid (HCl(aq)). The reaction produced carbon dioxide gas, water and calcium chloride solution according to the chemical equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$$

a Assuming that the hydrochloric acid is in excess, **calculate** the volume of carbon dioxide gas produced at SLC. (4 marks)

b The volume of gas generated was measured and recorded in Table 1. **Sketch** a graph of the rate of reaction. (2 marks)

TABLE 1 Results data of volume of gas produced from a reaction between calcium carbonate and hydrochloric acid

Time (min)	Volume of gas (L) produced
0.00	0.000
0.50	0.200
1.00	0.350
1.50	0.450
2.00	0.500
2.50	0.525
3.00	0.530
3.50	0.530
4.00	0.540
4.50	0.540
5.00	0.540

c **Calculate** the overall rate of reaction. (1 mark)

d **Compare** the theoretical and experimental results. (4 marks)

The students wish to repeat the experiment to increase the rate of reaction.

e **Identify** three things that they could incorporate into their methodology and **explain** how each would increase the rate of the reaction. (3 marks)

f **Identify** factor/s that would have no effect on the rate of this reaction. (1 mark)

4 In the presence of water, iron rusts according to the equation:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

A car manufacturer has found that their car parts are rusting before the painting and assembly processes. **Determine** what the manufacturer could do to slow down the rust or avoid it entirely. (2 marks)

Knowledge utilisation

5 In a reaction between sodium thiosulfate ($Na_2S_2O_3(aq)$) solution and hydrochloric acid ($HCl(aq)$), a yellow precipitate forms according to the equation reaction:

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow S(s) + SO_2(g) + 2NaCl + H_2O(l)$$

The precipitate is so fine that it remains suspended in the reaction mixture rather than settling to the bottom. **Develop** a method for measuring the rate of the reaction. (3 marks)

6 **Discuss** the benefits of using granulated sugar instead of sugar cubes to sweeten tea and coffee. (3 marks)

Practical

Lesson 16.2

Investigating the rate of chemical reactions

Learning intentions and success criteria

Video demonstration

oxforddigital

This practical lesson is available on Oxford Digital. It is also provided as part of a printable resource that can be used in class.

Lesson 16.3

Energy profile diagrams

Key ideas

→ Energy profile diagrams show the change in energy as the reaction progresses over time.

→ If energy is released in a reaction, it is exothermic. If energy is absorbed in a reaction, it is endothermic.

→ The stronger the bonds broken in a reaction, the higher the activation energy and the slower the reaction.

How does energy change during a reaction?

Learning intentions and success criteria

The law of conservation of energy states that energy cannot be created or destroyed. Chemical energy is potential energy. This energy can be most readily converted into heat, observed as an increase in temperature, or motion, as kinetic energy. Energy differences between reactants and products result in either an exothermic reaction (where the products have less energy than the reactants) or an endothermic reaction (where the products have more energy than the reactants).

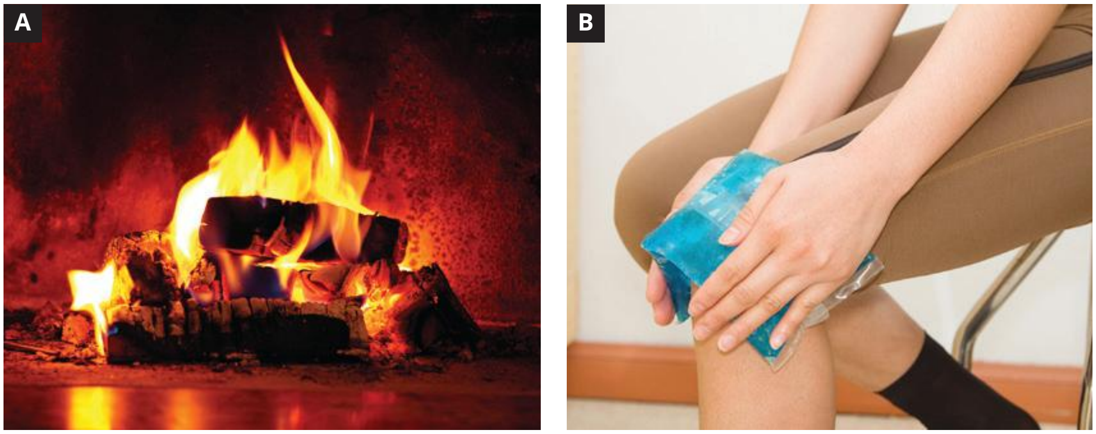


FIGURE 1 (A) The burning of firewood is an exothermic reaction and (B) the cooling of a chemical ice pack is an endothermic reaction.

Study tip

Think of the terms *exothermic* and *endothermic* like a party. *Exo* means "exit", when friends leave the party – it's bad, so it is negative. *Endo* means "enter", when friends arrive at the party – it's great, so it is positive.

What do energy profile diagrams show?

Energy profile diagrams represent the energy changes that occur in chemical reactions. These diagrams show the difference in energy within the bonds of reactants and products. They are similar to the enthalpy level diagrams you learnt about in Module 8.

In an energy profile diagram (Figure 2), energy is plotted on the y-axis and time, or reaction progress, is plotted on the x-axis. If the reactants contain more energy within their bonds than the products, energy is released as heat and the reaction is exothermic. If the reactants contain less energy than the products, energy must be absorbed from the environment and the reaction is endothermic.

energy profile diagram
a theoretical representation of the energy pathway that reactants must follow in order to form products

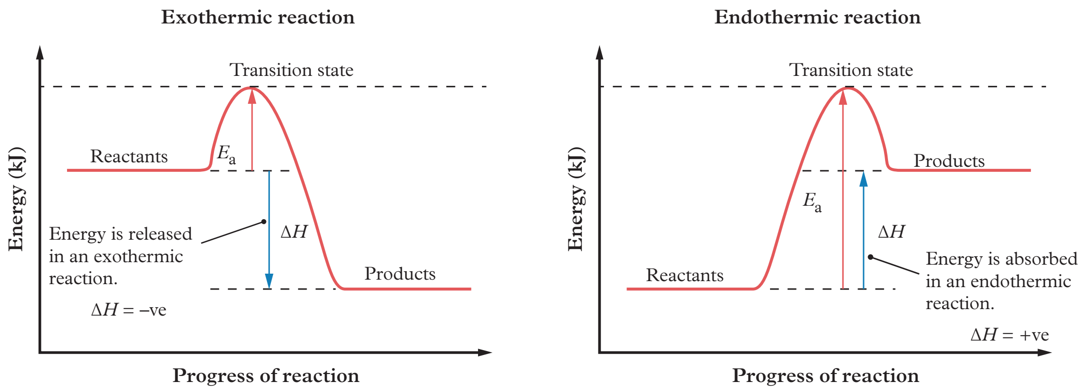


FIGURE 2 Energy profile diagrams for exothermic and endothermic reactions, including E_a (activation energy, energy required to break bonds) and ΔH (enthalpy)

Remember from Module 8 that the change in enthalpy is measured as:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

An energy profile diagram includes a number of key features.

1 The left side of the curve or "hill" is the kinetic energy of the colliding reactant particles that must be converted to potential energy to break reactant bonds.
2 The peak of the curve is the **transition state** or activated complex where potential energy is at its greatest and the arrangement of atoms is the most unstable. At this point, bonds are both breaking and forming into products.
3 The right side of the curve is the energy released when product bonds are formed.
4 If reactants have more chemical potential energy than products, their energy is released into the environment as heat. Therefore, the change in enthalpy between reactants and products is negative (–ΔH).
5 If products have more energy than reactants energy is absorbed from the environment as heat. Therefore, the change in enthalpy between reactants and products is positive (+ΔH).

transition state
an intermediate state in which unstable, loosely arranged atoms reach the highest energy level (activation energy) of a reaction pathway, where bonds are breaking and reforming; also called the activated complex

Study tip

The Maxwell–Boltzmann distribution and energy profile diagrams appear similar in shape. Take the time to identify and understand the differences. The Maxwell–Boltzmann distribution shows the kinetic energy distribution of all of the particles in the reaction mixture, with kinetic energy on the *x*-axis and number of particles on the *y*-axis. Energy profiles show the potential energy of reactants, the transition state, and products as the reaction proceeds, with the progress of the reaction on the *x*-axis and potential energy on the *y*-axis.

Worked example 16.3A

Constructing energy profile diagrams

Worked example 16.3A: Watch a video that shows how to solve this problem.

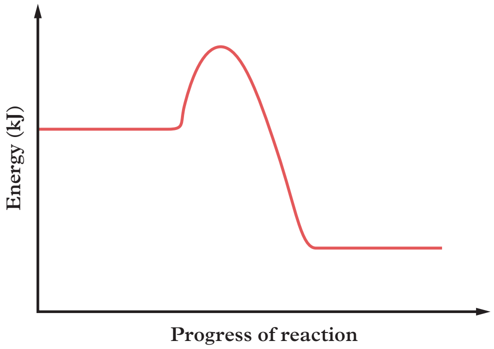
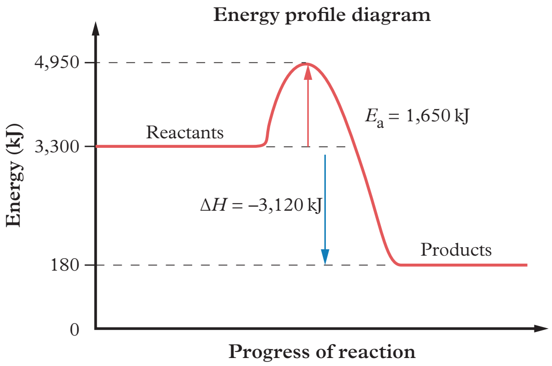
The combustion of ethane forms carbon dioxide and water based on the following equation:

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l) \qquad \Delta H = -3{,}120\text{ kJ}$$

The reactants have 3,300 kJ of energy and the activation energy is 1,650 kJ.

Construct an energy profile diagram for the reaction including labelled axes, reactants and products, activation energy (E_a) and enthalpy (ΔH). (5 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	"Construct" means to display information in a diagrammatical or logical form. We need to use the information given to construct an energy profile diagram. The question is worth 5 marks, so we must represent the information correctly.

Think	Do
Step 2: Draw a graph and label the *y*-axis “Energy (kJ)” and the *x*-axis “Progress of reaction”. Look at the sign in front of the ΔH value to determine whether the reaction is exothermic or endothermic – this will determine the shape of your graph.	The reaction is exothermic, so the energy at the start of the reaction is greater than the energy at the end. 
Step 3: Select the correct formula and gather any data required.	$\Delta H = H_{\text{products}} - H_{\text{reactants}}$ $\Delta H = -3{,}120\,\text{kJ}$, $H_{\text{reactants}} = 3{,}300\,\text{kJ}$, $H_{\text{products}} = ?$ $E_a = 1{,}650\,\text{kJ}$
Step 4: Substitute the known values into the formula and solve for the unknown value.	$-3{,}120\,\text{kJ} = H_{\text{products}} - 3{,}300\,\text{kJ}$ $H_{\text{products}} = 1{,}80\,\text{kJ}$
Step 5: Use the values obtained to label the diagram.	 (1 mark for labelled axes, 1 mark for correct graph shape, 1 mark for reactants and products indicated with correct energy, 1 mark for ΔH label, 1 mark for E_a label)

Your turn

Hydrogen and oxygen are formed from the electrolysis of water according to the following equation:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g) \qquad \Delta H = +484\,\text{kJ}$$

The reactants have 700 kJ of energy and the activation energy is 1,000 kJ. **Construct** an energy profile diagram for the reaction including labelled axes, reactants and products, activation energy (E_a) and enthalpy (ΔH). (5 marks)

What is the relationship between the magnitude of activation energy and reaction rate?

Reaction rates depend on the size of the activation energy. If the energy contained within the bonds of the reactants is relatively small (i.e. the bonds are weak), a smaller activation energy is required. Conversely, strong bonds contain larger amounts of energy, are harder to break apart and require a larger activation energy.

Bond energy, or bond strength, is the amount of energy required to separate atoms that are bonded together in the gaseous phase. The higher the bond energy, the more energy required to break the bond. Therefore, bond energy (strength) and the number of bonds that must be broken can also affect the activation energy and how readily a reaction occurs.

Consider octane and propane (Figure 3). The C–C bonds in both molecules take around 348 kJ mol^{-1} to break (in the gas phase), and the C–H bonds are also of similar energy (413 kJ mol^{-1}). We would therefore expect octane and propane to react at similar rates for reactions that involved breaking C–C or C–H bonds.

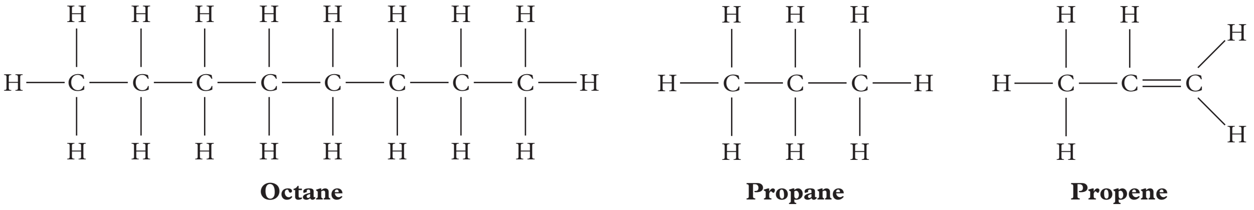


FIGURE 3 Structural formulas of octane, propane and propene

Considering the difference between propane and propene, the double bond in propene has two components. For an addition reaction (where new bonds are formed and the double bond between C atoms is replaced with a single bond), only one of these components needs to be broken, requiring about 264 kJ mol^{-1} of energy, so we would predict that an addition reaction with propene would be faster than a reaction with propane that requires breaking a C–C single bond, which requires more energy.

Skill drill

Evaluating an experiment involving an acid–metal reaction

Science inquiry skill: Evaluating evidence (Lesson 1.8)

A commonly used experiment is the reaction of magnesium metal with hydrochloric acid to produce magnesium chloride and hydrogen gas. In this experiment, 1.00 g of powdered magnesium reacted with 50.0 mL of 2.0 M hydrochloric acid produces 0.050 g of hydrogen gas.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

As there are many factors which can affect the rate of a chemical reaction, scientists need to ensure that all conditions are optimised to produce the highest quantity of gas in the smallest amount of time.

Practise your skills

1 **Calculate** the theoretical volume of hydrogen gas produced at STP. (7 marks)
2 **Use** the data to **comment** on the accuracy of the experiment. (6 marks)
3 **Describe** how the precision of the experiment can be increased. (2 marks)
4 **Explain** what it means to conduct a repeat of an experiment. **Describe** the effect this will have on the precision and accuracy of the data. (3 marks)
5 **Explain** the effect of repeats on experimental errors, with reference to systematic and random errors. (4 marks)

Real-world chemistry

Cold and hot packs

It is possible to buy small plastic pouches that contain chemicals that can either react and consume or produce heat. These cold and heat packs can be used for a variety of purposes such as cooling injuries and warming hands. The reactants are kept apart within the packaging and are reacted when needed.

In heat packs, two main chemical reactions can be used. These exothermic reactions occur quickly, releasing heat as needed before the energy is transferred into the environment around the pack. They are the reaction between iron and oxygen and the combination reaction involving anhydrous sodium acetate and water:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

$$CH_3COONa{\bullet}3H_2O(aq) \rightarrow CH_3COONa(s) + 3H_2O(l)$$

Cold packs do not need to be kept in a freezer before use. They rely on endothermic reactions which absorb heat from their surroundings and therefore, feel cool to the touch. The main reaction used is the dissociation of ammonium nitrate in water:

$$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

FIGURE 4 Heat packs work when two exothermic reactions produce heat.

Apply your understanding

1 **Construct** endothermic reactions from the two exothermic reactions involved in heat packs. (2 marks)

2 **Construct** a general energy profile for the reaction of iron. (5 marks)

3 With reference to the energy of reactants and products, **explain** why endothermic reactions feel cold. (2 marks)

Check your learning 16.3

Check your learning 16.3: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Describe** the term "activation energy" and **explain** how it affects the rate of a chemical reaction. (2 marks)

2 **Define** "transition state". (1 marks)

3 **Construct** an energy profile diagram for an endothermic reaction. Label both axes, reactants, products, E_a, ΔH, energy absorbed and energy released. (5 marks)

4 **Construct** an energy profile diagram for an exothermic reaction. Label both axes, reactants, products, E_a, ΔH, energy absorbed and energy released. (5 marks)

Analytical processes

5 **Consider** the following equation:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Methane and oxygen have an enthalpy of 877.1 kJ mol^{-1}. Carbon dioxide and water have an enthalpy of 74.8 kJ mol^{-1}. The activation energy for this reaction is 2,632 kJ mol^{-1}.

a **Use** this information to **construct** an energy profile diagram for the reaction. (5 marks)

b **Determine** the E_a and ΔH for the reverse reaction. **Classify** each reaction as endothermic or exothermic. (3 marks)

Lesson 16.4

Catalysts

Key ideas

→ Catalysts increase the rate of a reaction by providing an alternative pathway that requires a lower activation energy to form the same products.

→ The effect of a catalyst can be graphically represented on a Maxwell–Boltzmann distribution and on an energy profile diagram.

Learning intentions and success criteria

catalyst
a substance that provides a lower activation energy for a reaction and increases the rate of reaction without participating in the chemical reaction

homogeneous catalyst
a catalyst that is in the same state as the reactants

heterogeneous catalyst
a catalyst that is in a different state from the reactants

surface energy
excess energy at the surface of a material

What are catalysts?

Catalysts increase the rate of a chemical reaction without being used up in the reaction. They do this by providing a different way for the reaction to occur that has a lower activation energy. This is the energy required for a successful collision. Catalysts are essential in the human body, medicines, pharmaceuticals, the chemical industry, car exhaust systems and many other science-related fields. Catalysts include enzymes and metal nanoparticles.

There are two types of catalyst:

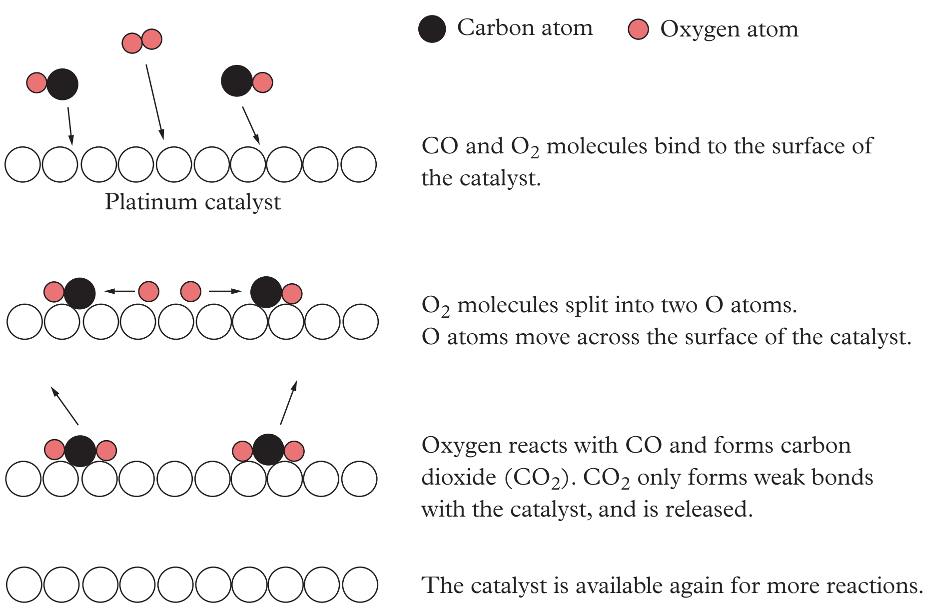
- **Homogeneous catalysts** are in the same state as the reactants. For example, in the dissociation of aqueous hydrogen peroxide ($H_2O_2(aq)$), an aqueous potassium iodide catalyst is used.
- **Heterogeneous catalysts** are in a different state from the reactants. For example, in the reaction between hydrogen and nitrogen gases to produce ammonia, a solid iron catalyst is used.

How do catalysts work?

Catalysts affect the rate of reactions by providing an alternative reaction pathway with reduced activation energy. The mechanisms by which catalysts can achieve a lower activation energy are of great interest to scientists.

For example, scientists know that heterogenous catalysts have high **surface energy**. When reactants strike the surface of the catalyst, they bond to it. The formation of these bonds weakens or breaks the bonds contained within the reactants. The atoms then rearrange on the catalyst's surface to form products (Figure 2).

FIGURE 1 Some solid catalysts



CO and O_2 molecules bind to the surface of the catalyst.

O_2 molecules split into two O atoms.
O atoms move across the surface of the catalyst.

Oxygen reacts with CO and forms carbon dioxide (CO_2). CO_2 only forms weak bonds with the catalyst, and is released.

The catalyst is available again for more reactions.

FIGURE 2 The action of a catalyst in lowering the energy required for reaction of carbon monoxide and oxygen

Effect of a catalyst on Maxwell–Boltzmann distributions

The catalysed reaction can be represented on a Maxwell–Boltzmann distribution as a line, similar to the one that represents activation energy. As a catalyst lowers the activation energy of a system, the number of particles with energy greater than the activation energy ($E > E_a$) increases (Figure 3), and the proportion of successful collisions increases but not the frequency of them.

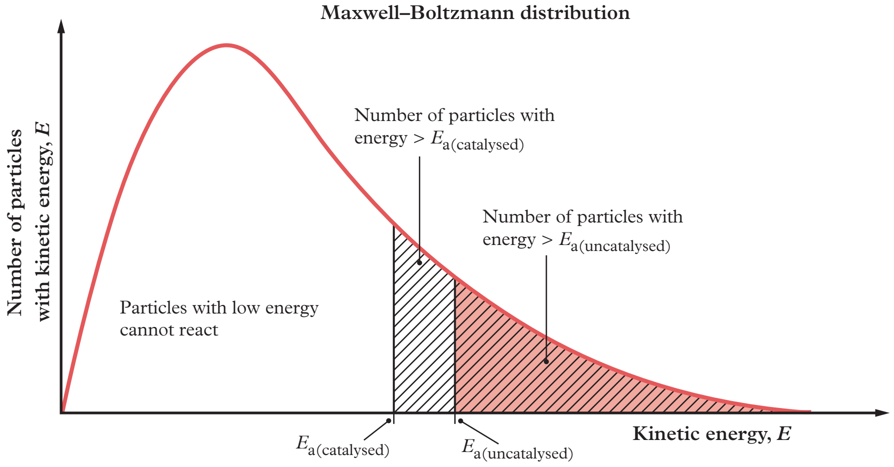


FIGURE 3 The Maxwell–Boltzmann distribution demonstrating activation energy (E_a) and the lowered E_a with a catalyst. The red shaded area contains particles that can react without a catalyst present. The entire black hashed area corresponds to particles that can react with a catalyst present.

Effect of a catalyst on energy diagrams

Because a catalyst provides a lower E_a of reaction, it can also be represented on an energy profile diagram. In Figure 4, the catalysed chemical pathway is represented as a broken line below the uncatalysed chemical pathway. The activation energy of the catalysed pathway (E_c) is lower than E_a, but the enthalpy of the reaction (ΔH) does not change.

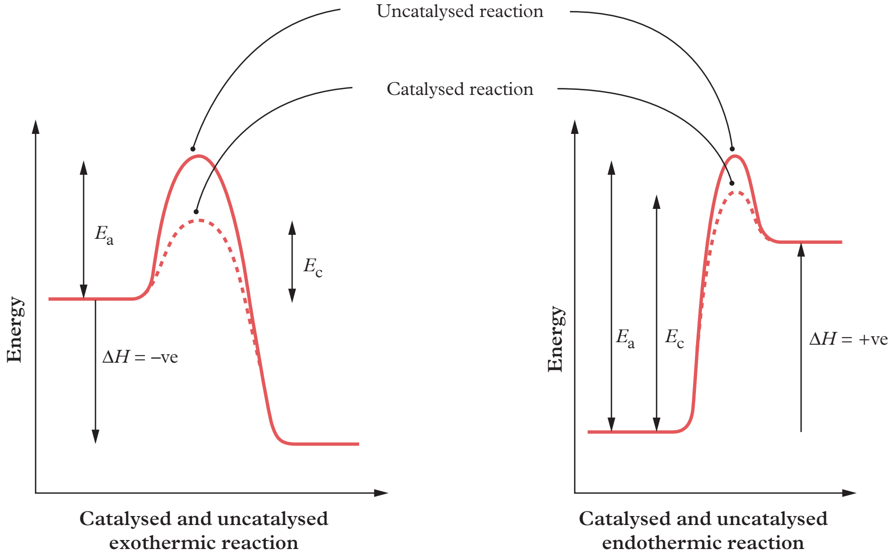


FIGURE 4 Energy profile diagrams of exothermic and endothermic reactions, demonstrating the regular reaction pathway and the catalysed reaction pathway with lowered activation energy(E_c)

Worked example 16.4A

Interpreting a Maxwell–Boltzmann distribution curve

Worked example 16.4A: Watch a video that shows how to solve this problem.

Consider this Maxwell–Boltzmann distribution.

For each of the following, **construct** a curve or line to represent the effect of the change on the same Maxwell–Boltzmann distribution. **Use** the diagram to **explain** the consequences on the rate of the reaction.

a an increase in temperature (3 marks)

b the addition of a catalyst. (3 marks)

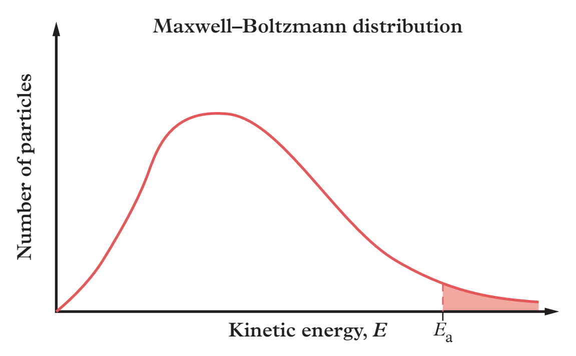
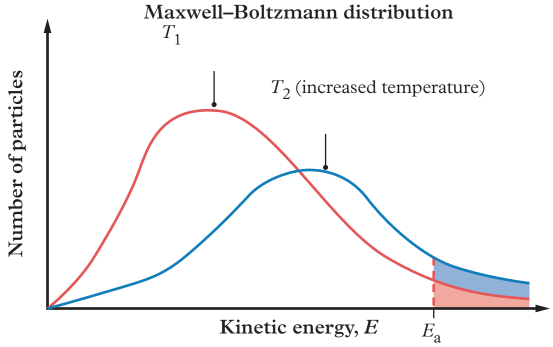
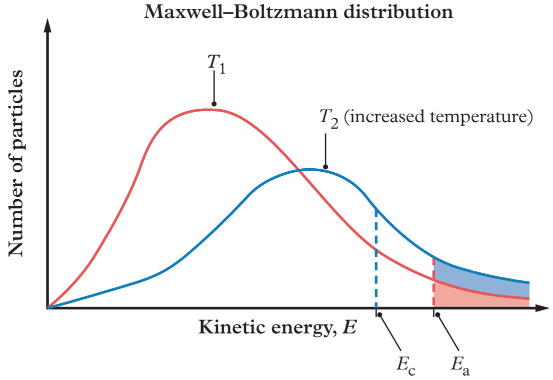


FIGURE 5 Maxwell–Boltzmann distribution

Think	Do
Step 1: Look at the cognitive verbs and mark allocation to determine what the questions are asking you to do.	"Construct" means to display information in a diagrammatical or logical form. "Use" means to apply knowledge or rules to put theory into practice. "Explain" means to make an idea or situation plain or clear by describing it in more detail or revealing relevant facts. We need to use the information given to construct new curves on the energy profile diagram, then interpret the meaning of the curves. The questions are worth 3 marks each, so we must represent the information correctly and explain the effect on reaction rate.

Think	Do
Step 2: For part **a**, consider what happens to the energy of the particles when temperature increases. Represent this on the Maxwell–Boltzmann distribution by drawing a new curve. Remember that the number of particles remains the same, so the area under the curve is constant but the distribution and its peak will shift.	When temperature is increased, the peak of the curve will shift to the right and become flatter. This is because all particles gain energy. (1 mark) 
Step 3: Refer to the shape and/or proportions of the curve to explain the effect of increasing temperature on the rate of the reaction. You must also discuss the frequency of collisions and proportion of successful collisions.	Increasing temperature results in an increase in the frequency of collisions, as there are more particles with a higher energy as shown by the area under the blue curve compared to the red. (1 mark) This corresponds to a higher proportion of collisions with $E \geq E_a$ and therefore, an increase in the rate of the reaction. (1 mark)
Step 4: For part **b**, consider what happens when a catalyst is added. Represent this on the same Maxwell–Boltzmann distribution by drawing a new curve.	When a catalyst is added, the curve remains the same but the activation energy is lower. The new activation energy is E_c. (1 mark) 
Step 5: Refer to the shape and/or proportions of the curve to explain the effect of adding a catalyst on the rate of the reaction. You must also discuss the frequency of collisions and proportion of successful collisions.	The addition of a catalyst provides a lower activation energy. A catalyst does not increase the frequency of collisions but it does increase the proportion of them that are successful. (1 mark) From the curve, there is a higher proportion of successful collisions with $E \geq E_a$, resulting in a faster rate of reaction. (1 mark)

Your turn

Consider the Maxwell–Boltzmann distribution.

Construct a curve to represent the effect of doubling the concentration on the Maxwell–Boltzmann distribution. **Use** the diagram to **explain** the consequences on the rate of the reaction. (3 marks)

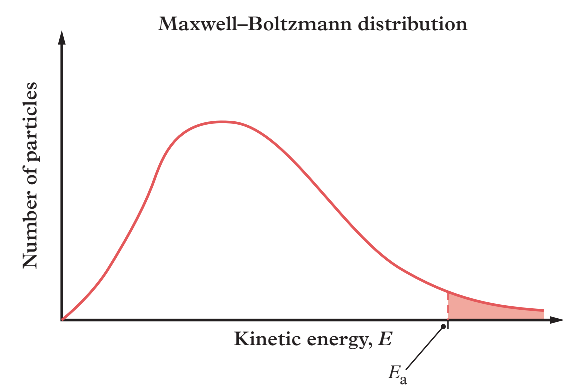


FIGURE 6 Maxwell–Boltzmann distribution

Worked example 16.4B

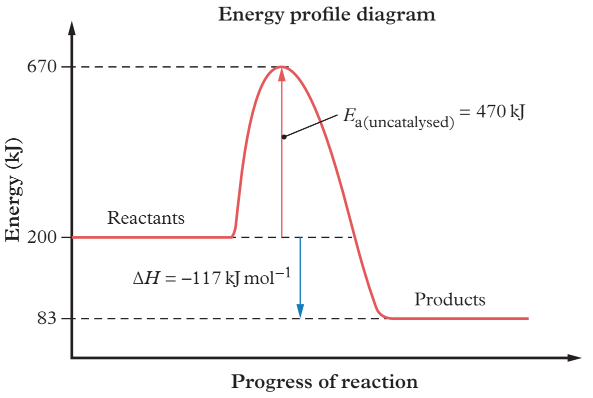
Interpreting an energy profile diagram

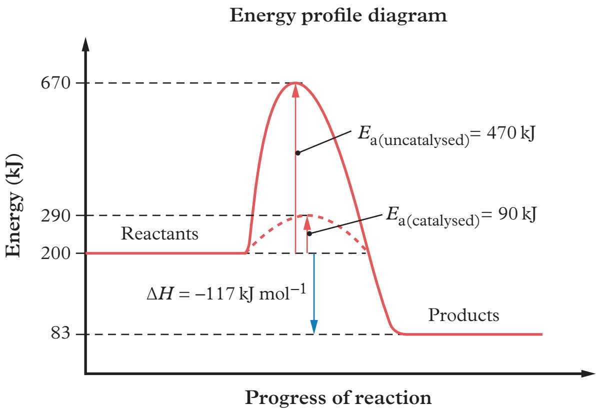
Worked example 16.4B: Watch a video that shows how to solve this problem.

Ethane reacts with chlorine gas to form chloroethane and hydrogen chloride gas according to the equation:

$$C_2H_6(g) + Cl_2(g) \rightarrow C_2H_5Cl(g) + HCl(g) \qquad \Delta H = -117\,\text{kJ mol}^{-1}$$

The reactants have an energy of 200 kJ and the activation energy is 470 kJ. The addition of a catalyst provides an alternative pathway that reduces the activation energy to 90 kJ. **Construct** an energy profile for the catalysed and uncatalysed reaction pathways. (7 marks)

Think	Do
Step 1: Look at the cognitive verb and mark allocation to determine what the question is asking you to do.	“Construct” means to display information in a diagrammatical or logical form. We need to use the information given to construct an energy profile diagram. The question is worth 7 marks, so we must represent the information correctly.
Step 2: Construct an energy profile diagram for the uncatalysed reaction. Refer to Worked example 16.2A for a refresher.	 (1 mark for labelled axes, 1 mark for correct graph shape, 1 mark for reactants and products indicated with correct energy, 1 mark for ΔH label, 1 mark for $E_{a(\text{uncatalysed})}$ labels)

Think	Do
Step 3: Draw a new curve on the same diagram to show the catalysed reaction. Remember that the enthalpies do not change, only the activation energy.	 (1 mark for correct graph shape, 1 mark for $E_{a(catalysed)}$ labels)

Your turn

In photosynthesis, carbon dioxide reacts with water in the presence of visible light to form glucose and oxygen according to the following equation:

$$6\,CO_2(g) + 6\,H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g) \qquad \Delta H = +2{,}816\,\text{kJ mol}^{-1}$$

The reactants have an energy of 120 kJ and the activation energy is 3,240 kJ. In the presence of the catalyst chlorophill, the activation energy is 190 kJ. **Construct** an energy profile diagram for the catalysed and uncatalysed reactions. (7 marks)

Study tip

Remember that catalysts provide an alternative pathways with a lower E_a and therefore result in a higher proportion of successful collisions. They do not lower the energy of reactants.

Real-world chemistry

Enzymes as catalysts

Enzymes are proteins (a type of biological molecule) that are produced from a code provided by the DNA of organisms. They catalyse biochemical reactions and are so specialised that each enzyme generally catalyses one specific chemical reaction. For example, amylase, an enzyme found in saliva, catalyses the breakdown of most of the starch in foods into glucose molecules before they are sent to the stomach.

Enzymes work best at a certain temperature and pH range. For example, human enzymes tend to work best at body temperature (37°C). Pepsin, an enzyme that catalyses the breakdown of proteins in the stomach, operates best at low pH such as the pH of 1.5 of stomach acid.

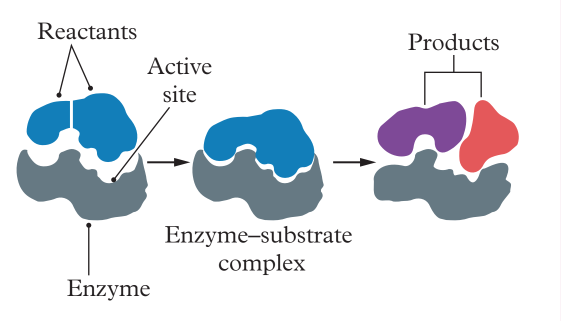


FIGURE 7 The enzyme-catalysed breakdown of a biological substrate occurs at the active site of the enzyme.

Enzymes act by interacting with a reactant called a substrate. The substrate enters the active site of the enzyme and forms intermolecular bonds with it (Figure 7). The formation of these intermolecular interactions stabilises the transition state for the reaction, providing a lower activation energy for the reaction and increasing the rate of reaction. Once the reaction is complete, the product is released from the enzyme to perform a biological function, leaving the enzyme unchanged.

Many enzymes catalyse the joining together of two or more substrate molecules through the formation of bonds. The way that the substrates interact with the enzyme ensures that the substrate molecules are aligned with one another in the right orientation to react correctly (Figure 8). The activation energy of the catalysed reaction is lowered, allowing the substrate molecules to bond together. The new product molecule is released, leaving the enzyme unchanged.

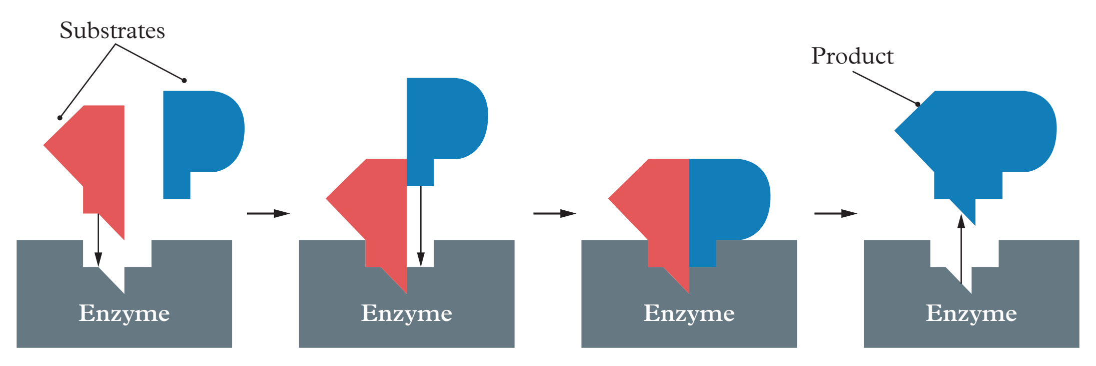


FIGURE 8 The enzyme-catalysed bonding of two biological substrates occurs at the active site of the enzyme.

Apply your understanding

1 **Describe** the types of attractive forces involved in a catalysis reaction. (2 marks)
2 **Summarise** the importance of an active site in an enzyme. (3 marks)
3 **Use** an example to **predict** the effect on bodily functions if catalysts did not exist. (2 marks)

Check your learning 16.4

Check your learning 16.4: Complete these questions online or in your workbook.

Retrieval and comprehension

1 **Explain** the effect a catalyst has on a chemical reaction. (2 marks)
2 **Construct** a Maxwell–Boltzmann diagram representing particles in a reaction at 30°C and 60°C. **Use** labels to **identify** the activation energy of the catalysed and uncatalysed reaction. (4 marks)

Analytical processes

3 **Contrast** a homogeneous and a heterogeneous catalyst. (1 mark)
4 The dissociation of hydrogen peroxide (H_2O_2), using a potassium iodide catalyst, is an exothermic process that forms water and oxygen gas. **Apply** your understanding to **construct** a fully labelled energy profile diagram of the dissociation of H_2O_2, including both the catalysed and uncatalysed pathways. (7 marks)

Knowledge utilisation

5 **Investigate** the production of ethanol by the fermentation of sugar and by the reaction of water with ethene.
   a **Construct** a balanced chemical equation for each chemical reaction. (2 marks)
   b **Identify** the catalyst involved in each chemical reaction. (2 marks)
   c **Comment** on the environmental impact of each reaction. (4 marks)
6 **Investigate** metal nanoparticles as catalysts and provide a summary of two examples of their use. (4 marks)

MODULE 16

Lesson 16.5

Review: Rates of chemical reactions

Summary

16.1
- Collision theory states that reactants must collide in the correct orientation and with sufficient energy to result in a successful chemical reaction.
- Temperature, pressure, concentration and surface area influence the rate of a chemical reaction.

16.2
- Practical: Investigating the rate of chemical reactions.

16.3
- Energy profile diagrams are used to show the change in energy as the reaction progresses over time
- If energy is released in a reaction, it is exothermic. If energy is absorbed in a reaction, it is endothermic.
- The stronger the bonds broken into a reaction, the higher the activation energy and the slower the reaction.

16.4
- Catalysts increase the rate of a reaction by providing an alternative pathway that requires a lower activation energy to form the same products.
- The effect of a catalyst can be graphically represented on a Maxwell -Boltzmann distribution and on an energy profile diagram.

Key formulas

Reaction enthalpy	$\Delta H = H_{\text{products}} - H_{\text{reactants}}$
Rate of reaction	$\text{rate of reaction} = \frac{\text{change in mass of reactant}}{\text{time}}$

Review questions 16.5A Multiple choice

Review questions: Complete these questions online or in your workbook.

(1 mark each)

1 Consider the chemical reaction:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Which of the following will not increase the rate of the reaction?

A Crush the magnesium into a powder
B Increase the concentration of oxygen
C Add water to make the reactants aqueous
D Increase the temperature by heating the magnesium

2 Consider the chemical reaction:

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Which of the following will not increase the rate of the reaction?

A Increase the volume of the solution mix
B Increase the concentration of NaCl
C Increase the temperature
D Increase the concentration of $AgNO_3$

3 Which of the following conclusions about the graph is correct?

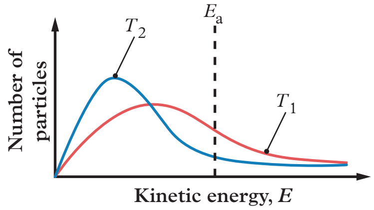


FIGURE 1 Maxwell–Boltzmann distribution

A At T_2, fewer particles have energy greater than E_a; therefore, there will be a slower reaction at this temperature, compared to T_1.
B At T_1, fewer particles have energy greater than E_a; therefore, there will be a faster reaction at this temperature, compared to T_2.
C At T_1, more particles have energy greater than E_a; therefore, there will be a slower reaction at this temperature, compared to T_2.
D At T_2, more particles have energy greater than E_a; therefore, there will be a faster reaction at this temperature, compared to T_1.

4 In an endothermic reaction
A the enthalpy value is positive.
B the reaction loses energy to the environment.
C the energy required to break bonds is smaller than the energy released when new bonds form.
D a catalyst will increase the activation energy.

5 In a catalysed chemical reaction
A the difference in enthalpy between products and reactants (ΔH) changes.
B an alternative reaction pathway with lower activation energy is provided.
C the kinetic energy of particles and therefore temperature increases.
D an alternative reaction pathway with higher activation energy is provided.

6 When an enzyme acts as a catalyst for chemical reactions, it
A provides a lower activation energy by stabilising the transition state.
B provides a higher activation energy by stabilising the transition state.
C provides a lower activation energy by destabilising the transition state.
D provides a higher activation energy by destabilising the transition state.

7 A catalyst increases the rate of a chemical reaction by
A increasing the frequency of collisions.
B decreasing the frequency of collisions.
C providing a lower activation energy.
D providing a higher activation energy.

8 An increase in the reactant concentration will
A increase the frequency of collisions and increase the number of successful collisions.
B increase the frequency of collisions and decrease the number of successful collisions.
C decrease the frequency of collisions and increase the number of successful collisions.
D decrease the frequency of collisions and decrease the number of successful collisions.

9 If pressure is doubled in a closed system
A volume is halved, causing the concentration to double and the frequency of collisions to increase.
B volume is doubled, causing the concentration to halve and the frequency of collisions to decrease.
C volume is constant, but the concentration increases, which increases the frequency of collisions.
D volume is constant, but the concentration decreases, which decreases the frequency of collisions.

10 Which of the following conclusions about the graph is correct?

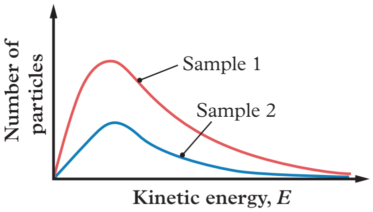


FIGURE 2 Maxwell–Boltzmann distribution

A Sample 2 is at a higher temperature than Sample 1 because there are fewer particles that have the average kinetic energy.
B Sample 2 is at a lower concentration than Sample 1 because it has fewer particles but the same average kinetic energy.
C Sample 1 is at a higher temperature than Sample 2 because there are more particles that have the average kinetic energy.
D Sample 1 is at a lower concentration than Sample 2 because it has a greater number of particles but the same average kinetic energy.

Short response 16.5B Short response

Review questions: Complete these questions online or in your workbook.

Retrieval and comprehension

11 **Define** an enzyme. (1 mark)

12 **Describe** how temperature alters the rate of a reaction. (3 marks)

13 A cold pack contains solid ammonium nitrate (NH_4NO_3) with water in separate pouches. When the pouches are broken, the water dissolves the ammonium nitrate in an endothermic process. Ammonium nitrate has an enthalpy value of 621 kJ mol^{-1}, while NH_4^+ and NO_3^- have a combined enthalpy value of 647 kJ mol^{-1}. The ΔH value for the endothermic reaction is +26 kJ mol^{-1}, and the activation energy for the reverse, exothermic reaction is 64 kJ mol^{-1}. **Sketch** a fully labelled energy profile diagram for the forward reaction, including E_a and ΔH values. (5 marks)

14 **Describe** the difference between activation energy of an uncatalysed reaction and the activation energy of a catalysed reaction. (1 mark)

15 A beaker containing a clear solution of a slow reacting precipitation reaction is placed on a piece of paper which has a black cross written on it.

- a **Describe** a method which could be used to measure the rate of reaction. (3 marks)
- b **Explain** two ways that the rate of reaction be increased. (4 marks)

16 Hydrogen peroxide decomposes into oxygen gas and water, but the process is very slow.

- a **Describe** a method that could be used to measure the rate of reaction. (3 marks)
- b **Explain** three ways that the rate of reaction could be increased. (6 marks)
- c Hydrogen peroxide solutions are normally stored in a refrigerator. **Consider** why this would be done. (2 marks)

17 **Explain** why the introduction of a catalyst does not increase the frequency of collisions. (1 mark)

18 A chemical reaction between 20.0 mL of 1 M sodium carbonate and 5.0 mL of 2.0 M hydrochloric acid produced the volume vs time graph.

$$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

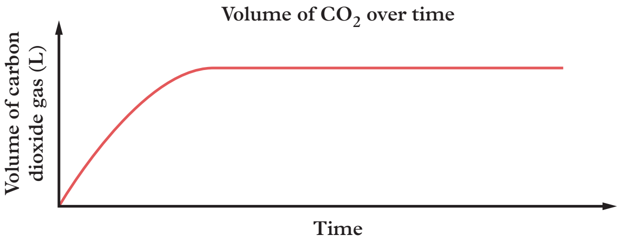


FIGURE 3 Change in volume of CO_2 over time

Sketch the resulting line on the graph when

- a the temperature of the reaction is increased (1 mark)
- b the volume of sodium carbonate is increased by 10 mL (1 mark)
- c the temperature of the reaction flask is decreased (1 mark)

19 Two experiments are conducted to investigate the factors affecting the rate of a chemical reaction between magnesium metal and hydrochloric acid:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

In the first experiment, a 6 cm strip of magnesium ribbon, with a mass of 0.046 g, is reacted with 3 mL of 1 M HCl. In the second experiment, a cubic piece of magnesium, measuring 0.3 × 0.3 × 0.3 cm, with a mass of 0.046 g, is placed in 3 mL of 2 M HCl.

- a **Identify** how the rate of reaction could be measured in these experiments. (2 marks)
- b **Identify** what factors are being investigated in this reaction. (2 marks)
- c **Explain** what the major flaw is in this experiment. (2 marks)
- d **Explain** whether it is possible to determine which reaction occurs faster. (2 marks)

Analytical processes

20 A chemistry student made the claim that “All collisions result in a chemical reaction. When the number of collisions is increased, the rate of reaction increases.” **Critique** this claim. (3 marks)

Knowledge utilisation

21 **Consider** the energy profile diagram.

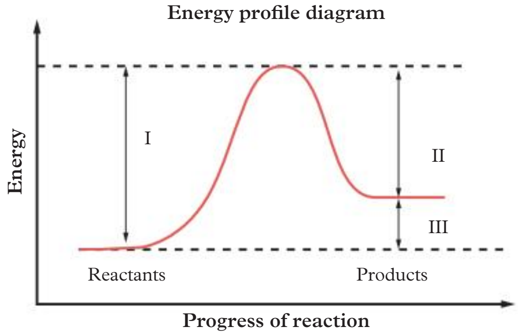


FIGURE 4 Energy profile diagram

a **Deduce** whether the reaction is endothermic or exothermic. **Justify** your answer. (2 marks)

b **Determine** which of the three quantities (I, II or III) is affected by the addition of a catalyst. (1 mark)

c **Determine** which of the three quantities is affected when the reaction is switched between endothermic/exothermic and endothermic/exothermic. (2 marks)

d **Consider** what would happen if the reaction was reversed. **Sketch** an energy profile diagram to **demonstrate** the effect. (3 marks)

Data drill

Rate of reaction – volume versus time

A student performed an experiment to calculate the rate of reaction between magnesium and hydrochloric acid. They set up the experiment as shown in Figure 1 and then used the results to create a volume versus time graph (Figure 2).

Apply understanding

1 **Use** the graph in Figure 2 to **calculate** the rate of reaction (in $\text{mL}\,\text{min}^{-1}$) between 1 and 2 minutes, 4 and 5 minutes, and 9 and 10 minutes. (3 marks)

Analyse data

2 **Identify** the time at which the end point of the reaction is reached. (1 mark)

3 **Contrast** Figure 2 with the expected shape of the graph for change in mass of magnesium over time. (2 marks)

Evaluate evidence

4 **Draw a conclusion** about the rate of reaction of magnesium metal with hydrochloric acid. (2 marks)

5 **Extrapolate**, by drawing on the graph, to **predict** the volume of gas that has been produced at 20 minutes. (2 marks)

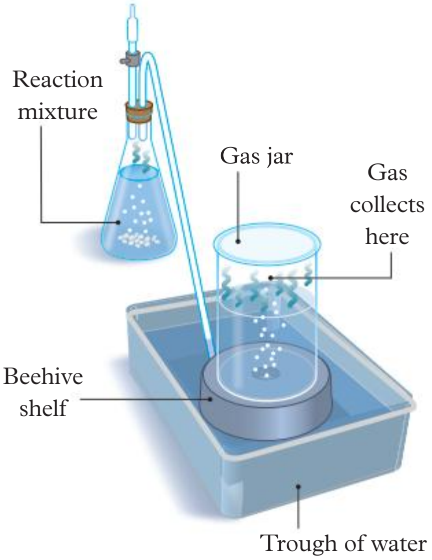


FIGURE 1 Experimental set-up for measuring volume of gas produced

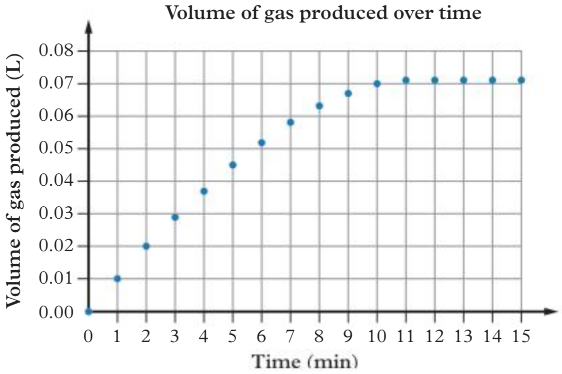


FIGURE 2 Volume of gas formed over time

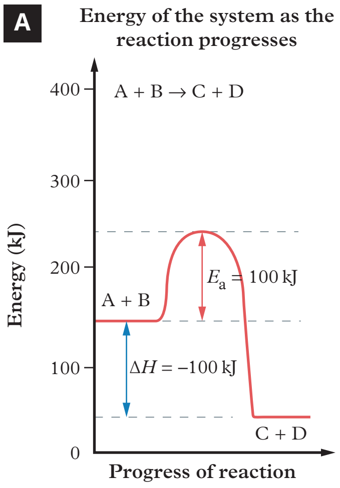
Module 16 checklist: Rates of chemical reactions

Quizlet: Revise key terms online to test your understanding

UNIT 2 Topic 3 review

Multiple choice

Use the following information to answer questions 1 and 2.



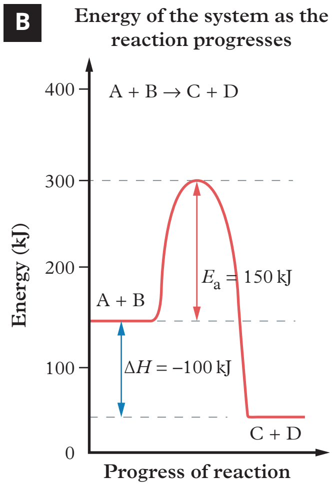


FIGURE 1 Energy profile diagrams for two reactions: (A) and (B)

(1 mark each)

1 Of the two reactions

- **A** A has faster rate of reaction due to its decreased activation energy.
- **B** A has a faster rate of reaction due to its increased activation energy.
- **C** B has a faster rate of reaction due to its increased activation energy.
- **D** B has a faster rate of reaction due to its decreased activation energy.

2 A student makes the following claims about the enthalpy of the reaction shown in the energy profile diagram.

- **I** The greater the bond enthalpy of the reactants, the greater the activation energy.
- **II** The greater the bond enthalpy of the products, the higher the enthalpy change of the reaction.
- **III** The addition of a catalyst provides an alternative pathway for the reaction in which less energy is required to break reactant bonds.

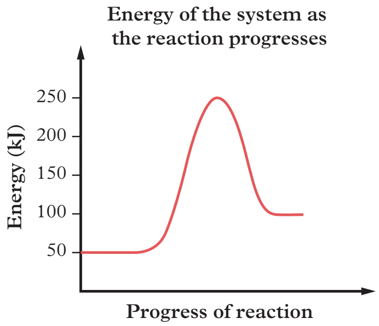


FIGURE 2 Energy profile diagram

Which of the student's claims are correct?

- **A** I only
- **B** III only
- **C** I and II
- **D** I, II and III

3 The same catalyst is used in both reactions. In both reactions, the catalyst lowers the activation energy by 90 kJ. What of the following is true?

A The enthalpies of both catalysed reactions are lower than the uncatalysed reactions.

B The rate of the uncatalysed reaction of B is faster than the catalysed reaction of A.

C The rate of the catalysed reaction of B is faster than the uncatalysed reaction of A.

D The enthalpies of both catalysed reactions are higher than the uncatalysed reactions.

Use the following information to answer questions 4 to 7.

Excess hydrochloric acid was reacted with sodium carbonate and the volume of gas generated was measured over time. The results are shown in Figure 3.

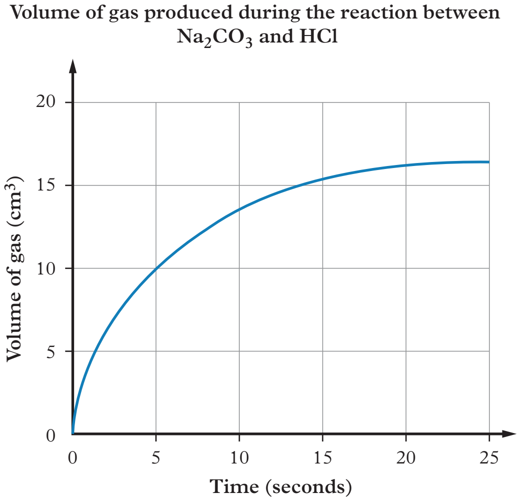


FIGURE 3 Volume of gas generated over time

4 What is the rate of reaction between time 0 and 5 seconds (in $cm^3 s^{-1}$)?

A 5

B 2

C 10

D 0.5

5 What is the rate of reaction between time 5 and 14 seconds ($cm^3 s^{-1}$)?

A 1

B 1.8

C 1.07

D 0.56

6 When is the rate of reaction at its greatest?

A Between 0 and 1 second

B Between 0 and 5 seconds

C Between 0 and 20 seconds

D Between 0 and 25 seconds

7 Why does the volume of gas plateau after 21 seconds?

A The reaction is endothermic and has stopped because it is too cold.

B All of the hydrochloric acid has reacted so no more gas can be produced.

C All of the sodium carbonate has reacted so no more gas can be produced.

D The concentration of HCl has decreased to a point where it is too low to react.

Use the following information to answer questions 8 and 9.

Two chemical species were reacted (Reaction A) and the rate of reaction was determined by measuring the mass of gaseous product lost over time. This is shown in Figure 4. The reaction was repeated with one change made and the rate of reaction was recorded on the same graph (Reaction B).

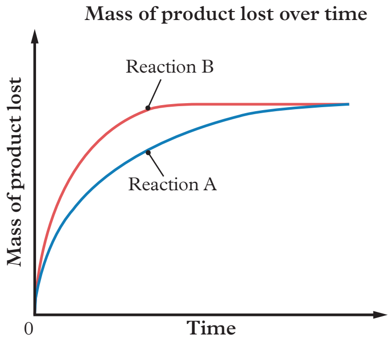


FIGURE 4 Mass of product lost over time

8 Which is true of the reactions?

A The steeper the line, the faster the rate of reaction.

B Reaction B occurs four times as rapidly as reaction A.

C The rate of the reaction is measured in mass of reactant lost per unit of time.

D The final amount of product is greater for reaction B than for reaction A.

9 What change has been made after Reaction A to result in the data obtained for Reaction B?

A Decrease temperature

B Increase volume of the excess reactant

C Increase surface area of the limiting reactant

D Decrease concentration of the limiting reactant

Use the following information to answer questions 10 and 11.

The hydrogen gas generated from a reaction between zinc and hydrochloric acid can be measured using two methods, A and B (Figure 5). In both cases, a stopwatch is used and changes are measured.

In method A, a glass jar is used to collect the gas and the time taken to fill the jar is recorded. In method B, the reaction takes place on top of a balance and the mass of the gas lost is recorded every 5 seconds.

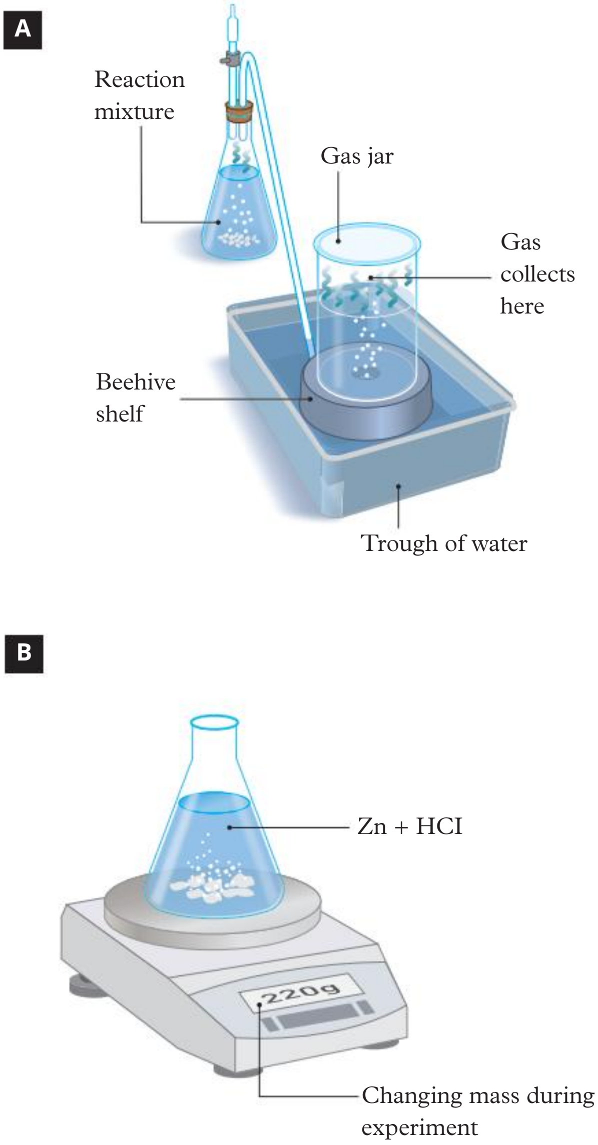


FIGURE 5 Hydrogen gas generated from two methods: (A) and (B)

10 Assuming that no other measurements were taken, which of the following is true?

A Method B is both more accurate and precise.

B Method A is both more accurate and precise.

C Method A is more accurate but B is more precise.

D Method B is more accurate but A is more precise.

11 Which statement is correct about the reliability and validity of the methods?

A Method A is valid but B is reliable.

B Method B is valid but A is reliable.

C Method A is both valid and reliable.

D Method B is both valid and reliable.

12 The Maxwell–Boltzmann distribution in Figure 6 represents the same reaction performed at two different temperatures.

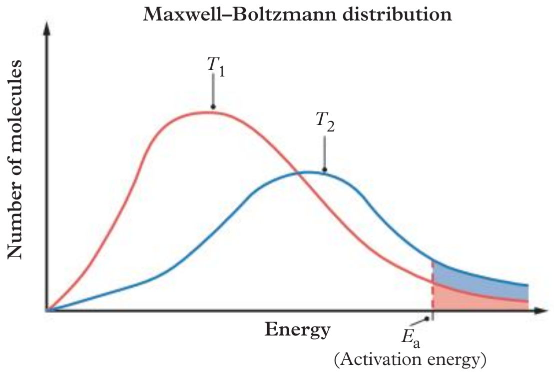


FIGURE 6 The Maxwell–Boltzmann distribution

Which of the following statements is correct?

A T_2 is greater than T_1. At T_2, there are more particles with energy greater than the activation energy.

B T_2 is greater than T_1. At T_1, there are more particles with energy greater than the activation energy.

C T_1 is greater than T_2. At T_2, there are more particles with energy greater than the activation energy.

D T_1 is greater than T_2. At T_1, there are more particles with energy greater than the activation energy.

13 The Maxwell–Boltzmann distribution in Figure 7 represents the same reaction performed in two different conditions. The blue curve is the original reaction, and the red curve is the reaction performed under the new conditions.

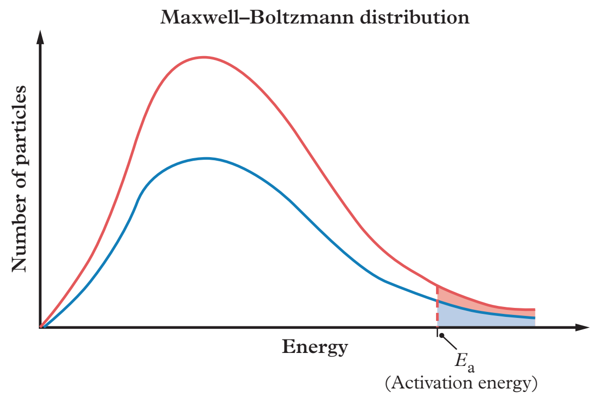


FIGURE 7 Maxwell–Boltzmann distribution

What change is represented by the red line?

A The addition of a catalyst
B A decrease in temperature
C An increase in temperature
D An increase in concentration

14 In the following reaction, which factor will have the smallest impact on increasing the rate of the chemical reaction?

$$H_2(g) + 3N_2(g) \rightarrow 2NH_3(g)$$

A The addition of a catalyst
B An increase in temperature
C An increase in surface area
D A decrease in the volume of the reaction system

15 In the following reaction, which factor will have the smallest impact on increasing the rate of the chemical reaction?

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

A An increase in temperature
B An increase in surface area of the zinc
C An increase in concentration of the HCl
D A decrease in the volume of the reaction system

Short response

16 In a reaction between H_2 and Cl_2, HCl is formed.

a Assuming that the activation energy of this reaction is 800 kJ mol^{-1}, **use** the bond enthalpies of the reactants and product to **sketch** an energy profile diagram. (8 marks)

b **Calculate** the enthalpy of the reaction and **identify** it on the energy profile by adding a label. (2 marks)

c When a catalyst is added, the reaction requires only 750 kJ to occur. **Sketch** the curve for the catalysed reaction pathway on the same energy profile and **explain** the effect that a catalyst has on the rate of a chemical reaction. (3 marks)

17 **Describe** what happens to the shape of a Maxwell–Boltzmann distribution curve when:

a temperature is increased (2 marks)
b temperature is decreased (2 marks)
c the concentration of the limiting reactant is doubled (2 marks)
d the volume of a closed gaseous system is doubled (2 marks)
e a catalyst is added. (2 marks)

18 **Consider** the following reaction between methane and chlorine to form chloromethane:

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$$

Predict the effect of each change below on the rate of production of chloromethane.

a Increasing temperature (at a constant volume) (3 marks)
b Increasing pressure (at a constant temperature) (3 marks)
c Adding a catalyst (2 marks)

19 Calcium carbonate and excess sulfuric acid react to form gaseous carbon dioxide according to the following equation:

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$$

In a reaction, 1.00 g of small $CaCO_3$ chips are reacted with 50.0 mL of 2.00 M H_2SO_4. The volume of the gas lost is recorded every 10 seconds and a graph of the results is plotted as shown in Figure 8. The dashed blue line represents this original reaction.

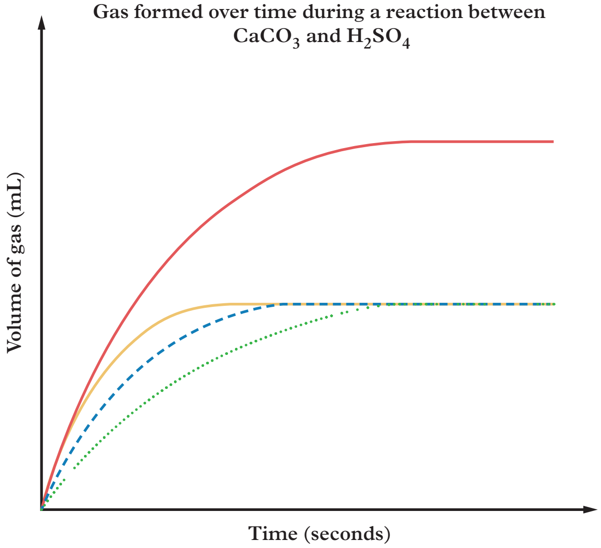


FIGURE 8 Gas formed over time during a reaction

- **a** **Determine** the limiting reactant. (5 marks)
- **b** **Determine** the theoretical volume of CO_2 produced at STP. (2 marks)

The experiment is repeated by changing the surface area of the chips.

- **c** **Identify** the curve on the graph that represents the rate of reaction when powdered $CaCO_3$ is used. **Justify** your selection. (3 marks)
- **d** An additional change is made to the reaction where the powdered $CaCO_3$ is reacted with 2.00 M H_2SO_4. This results in the red curve on the graph. **Deduce** what this change was and **justify** your decision, then **explain** how it affects the rate of the reaction. (4 marks)

20 Ethanol is formed from the fermentation of glucose, producing carbon dioxide as a waste product according to the following equation:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

- **a** **Propose** one change that can be made to increase the rate of ethanol production from glucose. **Justify** your choice. (2 marks)
- **b** **Describe** how the rate of the fermentation reaction can be measured. (2 marks)
- **c** **Explain** why the fermentation process often requires yeast. (3 marks)

21 A 1.5 g sample of iron granules is reacted with 50.00 mL of 4.00 M nitric acid according to the following equation:

$$Fe(s) + 2HNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + H_2(g)$$

The volume of hydrogen gas generated was measured every 2 seconds and Figure 9 was drawn.

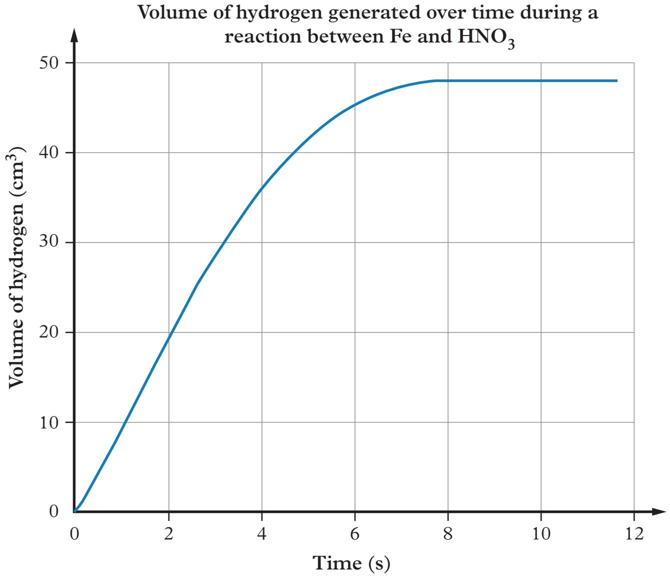


FIGURE 9 Volume of hydrogen generated over time

- **a** Assuming the reaction was completed at STP (V_m = 22.7 L mol^{-1}), **determine** the mass of iron that reacted to generate the graph. (3 marks)
- **b** **Identify** when the rate of reaction was at its greatest, using a calculation to support your answer. (2 marks)

- **c** The iron granules were crushed into a powder and the reaction was repeated.
  - **i** **Sketch** and label a curve on the same graph to show how the volume of gas changes over time. (1 mark)
  - **ii** **Explain** the effect of this change on the rate of the reaction. (3 marks)
- **d** The concentration of the nitric acid was halved and the experiment repeated.
  - **i** **Sketch** and label a curve on the same graph to show how the volume of gas changes over time. (1 mark)
  - **ii** **Explain** the effect of this change on the rate of the reaction. (3 marks)

22 Sulfur dioxide reacts with oxygen to form sulfur trioxide, according to the following equation:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

The Maxwell–Boltzmann distribution of the SO_3 particles in the container at a particular temperature is shown.

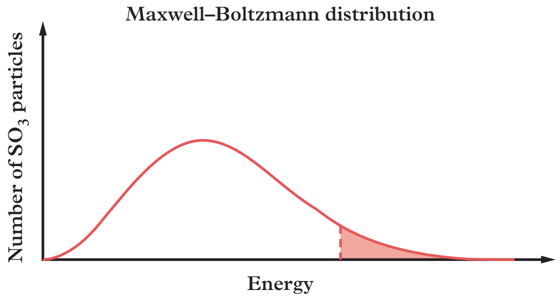


FIGURE 10 Maxwell–Boltzmann distribution

- **a** **Identify** what the area under the curve represents. (1 mark)
- **b** **Sketch** and label a curve on the same Maxwell–Boltzmann distribution to show the effect of a significantly lower temperature. (1 mark)
- **c** **Explain** the effect of this decrease in temperature on the rate of a reaction, with reference to the curve you constructed in part **b**. (3 marks)
- **d** The vertical line on the graph represents the activation energy of the reaction. **Sketch** the effect of a catalyst on the graph and **explain** how it affects the rate of a reaction. (3 marks)
- **e** **Explain** the effect of halving the volume of the container on the rate of the reaction. (3 marks)

23 Propane is used as a fuel for barbeques. Each mole of propane produces 2,220 kJ of energy.

- **a** **Determine** the thermochemical equation for the combustion of propane. (2 marks)
- **b** **Calculate** the amount of energy released if a 40.0 L tank of propane, at 100 kPa and 18°C, is completely combusted. (4 marks)
- **c** **Sketch** the energy profile for the combustion of propane. (1 mark)
- **d** In the incomplete combustion of propane, oxygen is limiting and carbon monoxide is formed rather than carbon dioxide. **Predict** and **sketch** the effect of incomplete combustion on the energy profile for the combustion of propane. (2 marks)
- **e** **Justify** the difference in the enthalpy of reaction for both the complete and incomplete combustion of propane. (2 marks)

24 A reaction between 100.00 mL of 0.0500 M sodium thiosulfate and 5.00 mL of 2.00 M hydrochloric acid was set up to determine the effect of temperature on the rate of a reaction.

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$$

The rate of the reaction is determined by sitting four flasks, at different temperatures, on a piece of white paper which has a black cross written on it. As both reactants are poured into the flasks, the solid sulfur precipitates and the time taken for the cross to completely disappear is recorded.

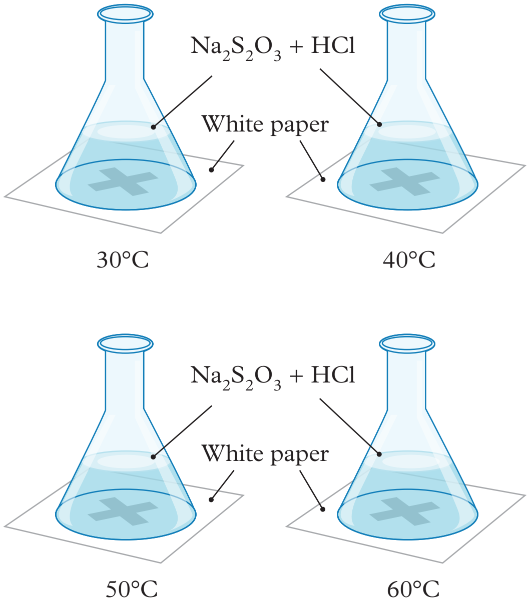


FIGURE 11 Two reactions

a **Identify** the independent variable. (1 mark)

b **Identify** the dependent variable. (1 mark)

c **Generate** a hypothesis for the reaction. (2 marks)

d **Justify** whether the experiment is qualitative or quantitative. (2 marks)

The following results were obtained:

TABLE 1 Experiment results

Temperature (°C)	30	40	50	60
Time taken for the cross to disappear (s)	51	42	34	22

e **Explain** what the trend in the data demonstrates about the rate of a reaction. (3 marks)

f **Evaluate** whether the experiment is valid. **Justify** your response. (2 marks)

g **Propose** how the experiment could be improved to measure the rate of the reaction. **Justify** your response. (3 marks)

25 Ethanol is used as a biofuel in cars. It is a cleaner, greener alternative to non-renewable fuels such as octane (C_8H_{18}). The ethanol is then combusted with a molar enthalpy of $-1{,}370\,\text{kJ}\,\text{mol}^{-1}$, which is lower than octane's $-5{,}470\,\text{kJ}\,\text{mol}^{-1}$.

a **Sketch** the energy profiles for both the combustion of ethanol and octane on separate graphs, labelling activation energy and enthalpy. (2 marks)

b **Explain** why ethanol has a lower molar enthalpy of combustion than octane. (3 marks)

c **Predict** the relative activation energies of octane and ethanol. **Justify** your response. (2 marks)

d **Deduce** whether ethanol or octane would have a faster rate of reaction. **Justify** your response. (2 marks)

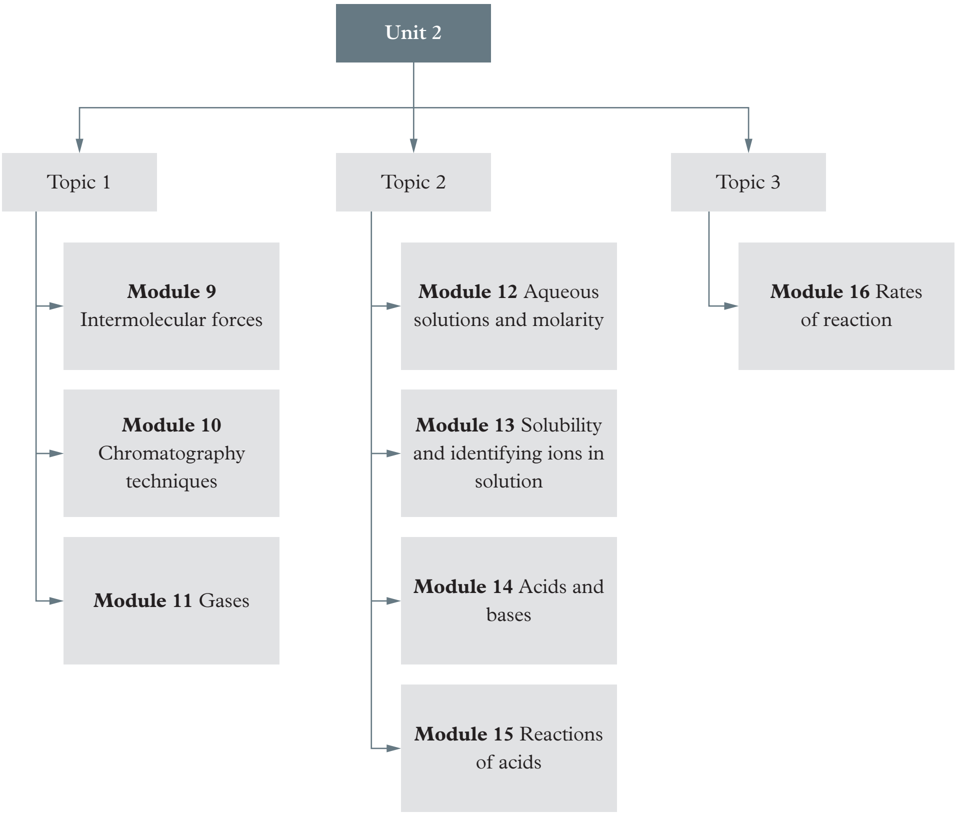
e **Propose** a reason to explain why people who live in colder climates cannot use ethanol as a fuel and instead must use octane. (3 marks)

TOTAL MARKS
/128

UNIT 2 Review

Part A – Revisit and revise

Part A of the Unit review asks you to reflect on your learning and identify areas in which you need more work.



Part B – Exam essentials

Now that you've completed your revision for Unit 2, it's time to learn and practise some of the skills you'll need to answer exam questions like a pro! To help you, our expert authors have created the following advice and tips to help you maximise your results on the end-of-year examination.

Exam tip 1: Use data to justify explanations

Where data is given in the question and conclusions must be drawn using it, you MUST refer directly to the data in your response. This helps you to clearly justify the reasons for your response.

If you think that there is not enough information or data to answer the question, there may be information that you can find in the Chemistry formula and data book.

See it in action

Read the real exam question below and see how the tip has made a difference between a response that has scored full marks and a response where marks have been lost.

QUESTION 27 (5 marks)

Five colourless 0.1 M solutions of NH_3, HCl, KOH, H_2SO_4 and CH_3CH_2COOH have lost their labels. The substances are randomly relabelled A, B, C, D and E. The conductivity of each solution and the colour of the solution when phenol red was added are shown.

Solution	Conductivity (S/m)	Colour with phenol red
A	4.1	Yellow
B	0.14	Red
C	0.08	Yellow
D	6.7	Yellow
E	4.9	Red

Identify the five solutions. Explain your reasoning.

Source: QCAA 2022 Chemistry Paper 1

Complete response

Realises that these are missing information; obtains the relevant information from the data book about the pH range of colour change for phenol red [1 mark]

Phenol red is yellow when pH < 6.8 and red when pH > 8.4. For the acids, C has a low conductivity so it must partially ionise in water and have a [H^+]. C must be propanoic acid since it is a weak acid. Solution D has a higher conductivity than solution A, so D must be sulfuric acid and A must be hydrochloric acid, so H_2SO_4 is diprotic and can donate 2 H^+ ions (HCl is monoprotic and can only donate 1). D will produce a higher concentration of ions in solution than C. Solution E is KOH since it is a strong based that completely dissociates in water and produces a higher concentration of OH^- ions. It will have a higher conductivity. B has a lower conductivity so it is ammonia. It only partially dissociates, giving a lower concentration of ions.

Uses conductivity data to identify the relative strength of the acids [1 mark]

Identifies the diprotic acid is more conductive than monoprotic acid [1 mark]

Uses conductivity data to identify the relative strength of the bases [1 mark]

Identifies all five solutions [1 mark]

Incomplete response

Has probably realised that this is missing information and has used the relevant information from the data book about the pH range of colour change for phenol red, but does not reference it as justification [0 marks]

Solutions A, C and D are acids and solutions B and E are bases.

Solution C is weak and is therefore propanoic acid.

Solution D is sulfuric acid due to its higher conductivity and solution A is HCl.

Solution E is KOH as it has a higher conductivity.

Solution B is ammonia.

Has identified solution C as the weak acid but has not provided any justification [0 marks]

Has referenced the data but has still not provided sufficient justification [0 marks]

Has referenced the data but has still not provided sufficient justification [0 marks]

Appears to have obtained this by process of elimination, without support from the data [0 marks]

Identifies all five solutions [1 mark]

Think like an assessor

To maximise your marks on an exam, it can help to think like a QCAA assessor. Consider how many marks each question is worth and what information the assessor is looking for.

A student has given the following response in a practice exam. Imagine you are a QCAA assessor and use the marking guide below to mark the response.

QUESTION 2 (4 marks)

A student is analysing the metal ions in four solutions in the laboratory using precipitation reactions. Each solution contains one of the following ions: Cu^{2+}, Pb^{2+}, Ag^{+} and Zn^{2+}. The results of the precipitation reacts with various anions are in the table. Assume no other reactions have occurred. Note: ppte = precipitate.

Unknown	**CO_3^{2-}**	**Cl^-**	**OH^-**	**I^-**	**SO_4^{2-}**
A	ppte	ppte	ppte	ppte	ppte
B	ppte	no ppte	ppte	no ppte	no ppte
C	not tested	no ppte	ppte	ppte	no ppte
D	ppte	no ppte	ppte	ppte	ppte

Determine the identity of the four metal ions. Explain your reasoning.

Ag^+ must be solution A. B and C both do not form a ppte with SO_4^{2-} so those two solutions must contain the Cu^{2+} and the Zn^{2+}. But B must have the Zn^{2+}. Solution D must contain Pb^{2+} because Pb^{2+} forms a ppte with most of the ions tested except for Cl^-.

Marking guide

Question 2

- identifies A as Ag^+ being the only one of the four ions to be insoluble/partly soluble with the anions used [1 mark]
- identifies B as Zn^{2+}, as it does not form a precipitate with I^- and the other metal ions do (or another valid reason that eliminates others) [1 mark]
- identifies D as Pb^{2+}, as it is forms a ppte with most ions except for Cl^- (or another valid reason that eliminates others) [1 mark]
- identifies Cu^{2+} as C as it does not form a ppte with Cl^- or SO_4^{2-}, so differs from the others (or another valid reason that eliminates others) [1 mark]

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved? Write your own response to the same question to receive full marks from a QCAA assessor.

Exam tip 2: Make sure quantities are in the correct units before completing calculations

The most common mistakes when using $PV = nRT$ are:

- not converting temperature to the Kelvin scale
- mismatching pressure and volume units. Since the value of R given by QCAA is 8.31 J K^{-1} mol^{-1}, if P is in kPa, V should be in L and if P is in Pa, V should be in m^3.

The most common mistakes when using $n = \frac{m}{M}$ and $c = \frac{n}{V}$ are:

- not converting masses to g before calculating
- not converting volume of solution to L before calculating.

See it in action

Read the exam-style question below and see how the tip has made a difference between a response that has scored full marks and a response where marks have been lost.

QUESTION 21 (3 marks)

Calculate the pH of a 0.1 M aqueous solution of $Ba(OH)_2$, assuming complete dissociation. Show your working.

pH = ________________ (to one decimal place)

Complete response

Has determined the balanced chemical equation

Correctly determines $[OH^-] = 0.2$ M [1 mark]

$\boxed{Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)}$

$[OH^-] = [Ba(OH)_2] \times 2 = 0.1 \times 2 = \boxed{0.2\ M}$

$10^{-14} = [H^+][OH^-]$

$[H^+] = \frac{10^{-14}}{[OH^-]} = \frac{10^{-14}}{0.2} = \boxed{5 \times 10^{-14}\ M}$

Correctly determines $[H^+] = 5 \times 10^{-14}$ M; can alternatively determine pOH [1 mark]

$pH = -\log_{10}[H^+] = -\log_{10}(5 \times 10^{-14}) = \boxed{13.3}$

Correctly determines pH = 13.3; can alternatively calculate pH using 14 – pOH [1 mark]

Incomplete response

$[H^+] = \boxed{0.1\ M}$ Has mistaken the concentration of the base for $[H^+]$ [0 marks]

$pH = -\log_{10}[H^+] = -\log_{10}(0.1) = \boxed{1.0}$ Arrives at a consequentially correct answer [1 mark]

Think like an assessor

To maximise your marks on an exam, it can help to think like a QCAA assessor. Consider how many marks each question is worth and what information the assessor is looking for.

A student has given the following response in a practice exam. Imagine you are a QCAA assessor and use the marking guide below to mark the response.

QUESTION 4 (3 marks)

A car tyre has a volume of 40.0 L and is at a temperature of 25°C. When it is fully pumped with air, the internal pressure is 3.43×10^5 Pa.

Calculate the mass of air in the tyre, if the average molar mass of air is 28.97 g mol^{-1}.

$PV = nRT$ so $n = \frac{PV}{RT}$

$n = \frac{3.43 \times 10^5 \times 40.0}{8.21 \times 298}$

$= 5.540$ mol

$n = \frac{m}{M}$ so $m = n \times M$

$n = 5.540 \times 28.97 = 160{,}493$ g

Marking guide

Question 4

- uses T in K and P in kPa in the rule $PV = nRT$ [1 mark]
- calculates n [1 mark]
- multiplies the calculated value of n and the given molar mass, to calculate the mass of air [1 mark]

Fix the response

Consider where you did and did not award marks in the above response. How could the response be improved? Write your own response to the same question to receive full marks from a QCAA assessor.

Exam tip 3: Know what is required for each cognitive verb

When cognitive verbs from the syllabus are used in exam questions, know what is required in the response. This is particularly important for seemingly similar cognitive verbs that are frequently confused. Examples of verbs that students confuse are:

- compare and contrast
- discuss and explain
- analyse, evaluate and justify.

See it in action

Read the exam-style question below and see how the tip has made a difference between a response that has scored full marks and a response where marks have been lost.

QUESTION 1 (3 marks)

Several variables can be altered to increase the rate of reaction.
Compare how increasing the concentration of reactants and increasing the temperature of the reaction mixture affect the rate of reaction, by considering collision theory.

Complete response

Identifies that both increasing concentration and increasing temperature allow an increase in the number of successful collisions [1 mark]

Similarity: Increase in the number of successful collisions.
Differences: Higher concentration increases the number of reactant particles, therefore increasing the frequency of collisions.
Higher temperature alters the Maxwell–Boltzmann distribution so that many more reactant particles possess the activation energy. Therefore, a much higher proportion of the collisions have sufficient energy to be successful and react.
To achieve a greater rate of reaction, both increasing temperature and increasing concentration of reactants can be used, but temperature increases have a much greater influence.

Highlights that a higher concentration increases the number of collisions [1 mark]

Refers to the increases in the number of particles with higher energies at higher temperatures (with reference to Maxwell–Boltzmann distribution) [1 mark]

Recognises the significance of the similarities and differences [1 mark]

Draws a link to the higher proportion of successful collisions due to particles possessing the required activation energy [1 mark]

Incomplete response

Fails to state that rate depends on the number of successful collisions for both variables [0 marks]

Both increasing temperature and increasing concentration increase the rate of reaction. An increase in concentration will mean there are more collisions, so the reaction rate increases. An increase in temperature increases the kinetic energy of the colliding particles.

Identifies that an increase in concentration results in a higher frequency of collisions, which increases the rate of reaction [1 mark]

Fails to recognise the significance of the similarities and differences [1 mark]

States that increasing the temperature increases the kinetic energy of colliding particles [1 mark] but fails to link to higher proportion of successful collisions with the required activation energy [0 marks]

Think like an assessor

To maximise your marks on an exam, it can help to think like a QCAA assessor. Consider how many marks each question is worth and what information the assessor is looking for.

A student has given the following response in a practice exam. Imagine you are a QCAA assessor and use the marking guide below to mark the response.

QUESTION 6 (4 marks)

Pentanal, $CH_3CH_2CH_2CH_2CHO$, and hexane, $CH_3CH_2CH_2CH_2CH_2CH_3$, have almost the same size and molar mass. Compare the intermolecular forces present in these two compounds and describe the significance of these forces on the boiling points.

Both of these substances consist of non-metallic elements bonded to each other. This means that they are held together by covalent bonds within their molecules. Between the molecules, there are intermolecular forces, which are much weaker than covalent bonds. Hexane is non-polar and pentanal is partly non-polar but the C=O bond is polar. So pentanal will have both weak dispersion forces and slightly stronger dipole-dipole attractions acting. Because both of these compounds are covalent molecules, their boiling points will be relatively low.

Marking guide

Question 6

- correctly identifies that hexane is non-polar and pentanal has a polar section and a non-polar section. [1 mark]
- correctly identifies that a similarity is that both hexane and pentanal will have relatively weak dispersion forces acting between the molecules [1 mark]
- correctly identifies the difference that pentanal will also have dipole–dipole attractions acting [1 mark]
- correctly identifies the significance that pentanal will have a higher boiling point than hexane has [1 mark]

Part C – Practice exam questions

Now it's time to put the tips and advice you've learned into practice while you complete these exam-style questions!

Multiple choice

(1 mark each)

1 What is the shape of sulfur difluoride (SF_2)?

A Bent

B Linear

C Pyramidal

D Tetrahedral

2 Ethanal is ________ and interacts with other ethanal molecules via ________________.

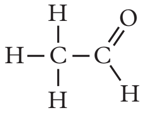


FIGURE 1 Ethanal

A polar; hydrogen bonding

B non-polar; dispersion forces

C polar; dipole–dipole attractions

D non-polar; dipole–dipole attractions

3 Which of the following lists the molecules from the smallest to the largest vapour pressure?

A H_2O, H_2S and CH_4

B H_2S, H_2O and CH_4

C CH_4, H_2O and H_2S

D CH_4, H_2S and H_2O

Use the following information to answer questions 4 and 5.

The following paper chromatography separation is produced when a leaf extract is applied to the origin of the paper and an organic solvent is used as the mobile phase.

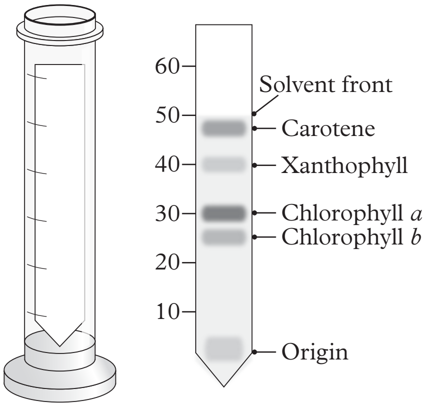


FIGURE 2 Paper chromatograph for leaf extract

4 Which of the following plays the largest role in the separation to produce the final chromatograph?

A The temperature of the solvent

B The length of the paper stationary phase

C The affinity of the pigments to the stationary phase

D The distance between the origin and the solvent front

5 What is the R_F of the chlorophyll *a* pigment?

A 0.8

B 0.6

C 0.5

D 0.3

6 In a balloon, as the temperature of the gas increases

A the balloon contracts, decreasing the pressure.

B the balloon volume remains constant and the pressure increases.

C the balloon expands, volume increases, to increase the pressure.

D the balloon expands, volume increases, to keep the pressure constant.

7 What is the temperature of a 40.0 L gas cylinder filled with 14.5 g of propane, C_3H_8, at 20.0 kPa?

A 6.64°C

B 19.86°C

C 292.86°C

D –266.36°C

8 What is the mass (in g) of butane, C_4H_{10}, gas contained within a 5.0 L tank at STP?

A 0.22

B 0.20

C 12.81

D 11.72

9 A 50.0 mL solution of 1.50 M NaOH has 5.00 g of NaOH added to it. What is the concentration of the new solution?

A 2.50 M

B 2.54 M

C 4.00 M

D 0.075 M

10 A 1.0 M nitric acid solution with a volume of 100 mL is diluted to 1,000 mL. Which statement about this dilution is incorrect?

A Although the concentration of H^+ ions in the solution decreases, the pH increases.

B The concentration of nitrate ions in the solution decreases when the dilution occurs.

C Nitric acid is a strong acid but it becomes a weak acid as the result of the dilution.

D The diluted solution has a lower concentration of particles and therefore a lower conductivity.

11 In which one of the following reactions is water acting as an acid?

A $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

B $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

C $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

D $H_2PO_4^-(aq) + H_2O(l) \rightarrow H_3PO_4(aq) + OH^-(aq)$

12 The equation that represents the precipitation reaction between barium nitrate and sodium sulfate is

A $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$.

B $NO_3^-(aq) + Na^+(aq) \rightarrow NaNO_3(s)$.

C $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3(aq) + BaSO_4(s)$.

D $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3(s) + BaSO_4(aq)$.

13 The ionic equation that represents the precipitation reaction between iron(III) chloride and sodium hydroxide is

A $Cl^-(aq) + Na^+(aq) \rightarrow NaCl(s)$.

B $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$.

C $FeCl_2(aq) + 2NaOH(aq) \rightarrow Fe(OH)_2(s) + 2NaCl(aq)$.

D $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$.

14 Consider the solubility curve in Figure 3.

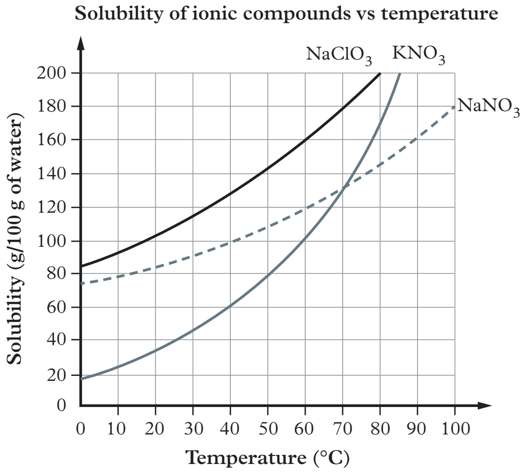


FIGURE 3 Solubility curve

100 g of a saturated solution of KNO_3 was cooled from 70°C to 40°C. What mass of KNO_3 would be expected to have crystallised from the solution?

A 56 g

B 60 g

C 70 g

D 80 g

15 When the pH of a solution changes from 4 to 2, the hydrogen ion concentration

A decreases by a factor of 2.

B increases by a factor of 2.

C increases by a factor of 100.

D decreases by a factor of 100.

16 What is the pH of a solution made by dissolving 1.36 g of barium hydroxide, $Ba(OH)_2$, in 100 mL of water?

A 11.9

B 12.2

C 12.9

D 13.2

17 Which one of the following combinations of aqueous solutions would not produce a precipitate when mixed?

A Barium nitrate and sodium sulfate

B Silver nitrate and sodium hydroxide

C Copper(II) nitrate and sodium chloride

D Copper(II) nitrate and sodium hydroxide

18 Which of the following chemical equations represents a metal carbonate reacting with an acid?

A $2NaOH(aq) + Li_2CO_3(aq) \rightarrow 2LiOH(aq) + Na_2CO_3(g)$

B $2LiOH(aq) + K_2CO_3(aq) \rightarrow 2KOH(aq) + Li_2CO_3(g)$

C $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

D $2H_3PO_4(aq) + 3CaCO_3(s) \rightarrow Ca_3(PO_4)_2(g) + 3CO_2(l) + 3H_2O(l)$

19 Which will not increase the rate of a reaction?

A Decreasing the volume of a gas cylinder

B Decreasing the pressure of a gas cylinder

C Increasing the temperature of the reaction

D Increasing the concentration of the limiting reactant

20 Consider the following Maxwell–Boltzmann distribution of a reaction.

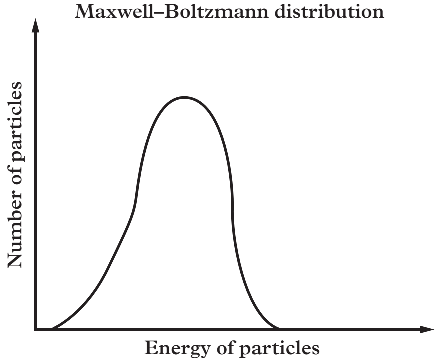
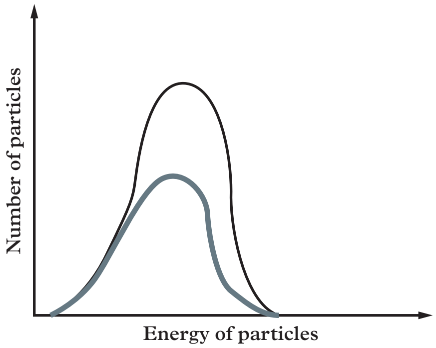


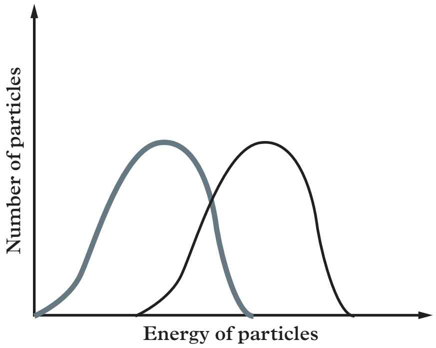
FIGURE 4 Maxwell–Boltzmann distribution

Which of the following distributions represents a system where the concentration was doubled?

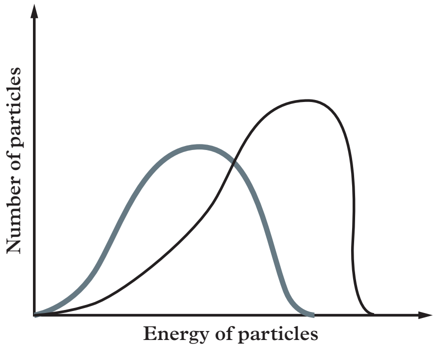
A



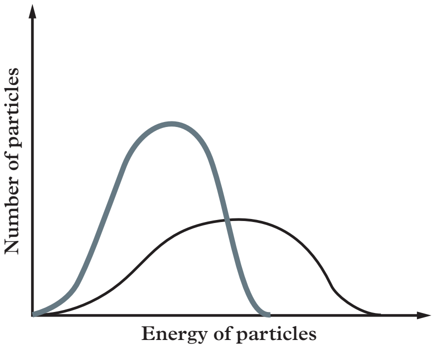
B



C



D



Short response

21 Explain why the shape of ammonia (NH_3) is not a trigonal planar. (3 marks)

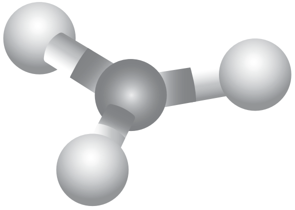


FIGURE 5 Ammonia

22 Fluorine is the most electronegative element on the periodic table. Use the molecular shape and polarity of HF to explain why it has a lower boiling point (19.5°C) than H_2O (100°C), but a higher boiling point than NH_3 (–33°C). (3 marks)

23 Explain why butane is a gas at room temperature, but octane is a liquid. (3 marks)

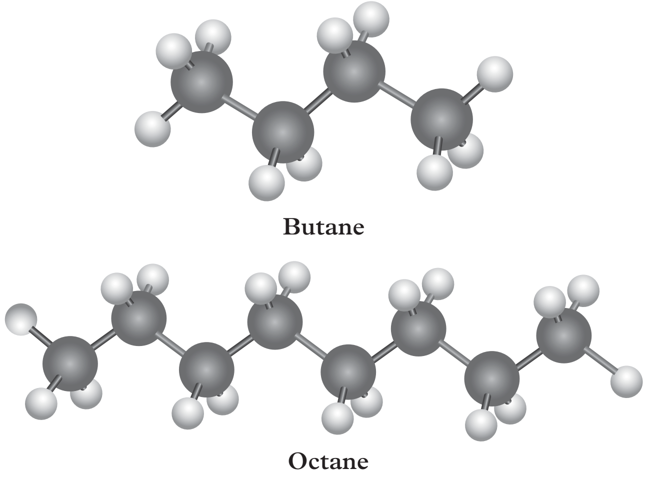


FIGURE 6 Butane and octane

24 Calculate the pressure of a 15.0 L cylinder filled with 9.50 g of gaseous methane, CH_4, at 18.0°C. (4 marks)

25 Explain how kinetic molecular theory can be used to predict the effect of increasing the temperature of a closed gaseous system on the pressure exerted on the walls of the container. (3 marks)

26 Explain why the walls of a gas tank must be very strong to increase the amount of time that divers can remain underwater. (4 marks)

27 Calculate the molar concentration of a solution of potassium dichromate that is made by adding 12.5 g of the salt to 250.0 mL of water at room temperature. (4 marks)

28 A solution of sodium hydroxide is found to be overconcentrated and is diluted by adding 200.00 mL of water. If the original solution was 75.00 mL of 3.25 M NaOH, calculate the concentration of the diluted solution. (2 marks)

29 Use solubility rules and balanced chemical equations to justify how you would determine whether a substance was potassium carbonate, potassium hydroxide or potassium phosphate. (5 marks)

30 A chemist has five colourless 1.0 M solutions of HNO_3, HF, H_3PO_4, NaOH and CH_3NH_2. The substances have been labelled A–E for the purpose of identification. The conductivity of each solution is measured in Siemens per metre ($S\,m^{-1}$) and their colour with bromothymol blue added is recorded.

TABLE 1 Conductivity of five solutions

Solution	Conductivity ($S\,m^{-1}$)	Colour with bromothymol blue
A	0.12	Yellow
B	0.21	Blue
C	4.62	Blue
D	5.51	Yellow
E	0.58	Yellow

Deduce the identities of the five solutions and justify your choices. (5 marks)

31 Consider methane (CH_4) and dichloromethane (CH_2Cl_2).

- **a** Determine the shape of the molecules. (1 mark)
- **b** Identify each the molecules as polar or non-polar. (1 mark)
- **c** Recognise the type of intermolecular forces the molecules interact through. (2 marks)
- **d** Explain the difference in boiling points of the molecules. (2 marks)

32 The data in Table 2 was obtained from a separation of three dye samples conducted using paper chromatography. The mobile phase used was 10% by volume of ethanol in water.

TABLE 2 Data from a separation of three dye samples

Sample/ standard	Component	R_F
Standard	Blue	0.12
Standard	Red	0.55
Standard	Yellow	0.33
Sample	Yellow	0.33
Sample	Blue	0.56
Sample	Red	0.89

- **a** Sketch a chromatograph of the resulting separation. (2 marks)
- **b** Draw a justified conclusion about the properties of the yellow and red components of the sample. In your response, refer to their R_F values and affinities to the mobile and stationary phases. (4 marks)
- **c** Draw a justified conclusion about the identity of the components in the sample. (2 marks)

33 A 20.00 mL solution of 0.750 M NaOH was added to 50.00 mL of a 0.250 M HCl solution.

- **a** Determine the balanced equation for this reaction. (1 mark)
- **b** Determine which reactant is in excess and by what amount (in mol). (6 marks)
- **c** Calculate the pH of the final solution. (4 marks)

34 The solubilities of three substances were measured and represented in Figure 7. The substances are ammonium chloride, potassium nitrate and nitrogen.

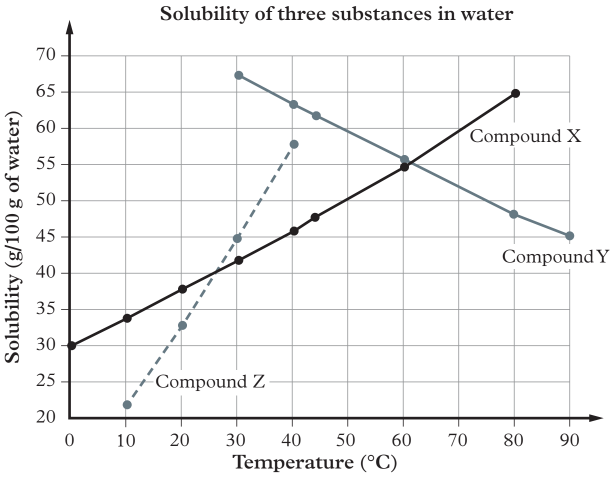


FIGURE 7 Solubility of three substances in water

- **a** One of the compounds can be identified from the graph alone. Identify the compound and explain how you could deduce its identity. (2 marks)
- **b** Calculate the concentration, in $g L^{-1}$, of a saturated 50.0 mL solution of compound X at 45°C. (2 marks)
- **c** Calculate the mass of compound Y that precipitates out of a 250.00 mL solution as its temperature increases from 60°C to 90°C. (3 marks)

35 A group of students conducted an investigation to determine the effect of concentration on the rate of a reaction. They run the trials outlined in Table 3.

TABLE 3 Effect of concentration on the rate of a reaction

Reaction	HNO_3 concentration (M)	Mass of Zn (g)
A	0.5	0.5
B	1.0	0.5
C	2.0	0.5

The students record their results and produce the graph in Figure 8.

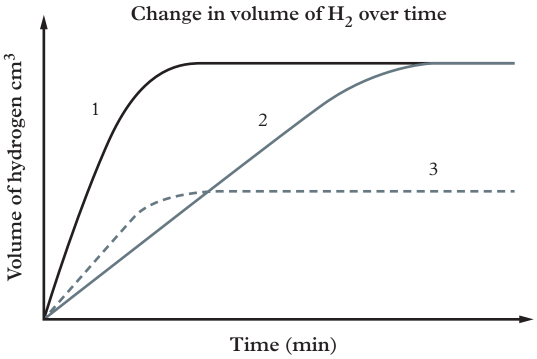
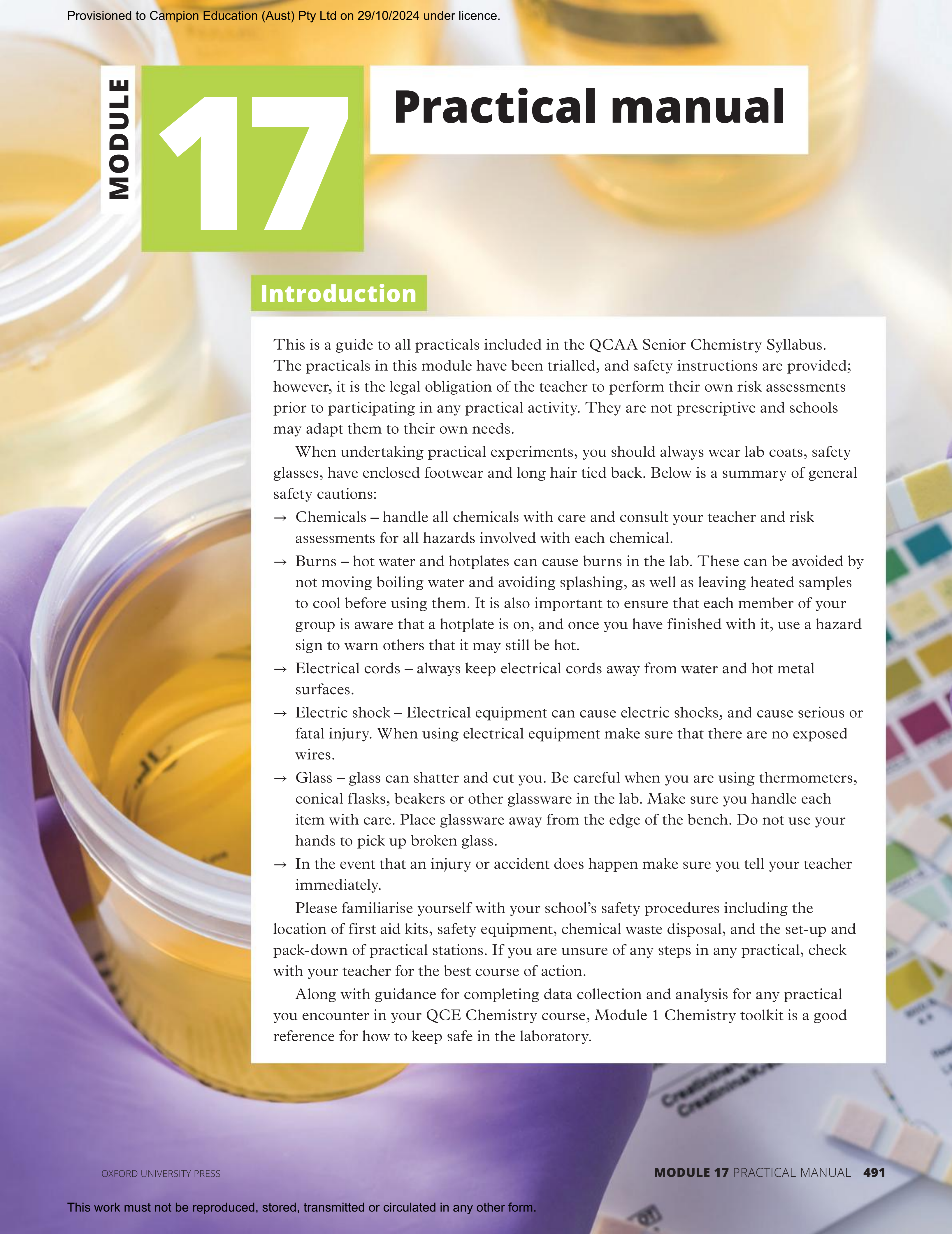


FIGURE 8 Change in hydrogen volume over time

- **a** Identify the dependent and independent variables. (2 marks)
- **b** Identify the variables that must be controlled. (2 marks)
- **c** Use the data to deduce which reactions (A, B or C) are represented by curves 1, 2 and 3. (3 marks)
- **d** Explain why curves 1 and 2 produce a greater volume of H_2 gas than curve 3. (3 marks)
- **e** Draw a justified conclusion for the experiment. (2 marks)

TOTAL MARKS

/100



MODULE 17

Practical manual

Introduction

This is a guide to all practicals included in the QCAA Senior Chemistry Syllabus. The practicals in this module have been trialled, and safety instructions are provided; however, it is the legal obligation of the teacher to perform their own risk assessments prior to participating in any practical activity. They are not prescriptive and schools may adapt them to their own needs.

When undertaking practical experiments, you should always wear lab coats, safety glasses, have enclosed footwear and long hair tied back. Below is a summary of general safety cautions:

- → Chemicals – handle all chemicals with care and consult your teacher and risk assessments for all hazards involved with each chemical.
- → Burns – hot water and hotplates can cause burns in the lab. These can be avoided by not moving boiling water and avoiding splashing, as well as leaving heated samples to cool before using them. It is also important to ensure that each member of your group is aware that a hotplate is on, and once you have finished with it, use a hazard sign to warn others that it may still be hot.
- → Electrical cords – always keep electrical cords away from water and hot metal surfaces.
- → Electric shock – Electrical equipment can cause electric shocks, and cause serious or fatal injury. When using electrical equipment make sure that there are no exposed wires.
- → Glass – glass can shatter and cut you. Be careful when you are using thermometers, conical flasks, beakers or other glassware in the lab. Make sure you handle each item with care. Place glassware away from the edge of the bench. Do not use your hands to pick up broken glass.
- → In the event that an injury or accident does happen make sure you tell your teacher immediately.

Please familiarise yourself with your school's safety procedures including the location of first aid kits, safety equipment, chemical waste disposal, and the set-up and pack-down of practical stations. If you are unsure of any steps in any practical, check with your teacher for the best course of action.

Along with guidance for completing data collection and analysis for any practical you encounter in your QCE Chemistry course, Module 1 Chemistry toolkit is a good reference for how to keep safe in the laboratory.

Full lessons for this module are available on Oxford Digital. They are also provided as part of a printable resource that can be used in class.

Unit 1 Practicals

Lesson 2.2	Simulating Geiger–Marsden's gold foil experiment
Lesson 2.7	Investigating periodic table trends using a database
Lesson 4.3	Conducting a flame test
Lesson 4.4	Simulating atomic absorption spectroscopy
Lesson 4.6	Simulating mass spectrometry
Lesson 5.3	Separating mixtures
Lesson 6.4	Testing for saturation
Lesson 6.6	Investigating physical properties
Lesson 7.4	Determining the empirical formula of magnesium oxide
Lesson 7.6	Investigating limiting reactants
Lesson 7.8	Investigating the percentage yield of copper carbonate
Lesson 8.3	Applying Hess's law
Lesson 8.5	Measuring the enthalpy of a reaction using calorimetry
Lesson 8.6	Measuring the heat of combustion

Unit 2 Practicals

Lesson 9.2	Constructing 3D models of molecules
Lesson 10.3	Separating components of a mixture using paper chromatography
Lesson 11.2	Investigating the properties of gases
Lesson 11.4	Investigating Boyle's law
Lesson 13.2	Investigating precipitation reactions
Lesson 13.4	Investigating the effect of temperature on solubility
Lesson 15.3	Investigating reactions of acids
Lesson 16.2	Investigating the rate of chemical reactions

Answers

Module 1: Chemistry toolkit

Lesson 1.1 Studying QCE Chemistry

Check your learning 1.1

1 Chemistry is the study of matter.
2 Student answers will vary.
3 No, not all observations of chemical phenomena come from experiments.
4 **a** Developing or selecting the right medication by understanding how drugs interact with the bacteria and within the koala's body
  b Checking that the water is safe for drinking
  c Analysing the original materials/pigments for accurate restoration and preservation
  d Knowing the chemical composition of the soil to harvest better yields of wheat crops
  e Understanding the pH of sea waters to ensure the survival of the habitat of marine organisms
  f Checking wastewater from mine sites at Mt Isa to ensure that there is no contamination of the town water supply
  g Using polymer chemistry knowledge to design shade structures that provide adequate sun protection
  h Knowing the characteristics of aromatic compounds so that when they are mixed, they produce a favourable perfume
  i Knowing which mosquito insecticides will work on certain surface water areas
5 Student answers will vary.

Lesson 1.7 Processing and analysing data

Worked example 1.7A

$25.2 \pm 0.05\,g$

Worked example 1.7B

±0.2%

Worked example 1.7C

$110 \pm 0.3\,mL$

Worked example 1.7D

$8.1 \pm 0.3\,g\,L^{-1}$ or $8.1\,g\,L^{-1} \pm 3.6\%$

Worked example 1.7E

±4%

Worked example 1.7F

a Go to Oxford Digital.
b $9.0\,g\,cm^{-3}$
c $y = 9x$
d $13\,cm^3$

Worked example 1.7G

1.96 ± 0.59

Check your learning 1.7

1 Absolute uncertainty is a measure of the range of uncertainty for a value. Percentage uncertainty is the uncertainty of measurements expressed as a percentage of the final value. Percentage error is the percentage difference between the measured value and the true value.
2 **a** 24.3 mL
  b ±0.3 mL
  c ±1.2%
3 ±3%
4 Transforming non-linear data by applying a mathematical function to one of the variables so that the relationship between the variables becomes closer to a straight line
5 Mean: 13.5 mg; mode: 13.4 mg; median: 13.4 mg
6 Positive: the value of one variable increases as the other variable increases; negative: when one variable's value decreases as the other increases
7 Correlation: a link between a change in the independent variable and a change in the dependent variable; causation: when a change in a single variable affects a change in a second variable
8 **a** 33.8 mL
  b ±1.0 mL
9 **a** Length: ±4.4%; width: ±8.7%, height: ±20.0%
  b $10.4 \pm 0.6\,cm^3$
  c ±1.9%
  d ±33.1%
10 **a** Go to Oxford Digital.
  b 17.3 cm
11 **a**

Volume (mL)	Mean Pressure (kPa)
60	101
55	110
50	120
45	132
40	146
35	165
30	190

  b Go to Oxford Digital.
  c Inverse power; plot y (pressure) against $\frac{1}{x}$ or $\frac{1}{\text{volume}}$
  d

$\frac{1}{V}$ (mL⁻¹)	Pressure (kPa)
0.017	101
0.018	110
0.020	120
0.022	132
0.025	146
0.029	165
0.033	190

  e Go to Oxford Digital.
12 **a** Go to Oxford Digital.
  b 0.0039 M

Lesson 1.14 Review: Chemistry toolkit

Multiple choice

1	B	**2**	B	**3**	B	**4**	A	**5**	A
6	C	**7**	D	**8**	C	**9**	D	**10**	C
11	C	**12**	A	**13**	A	**14**	A		

Short response

15 Significant figures indicate the level of precision and reliability of measurements, while scientific notation makes use of exponents to efficiently handle very small or very large values.
16 Systematic errors: the contribution to the uncertainty in a measurement result that are identifiable and quantifiable; random errors: those due to measurement uncertainty and uncontrollable effects of a method
17 The method used to estimate how uncertainties influence the final calculated result
18 Improve the precision and validity results

19 **a** No correlation
  b High negative correlation
  c Low positive correlation
  d High positive correlation
  e High negative correlation
  f No correlation
20 **a** 14
  b 14
  c 0.01
  d 116
21 High precision, low accuracy
22 Random error
23 $2.3 \pm 0.1\,cm^3\,s^{-1}$
24 ± 0.1 mm
25 Systematic error; clean, service and calibrate the electronic balance
26 Precise but inaccurate
27 **a** Playground = 9:00 am; Library and classroom = 12:00 pm
  b With the exception of 11:00 am and 12:00 pm, the playground consistently has the highest pollution levels throughout the day. The library and classroom locations generally have lower pollution levels, with the classroom having slightly lower levels than the library in the afternoon.
  c The concentration of air pollutants will be higher in school locations closer to major emission sources throughout the school day.
  d Student answers will vary.
28 6%; inaccurate and invalid
29 **a** Over time, coins in circulation will experience different levels of everyday wear and tear.
  b The mean should be reported to the same number of decimal places as the precision of the original measurements, i.e. 2.980 ± 0.333 g
30 Student answers will vary.

Module 2: Atomic structure, isotopes and the periodic table

Lesson 2.1 The atomic model

Challenge

1 mm to 1 cm

Worked example 2.1A

20 neutrons, 18 electrons

Real-world chemistry

1 Student answers will vary.
2 Student answers will vary.

Check your learning 2.1

1 Protons: positive charged and have a mass of 1.67×10^{-24} g; neutrons: no charge but same mass as a proton; electrons: negative charge and insignificant mass
2 Atomic number (Z) = total number of protons in nucleus; mass number (A) = total number of protons and neutrons in the nucleus; isotope = atoms with the same number of protons and electrons but different numbers of neutrons
3 **a** $^{49}_{22}Ti$
  b $^{79}_{35}Br$
4 11 protons, 12 neutrons, 11 electrons
5 13 protons, 14 neutrons, 10 electrons
6 1833 electrons

Lesson 2.3 Isotopes

Worked example 2.3A

^{54}Fe, ^{56}Fe, ^{57}Fe, ^{58}Fe

Worked example 2.3B

24.31

Worked example 2.3C

^{14}N = 99.31%; ^{15}N = 0.69%

Check your learning 2.3

1 Atoms of an element have the same number of protons but different number of neutrons
2 Same proton and electron configurations, similar chemical properties, different physical properties (e.g. water made with ^2H is heavier than water made with ^1H)
3 ^{A}X or X-A
4 The weighted average mass of an atom once the mass and percentage abundance of all isotopes have been considered, expressed relative to carbon-12
5 207.2169
6 ^{203}Tl = 28.7%; ^{205}Tl = 71.3%
7 ^{63}Cu (copper-63) and ^{65}Cu (copper-65)
8 He isotope 2 (99.797%), C isotope 1 (99.298%), N isotope 1 (99.298%), Br isotopes 1 and 2 are equally abundant (50.85% and 49.15%)

Lesson 2.4 The periodic table

Challenge

Student answers will vary.

Real-world chemistry

1 Student answers will vary.
2 Student answers will vary.
3 Argon; it is inert and abundant.

Check your learning 2.4

1 The periodic table is arranged into vertical groups and horizontal periods, with elements divided into sections depending on their characteristics and in ascending order by atomic number.
2 Mercury is a liquid at room temperature and has high surface tension, a low melting point and is toxic.
3 The atom is the building block for all matter, whereas an element is a type of atom as defined by its atomic number.
4 Non-metal

Lesson 2.5 Electron configurations

Challenge

For every electron in a sublevel, there are half the number of orbitals.

Worked example 2.5A

$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^7$, $[Ar]4s^2\,3d^7$

Check your learning 2.5

1 1 orbital in s, 3 in p, 5 in d, 7 in f
2 2 in s, 6 in p, 10 in d, 14 in f
3 Electrons in higher energy levels (7s), experience less electrostatic attraction to the positively charged nucleus than electrons in the lower energy levels (2s), which are closer to the nucleus.
4 s < p < d
5 $[Ar]\ 4s^2\ 3d^4$
6 Germanium (Ge)
7 Potassium (K)
8 **a** 6
  b 10
  c 8
9 4
10 Boron: $1s^2\ 2s^2\ 2p^1$; chlorine: $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$; lithium: $1s^2\,2s^1$
11 The bromine 4s and 3d sublevel orbitals are very close in energy, so there are exceptions to the Aufbau principle.
12 Once a noble gas is referenced, the electrons in the highest remaining energy levels (for the calculated element) can be added up more easily.

Lesson 2.6 Trends in the periodic table

Skill drill

1 Go to Oxford Digital.
2 As the atomic number increases, the number of electrons increases. For the first two elements, an increase in radius occurs. As the atomic number increases to 3, the radius increases again. As the atomic number increases from 3 to 7, the radius decreases.
3 Sodium has a lower atomic number than potassium, and therefore fewer electrons.
4 Lithium has fewer protons and therefore, a smaller attraction to the outer electron, resulting in a larger atomic radius. Phosphorus has more protons and a greater attraction to its electrons, resulting in a smaller atomic radius than lithium.

Worked example 2.6A

Boron; fewer filled orbitals and its electrons are closer to the nucleus

Skill drill

1 558.3 kJ mol^{-1} (In), 834 kJ mol^{-1} (Sb), and 1,170.4 kJ mol^{-1} (Xe)
2 Indium's one p valence electron is easier to lose. Antimony has three p valence electrons which are more difficult to lose. Xenon has no free valence electrons; a very, very high amount of energy is needed to lose electrons.

Check your learning 2.6

1 The atomic radius refers to the size of an atom. The ionic radius is smaller or larger than the atomic radius when the ion is positive or negative, respectively.
2 The electron shielding effect occurs when the electrons within an atom are "shielded" from the pull of the nucleus.
3 Basic to amphoteric to acidic
4 Metallic and non-metallic characters are properties of elements. Metals are usually shiny, solid, with high melting points, heat and electrical conductivities, and densities, are very malleable and ductile, and lose electrons in chemical reactions (e.g. iron). Non-metals commonly have low melting and boiling points, do not conduct heat and electricity, and do not have a shiny surface (e.g. chlorine).
5 **a** Cl
  b N
  c As
6 **a** K
  b Sr
  c Te
7 **a** Cl
  b F
  c N
8 Electronegativity increases across a period, except for the noble gases which have full valence shells. Electronegativity decreases down a group because of the electron shielding effect.
9 More energy is therefore required to remove valence electrons from the bottom left of the periodic table, and first ionisation energy increases.

Lesson 2.8 Review: Atomic structure, isotopes and the periodic table

Multiple choice

1 C **2** B **3** D **4** B **5** C
6 C **7** D **8** C **9** B **10** A
11 D **12** C **13** C **14** B **15** D

Short response

16 An atom contains a nucleus which consists of protons and neutrons. Electrons are located outside of the nucleus but remain close due to an electrostatic attraction to the protons.
17 Atomic number is the number of protons in an atom. Mass number is the total number of subatomic particles in the nucleus (the protons and neutrons).
18 Isotopes are atoms of an element that have the same number of protons but different numbers of neutrons.
19 The actual electron configurations contain both 4s and 3d sublevels half-full, which are more stable than full 4s and partially full 3d sublevels.
20 s < p < d < f
21 **a** The ability of a positive nucleus to electrostatically attract the valence electrons of the adjacent atom
  b The energy required to remove an electron from a gaseous atom, forming a positively charged ion
22 **a** 63.55
  b 28.10
  c 14.01
  d 39.11
23 35.22
24 **a** $^{23}_{11}\text{Na}^{+}$
  b $^{19}_{9}\text{F}$
  c $^{59}_{27}\text{Co}$
  d $^{48}_{22}\text{Ti}$
  e $^{37}_{17}\text{Cl}^{-}$
25 **a** $1s^2 2s^2 2p^6 3s^2 3p^3$
  b $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
  c $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
26 **a** [Ne] $3s^2\ 3p^3$
  b [Ar] $4s^1$
  c [Ar] $4s^1 3d^5$
  d [Ar] $4s^2 3d^{10} 4p^5$
27 ^{107}Ag = 51.85%; ^{109}Ag = 48.15%
28 Zinc (Zn) and gallium (Ga)
29 **a** $1s^1$; should be $1s^2$
  b $3s^3$; should be $3s^2$
  c $4p^2$; should be $3d^2$
30 [He] $2s^2 2p^4$, [Ar] $4s^2$, [Ar] $4s^2 3d^2$
31 $1s^2 2s^2 2p^6$
32 Atomic radii decrease across a period and increase down a group. Ionic radii are smaller than atomic radii when the ion is positive. Ionic radii larger than atomic radii when the ion is negative.

Data drill

1 Caesium
2 91 pm
3 Go to Oxford Digital.
4 Student answers may vary.
5 Student answers may vary.
6 Approximately 42°C

Module 3: Introduction to bonding

Lesson 3.1 Types of bonding

Worked example 3.1A

$(NH3)_3PO_4$

Skill drill

1 Sodium and the halides (except for sodium iodide) as sodium has a greater electronegativity than potassium
2 The melting point decreases as molecular size increases.

Worked example 3.1B

Ammonium sulfide

Real-world chemistry

1 Each ion deficiency has characteristic observable effects but some symptoms are symptomatic of multiple deficiencies, and therefore not usable as a diagnosis.
2 Student answers will vary.

Worked example 3.1C

The nitrogen atom contributes 3 electrons, and each hydrogen atom contributes 1 electron to form three single bonds.

Challenge

Carbon tetrachloride, dinitrogen tetrachloride, carbon tetrabromide, sulfur hexafluoride, diphosphorus tetroxide

Check your learning 3.1

1 Inner electrons are electrons located in the inner orbitals, while valence electrons are the electrons located in an atom's outermost energy level only.

2 Transition metals

3 Atoms lose, gain or share electrons to achieve a stable configuration of eight electrons in the valence energy level.

4 Atoms lose, gain or share electrons to achieve eight electrons in the valence energy level.

5 Positively charged cations form electrostatic attractions with oppositely charged anions. The opposite charges cancel out to form a neutral ionic compound. The combined charges of Na^+ and Cl^- (1+ and 1–) cancel out to form NaCl.

6 Both ionic and covalent bonding occur between atoms and involve electrons. However, ionic bonding involves the transfer of electrons, while covalent bonding involves the sharing of electrons to form bonds.

7 **a** Gained 1 electron
  b Lost 1 electron
  c Gained 3 electrons
  d Gained 1 electron
  e Lost 2 electrons

8 **a** Bromine trichloride
  b Ammonia

9 **a** Li_2CO_3
  b MgO
  c $Ca(NO_3)_2$
  d KCl.

10 Go to Oxford Digital.

11 Superscript = overall ionic charge (negative = anion); subscripts = number of atoms of each element

Lesson 3.2 Lewis structures

Worked example 3.2A

a Go to Oxford Digital.
b Go to Oxford Digital.

Worked example 3.2B

Go to Oxford Digital.

Check your learning 3.2

1 Valence electrons are drawn as dots surrounding the atom's elemental symbol. Lines represent covalent bonds.

2 Hydrogen has a full valence 1s sublevel of two electrons when covalently bonded.

3 **a** Go to Oxford Digital.
  b Go to Oxford Digital.
  c Go to Oxford Digital.
  d Go to Oxford Digital.
  e Go to Oxford Digital.
  f Go to Oxford Digital.

4 Determine the Lewis structures of the following ions:
  a Go to Oxford Digital.
  b Go to Oxford Digital.
  c Go to Oxford Digital.
  d Go to Oxford Digital.

5 Methane has a tetrahedral three-dimensional shape and is not planar, as shown in the diagram on the right.

Lesson 3.3 Review: Introduction to bonding

Multiple choice

1	B	**2**	B	**3**	B	**4**	C
5	C	**6**	B	**7**	B	**8**	C

Short response

9 Ions are atoms or groups of atoms that are charged because of an imbalance in the number of electrons to its protons.

10 Covalent bonding occurs when non-metal atoms share electrons to form bonds.

11 s and p sublevels

12 Barium (Ba) loses two electrons to form a Ba^{2+} cation and forms ionic attractions with the hydroxide ion, a polyatomic anion (OH^-) formed by the covalent bonding between oxygen and hydrogen, to form a neutral compound, $Ba(OH)_2$.

13 Two sodium (Na) atoms lose one electron each to form two Na^+ cations and form ionic attractions with the thiosulfate ion ($S_2O_3^{2-}$), which is a polyatomic anion formed by the covalent bonding between oxygen and sulfur with two extra electrons, to form a neutral compound, $Na_2S_2O_3$.

For questions 14 to 20 go to Oxford Digital.

21 Both ionic and covalent bonding occur between atoms and involve electrons. However, ionic bonding involves the transfer of electrons, while covalent bonding involves the sharing of electrons to form bonds.

22 Both monatomic and polyatomic ions have either a net positive or negative charge. Monatomic ions are single atoms, whereas polyatomic ions are composed of more than one atom covalently bonded together.

23 **a** K_2S
  b Ra_3P_2
  c AlN
  d $Sr(ClO_2)_2$
  e BPO_4
  f $(NH_4)_2SO_3$

24 Go to Oxford Digital.

25 **a** BI_3
  b MgO

26 **a** Dinitrogen monoxide
  b Boron trichloride
  c Ammonium chromate

27 **a** Phosphorus does not form a double bond.
  b The carbon atom will only form a single bond with chlorine.

28 **a** Carbon is missing electrons.
  b Boron only has three valence electrons.

29 **a** Go to Oxford Digital.
  b The electronegativity difference is slightly lower for HF than for NaBr. Bonds are more covalent for HF than NaBr, which is more ionic in character. However, bonding is often not clearly ionic or covalent for polar covalent bonding elements.
  c Ionic bonds tend to be very strong and form solids with higher melting points and that are good conductors of heat and electricity when molten. Polar covalent bonds tend to form liquids and gases with lower melting points and are not very good conductors.

Data drill

1 BCl_3

2 823°C

3 Ionic: lithium chloride, sodium chloride, potassium chloride; covalent: beryllium chloride, boron trichloride, carbon tetrachloride

4 Student answers will vary.

5 –107.3°C; poor conductivity

Module 4: Analytical techniques

Lesson 4.1 Principles of spectroscopy

Real-world chemistry

1 Heat energy causes electrons to move from their ground state to an excited state. The emission spectrum is created when the electrons return to the ground state, releasing the extra energy as specific wavelengths of light.

2 Each spectral line represents a particular wavelength that would be specific to the transitions that occur when the electron returns to the ground state for that particular element.

3 An impure sample will change the spectrum, producing more spectral lines for each of the elements in the sample.

Check your learning 4.1

1 **a** The emission spectrum of hydrogen shows distinct wavelengths of light that are emitted as electrons return to their ground state.

b Electrons can only move to higher energy orbits if they gain enough energy to reach the excited state of the higher orbital and will emit that specific wavelength when returning to the ground state.

2 The orbits of electrons are further apart when low in energy, close to the nucleus, but as they move further from the nucleus, becoming higher in energy, the orbits get closer together.

3 Go to Oxford Digital.

4 **a** Five

b Yes. Higher energy transitions would occur in the UV range.

Lesson 4.2 Techniques for identifying elements

Skill drill

1 Student answers will vary.

2 Student answers will vary.

3 Student answers will vary.

Real-world chemistry

1 It is expensive to mine and process ores.

2 Hyperkalemia: too much potassium in the blood; hypokalemia: too little potassium in the blood. Symptoms are not always easy to diagnose, so rapid analysis techniques like AAS can help identify potassium levels in a blood sample and inform medical solutions.

3 It is important to analyse the mercury content in fish to make sure that they will not be harmful to consume.

4 To check the quality of the land

Worked example 4.2A

Diluted sample concentration: 68 μg L^{-1}; undiluted sample concentration: 6,800 μg L^{-1}

Challenge

15 mg L^{-1}

Check your learning 4.2

1 **a** When a metal sample is placed in a Bunsen flame, the electrons will absorb then emit specific wavelengths of energy that will allow you to identify the metal that is present.

b Electrons gain energy to reach an excited state in a higher electron shell, producing black lines on an absorption spectrum. Electrons that are in an excited state return to the ground state and release the extra energy they have gained, producing coloured spectral lines in the emission spectrum.

c A prism or diffraction grating will separate out the wavelengths of light into distinct lines for each wavelength emitted.

2 Similar to a flame test, the heat from a flame excites the electrons in a sample and causes them to absorb energy, in line with its absorption spectrum. The intensity of light transmitted after electrons return to their ground state is detected. Lower intensity of transmitted light corresponds to a higher absorbance, and a higher concentration of the metal atoms in the sample.

3 34.6 μg L^{-1}

4 Sample 1 is within the recommended levels for lead(II) contamination. However, as sample 2 has higher than the recommended levels, more testing is recommended to verify the extent of contamination and investigate a possible clean-up of the area.

Lesson 4.5 Mass spectrometry

Worked example 4.5A

24.3

Worked example 4.5B

39.11

Worked example 4.5C

^{79}Br = 51%; ^{81}Br = 49%

Challenge

48.161%, 108.97

Check your learning 4.5

1 spectrometry; isotope; ionisation; smaller/lower; radius; separates; ratio; abundance; atomic

2 ^{69}Ga = 60.30%; ^{71}Ga = 39.70%

3 ^{151}Eu = 48%; ^{153}Eu = 52%

4 107.86

5 28.11

6 10.80

7 121.76

Lesson 4.7 Review: Analytical techniques

Multiple choice

1 B **2** B **3** D **4** C **5** D

6 A **7** A **8** C **9** C **10** C

Short response

11 An absorption spectrum shows the band of lines observed when an element absorbs light. An emission spectrum shows the band of lines observed when an element emits light.

12 RIM: the mass of a particular isotope when compared to carbon-12; RAM: the weighted average mass of the atoms of an element, relative to carbon-12; mass number: is number of protons and neutrons in the nucleus of an atom

13 The mass of each isotope is measured by comparing it to the carbon-12 isotope, which is called relative isotopic mass. All the isotopes of elements and their different relative masses and abundances are considered to determine relative atomic mass. Each isotope has a different number of neutrons and therefore a different relative isotopic mass.

14 **a** The light beam from a lamp is pulsed at the sample which has been sprayed into a flame, exciting the electrons. They absorb some of the light beam; the remaining light is transmitted and detected.

b They would be the same because electrons absorb and emit the same amount of energy.

15 **a** 28.09

b 151.96

c 65.40

16 **a** ^{203}Tl = 29.5%; ^{205}Tl = 70.5%

b ^{113}In = 5.2%; ^{115}In = 94.8%

c ^{85}Rb = 70.6%; ^{87}Rb = 29.4%

17 **a** 24.54

b 69.72

c 20.18

18 **a** 47.87

b 51.997

19 **a** 186.17

b 207.011

20 **a** 9.22 mg L^{-1}

b 3.33 mmol L^{-1}

21 **a** ^{6}Li = 11.59%; ^{7}Li = 88.81%

b ^{191}Ir = 38%; ^{193}Ir = 62%

Data drill

1 ~16 mg L^{-1}

2 50 mg L^{-1}

3 As concentration of magnesium ions increases, the absorbance increases.

4 0.74

5 AAS has difficulty detecting accurate absorbance of samples at high concentration.

Unit 1 Topic 1 review

Multiple choice

1	C	**2**	A	**3**	B	**4**	D	**5**	B
6	D	**7**	D	**8**	C	**9**	D	**10**	A
11	D	**12**	D	**13**	D	**14**	C	**15**	C

Short response

16 **a** In a group (vertical column in the periodic table), elements have the same number of valence electrons. In a period (horizontal row in the periodic table), elements have the same number of electron shells.

b Nitrogen: period 2; phosphorus: period 3

17 **a** Chromium has the same electron configuration as argon plus additional electrons.

b $x = 2, y = 5$

c It would be more difficult for calcium to lose the second electron because of the increasing imbalance between the number of positively charged protons and the number of electrons remaining.

18 **a** They have the same number of protons but different numbers of neutrons.

b 28.11

c Silicon

19 Both consist only of covalent bonds between more than one type of atom. Silicon dioxide forms a giant covalent network, whereas carbon dioxide is a small covalent compound.

20 **a** $Al(g) \rightarrow Al^{+}(g) + e^{-}$

b As each electron is removed, the excess positive nuclear charge increases and the nucleus pulls more strongly on the valence electrons. More energy is required to remove subsequent electrons.

c Removing the last s orbital electron (third electron) means the shielding effect would decrease and attraction to nucleus would increase, requiring greater amount of energy to remove the fourth electron (in the p orbital).

21 **a** Mercury and small amounts of krypton

b Energy causes electrons to move from their ground state to an excited state in higher electron shells. When the electrons return to the ground state, they release the extra energy as specific wavelengths of light.

22 **a** Ca loses two electrons to form Ca^{2+}. Its number of electron shells decreases from four to three.

b Si^{4+} has lost all four valence electrons in its third shell, so its overall radius decreases compared to phosphorus, which has electrons in this shell. The three extra valence electrons in P^{3-} repel one another and cause the ionic radius to be even larger.

23 Go to Oxford Digital.

24 Left to right: basic ($Na_2O + H_2O \rightarrow 2NaOH$) > amphoteric > acidic ($SO_3 + H_2O \rightarrow H_2SO_4$)

25 **a** Ammonium sulfate; ionic bonding

b Dibromine monoxide; covalent bonding

c Potassium dichromate; ionic bonding

26 93.12%

27 **a** Go to Oxford Digital.

b $22.2\,\mu g\,L^{-1}$

c The water sample is unsafe to drink.

Module 5: Compounds and mixtures

Lesson 5.1 Pure substances

Worked example 5.1A

a 18.45 g

b $0.9973\,g\,cm^{-3}$

Check your learning 5.1

1 A pure substance can either contain only one type of element (e.g. Ar) or one type of compound (e.g. NaCl).

2 Go to Oxford Digital.

3 Student answers may vary.

4 **a** $7.194\,g\,cm^{-3}$

b Not pure

5 **a** Substance C

b Melts over a range of temperatures instead of a single discrete temperature

c It is not pure.

Lesson 5.2 Mixtures

Skill drill

1 Pure: B and C (small melting point range); mixtures: A and D (larger melting point range)

2 Sample A – benzoic acid and aspirin; sample B – ferrocene; sample C – biphenyl or stearic acid; sample D – 2-napthol and benzoic acid

3 Some of the substances could not be clearly identified due to impurities and their ratios, and similarities in melting point.

Real-world chemistry

1 Ensures that variables are controlled and the data collected can reliably be used to draw conclusions

2 Student answers will vary.

Check your learning 5.2

1 A mixture with a uniform composition

2 Student answers will vary.

3 It is an inconsistent mixture of differing components.

4 The technique chosen for separation will depend on the properties of the substance being separated from the mixture.

5 A homogeneous mixture appears consistent throughout the mixture (e.g. cordial). A heterogeneous mixture has an inconsistent composition (e.g. a salad).

6 **a** Homogeneous

b Heterogeneous

c Heterogeneous

d Heterogeneous

e Homogeneous

f Heterogeneous

7 Air is a mixture of the elements nitrogen, oxygen and hydrogen, and small amounts of other gases, which are all gases of uniform composition on a molecular level.

8 Student answers will vary.

9 Student answers will vary.

Lesson 5.3 Review: Compounds and mixtures

Multiple choice

1	C	**2**	C	**3**	D	**4**	D	**5**	D
6	C	**7**	C	**8**	C	**9**	D	**10**	B

Short response

11 A substance made up of a single type of element (e.g. Ar) or a single type of compound (e.g. NaCl).

12 Composition, phase, particle size, concentration or ratio of components, chemical interactions between components

13 The mass of a substance that occupies a specific volume

14 Changes phase from a liquid to a gas with bubbles of gas forming below the surface of the liquid

15 Homogenous mixtures have a uniform composition and the same phases throughout, while heterogenous mixtures may have non-uniform composition or different phases throughout.

16 **a** Pure substance

b Heterogeneous mixture

c Homogeneous mixture

d Homogeneous mixture

e Heterogeneous mixture

17 **a** Heterogeneous

b Homogeneous

c Homogeneous

d Heterogeneous

e Heterogeneous

18 Water (solid), oxygen (liquid), nitrogen (liquid), mercury (solid) and iron (solid)

19 **a** 4°C

b Volume would increase.

c 12°C

d No, the temperature can be read as 1°C and 7°C.

20 Student answers will vary.

21 Remove solid contaminants to ensure that the air mixed with fuel is a clean gas.

b The build-up of dust and debris can lead to poor combustion/loss of power.

Data drill

1 ≈18 g mL^{-1}

2 Aluminium = 2.7 g mL^{-1}; wood = 0.5 g mL^{-1}; water = 1 g mL^{-1}

3 They have densities that are significantly higher than the other substances. A limited volume of mass could be measured compared to the other substances.

4 Cyclohexane, petrol, vegetable oil and wood, which have lower densities than water

5 Go to Oxford Digital.

Module 6: Properties and structure of materials

Lesson 6.1 Properties of ionic compounds

Worked example 6.1A

It is made of Cu^+ and SO_4^{2-} ions which are free to move and carry current.

Skill drill

1 Student 1: anecdote; student 2: evidence; student 3: anecdote and limited evidence

2 Student 1: scientific; student 2: scientific; student 3: partially scientific

Check your learning 6.1

1 **a** CsI

b AlF_3

c Ga_2O_3

d Li_3N

e Na_2S

f $CaBr_2$

2 **a** Sodium nitride

b Potassium nitrate

c Magnesium sulfide

d Lithium carbonate

e Iron(II) hydroxide

f Iron(III) nitrate

3 **a** Student answers will vary.

b Student answers will vary.

4 Go to Oxford Digital.

5 When a force is applied, like charges approach each other and then repel away from each other, causing the ionic solids to split or shatter along smooth planes.

6 A large amount of energy is needed to break the multiple electrostatic interactions between cations and anions.

7 Student answers may vary.

Lesson 6.2 Properties of metallic substances

Check your learning 6.2

1 **a** Zn^{2+}

b Ag^+

c Cu^+ and Cu^{2+}

d Mg^{2+}

e Fe^{2+} and Fe^{3+}

2 **a** The sea of electrons are free to move and carry electric current.

b Strong attractive forces between the delocalised electrons and the metal cations must be overcome to melt or boil metals.

c The sea of delocalised electrons reflect all wavelengths of light.

3 **a** Go to Oxford Digital.

b Go to Oxford Digital.

4 **a** All metals and ionic compounds have relatively high melting and boiling points, metals share the same colour and ionic compounds share the same colour. However, metals have a lower melting point than their compound form. Metals are good conductors of electricity, whereas ionic compounds are not.

b Delocalised electrons in metals reflect all wavelengths of light well, whereas light is not reflected as well in the compounds. Delocalised electrons in metals allow easy flow of electric current, while the ionic bonds in both compounds resist the free flow of current.

5 **a** Metals conduct electricity in the solid state because the metal has delocalised **electrons** that allow the free flow of electric current.

b Metals are held together by **metallic bonds**, which are the attractive forces between metal ions and delocalised electrons in a metal lattice.

c The strength of the metallic bonds depends on the number of delocalised electrons, the charge of the cation and the size of the cation.

d Metals melt once sufficient heat energy has been provided to break the ordered arrangement of metal ions in the lattice.

e The boiling point of a metal is a better indication than the melting point of the strength of the metallic bonding within that metal.

Lesson 6.3 Properties of simple covalent molecules

Worked example 6.3A

C_8H1_8

Real-world chemistry

1 $C_{12}H_{22}O_{11}$

2 All four amino acids have a chain in common, with unique (residue) sections that extend away from the protein chain. Glycine and alanine have smaller chains, while aspartic acid and glutamine have longer, more complex chains.

3 Saturated fatty acids contain only carbon–carbon single bonds. Unsaturated fatty acids contain at least one carbon–carbon double bond.

Check your learning 6.3

1 Very pure water contains almost no dissolved ions and, as a covalent molecule, is resistant to ionisation. Tap water has a higher concentration of dissolved ions, which can carry charge.

2 The intramolecular bonds are very strong and the intermolecular forces are significantly weaker. Heating covalent molecular substances disrupts the weak intermolecular forces first, so melting and boiling will occur more readily, before the intramolecular bonds are disrupted.

3 Go to Oxford Digital.

4 Certain reactants can add across the double bond in alkenes.

Lesson 6.5 Properties of giant covalent networks

Real-world chemistry

1 10,000

2 ZnO and TiO_2

3 Each carbon atom is bonded to its neighbouring atoms with very strong covalent bonds in a single layer hexagonal arrangement, giving graphene a higher melting and boiling point, thermal and electrical conductivity, and hardness.

Challenge

To convert graphite into diamond, energy and pressure would be needed to break the covalent bonds in graphite and force them into a new configuration.

Skill drill

1 Student answers will vary.

2 Student answers will vary.

3 Student answers will vary.

Check your learning 6.5

1 Student answers will vary.

2 Each silicon atom is covalently bonded to four oxygen atoms in a tetrahedral arrangement and each oxygen atoms bridges two silicon atoms. The bonds require a large amount of heat energy to break.

3 Diamond and quartz both form giant covalent network solids. In diamond, each carbon atom is covalently bonded to four other neighbouring carbon atoms, whereas in quartz, each silicon atom is surrounded by four oxygen atoms in a tetrahedral arrangement and each oxygen atoms bridges two silicon atoms.

4 **a** Go to Oxford Digital.

b Go to Oxford Digital.

Lesson 6.7 Review: Properties and structures of materials

Multiple choice

1	C	**2**	B	**3**	D	**4**	B	**5**	B
6	B	**7**	C	**8**	D	**9**	A	**10**	A
11	D	**12**	B	**13**	D	**14**	B		

Short response

15 KI is formed by ionic bonding into a crystalline lattice structure. ICl has a covalent bond between the I and Cl atoms.

16 If benzene were to undergo an addition reaction, then the extra stability that comes from having resonance would be lost.

17 **a** Go to Oxford Digital.

b In the solid state, ions cannot move freely to carry electrical charge.

18 Methane is the smallest of the alkanes, so the intermolecular forces will be weaker than for larger alkane molecules.

19 Aqueous ionic solutions will dissociate to form ions which carry a charge. Although most covalent compounds do not dissociate into ions when dissolved in water, ethanoic acid can form H^+ and CH_3COO– ions, which can carry charge.

20 Student answers will vary.

21 **a** Metals contain positively charged ions and delocalised electrons, while ionic substances are made from positively charged cations and negatively charged anions.

b Covalent molecular substances are held together internally by covalent bonds, and to each other by intermolecular forces. Covalent network substances are held together by strong covalent bonds that extend to create a strong lattice of covalent bonds.

22 Sample A: covalent bonding; sample B: ionic bonding; sample C: covalent bonding; sample D: ionic bonding

23 Substance U: lithium chloride; substance V: lead; substance W: aluminium; substance X: silicon dioxide; substance Y: calcium carbonate; substance Z: graphite

24 Go to Oxford Digital.

25 Go to Oxford Digital.

Data drill

1 2,230°C

2 94°C

3 B < D < A < C

4 A: $MgCl_2$; B: C_2H_6; C: quartz; D: Mg

5 Will not dissolve; has a very high melting and boiling point

Unit 1 Topic 2 review

Multiple choice

1	B	**2**	D	**3**	B	**4**	B	**5**	B
6	D	**7**	D	**8**	D	**9**	B	**10**	D
11	D	**12**	D	**13**	B	**14**	B	**15**	B

Short response

16 **a** Distillation

b Evaporation

c Separating funnel

17 **a** Homogeneous

b Heterogeneous

18 Both are covalent network compounds where all the atoms are bonded to their neighbours with very strong covalent bonds, which require a large amount of energy to overcome. The electrons in graphite can become delocalised and conduct electricity, whereas all valence electrons in diamond are involved in bonding.

19 "The starting liquid had a freezing point 5°C lower than that of the condensed liquid." "The liquid remaining in the starting flask had a density 1.2 times that of the liquid which distilled over and condensed in the receiving flask."

20 **a** Ionic compounds are brittle, not malleable, because when struck, ions of the same charge can be brought close together, resulting in repulsion. This repulsion causes ionic crystals to shatter.

b The electrical conductivity of molten ionic compounds is due to ions being able to move freely through the molten liquid.

c In solid sodium chloride, each sodium ion in the lattice is bonded to six chloride ions in the lattice.

d Some ionic compounds, such as potassium chloride, are soluble in water because the partially negative oxygen end of the water molecule is attracted to the cations, and the partially positive end of the molecule is attracted to the anions.

21 Gold forms metallic bonds. As it is hammered or stretched, the positively charged ions can slip past each other without causing repulsion as the delocalised electrons are able to move freely between the ions and maintain their attraction to the ions, holding the metal together without breaking.

22 **a** Carbon has four valence electrons that can all undergo strong covalent bonding with other carbon atoms, leading to a very high melting point when compared to other elements.

b Its melting point will be much lower because it only forms weak attractive forces.

23 W: glucose; X: sodium chloride; Y: citric acid; Z: calcium carbonate

24 **a** The ratio of ionic radii of about 3.2 is either the point at which compounds change solubility or there are other factors causing a substance to become soluble.

b For a substance to be soluble, at least one of the ions present must have a 1+/– charge.

c Insoluble; has a high ratio of ionic charge

Module 7: Mole concept and law of conservation of mass

Lesson 7.1 The mole

Worked example 7.1A

a 4.8×10^{24} molecules
b 1.9×10^{25} atoms

Worked example 7.1B

a 4.8×10^{24} formula units
b $N(Fe^{3+}) = 4.8 \times 10^{24}$ ions; $N(NO_3^-) = 1.4 \times 10^{25}$ ions

Worked example 7.1C

1.17 mol

Check your learning 7.1

1 One mole is equal to 6.02×10^{23} representative particles (atoms, molecules, ions, formula units). This is Avogadro's number.
2 6.02×10^{23}
3 **a** 6.0×10^{23} atoms
  b 1.2×10^{24} atoms
  c 3.0×10^{23} atoms
  d 4.2×10^{24} atoms
4 **a** 2.41×10^{24} atoms
  b 4.82×10^{24} atoms
  c 2.41×10^{24} atoms
5 **a** 1.20×10^{23} atoms
  b 2.41×10^{23} ions
6 **a** 0.585 mol
  b 0.424 mol

Lesson 7.2 The mass of a mole

Worked example 7.2A

a 80.06 g mol^{-1}
b 30.08 g mol^{-1}

Worked example 7.2B

a 0.0831 mol
b 92.0 g
c 44.0 g mol^{-1}

Check your learning 7.2

1 **a** A measure of how much matter is contained within a substance
  b The weighted mean of the masses of atoms of an element once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope
  c The mass of one mole of a substance
2 **a** 16.05 g mol^{-1}
  b 241.88 g mol^{-1}
  c 60.09 g mol^{-1}
  d 249.71 g mol^{-1}
3 **a** 6.231 mol
  b 0.4134 mol
  c 1.664 mol
  d 0.4005 mol
4 **a** 1.5 mol
  b 0.061 mol
  c 0.034 mol
  d 0.10 mol
5 **a** 4.5 g
  b 13 g
  c 73 g
  d 285 g
6 Ammonia

Lesson 7.3 Empirical and molecular formulas

Worked example 7.3A

11.2%

Worked example 7.3B

N_2O

Skill drill

1 Student answers will vary.
2 Student answers will vary.
3 Student answers will vary.

Worked example 7.3C

$C_2H_4O_2$

Check your learning 7.3

1 The percentage that each element contributes to the mass of a compound
2 The empirical formula is the simplest whole-number ratio of the elements in a compound, while the molecular formula is the chemical formula of a covalent molecular compound that indicates the number of each element present in one molecule.
3 %C = 40.0%; %H = 6.73%; %O = 53.3%
4 $C_3H_8O_2$
5 B_2H_3
6 Empirical formula = HO; Molecular formula = H_2O_2

Lesson 7.5 Stoichiometric ratios

Worked example 7.4A

104 g

Worked example 7.4B

a KI
b 1.39 g

Real-world chemistry

1 Designers would try to test fuel injectors to get an exact match for the 2 : 25 ratio between octane and oxygen.
2 Student answers will vary.
3 A fine mist will provide more even burning and precision timing would lead to fuel burning at the most evenly spread time.

Check your learning 7.5

1 **a** The ratio of the amounts in moles of the reactants and products in a reaction
  b The reactant that is totally consumed when a chemical reaction is complete
  c The reactant that is not totally consumed by a chemical reaction
2 17.04 g
3 196 g
4 **a** $FeCl_2 + H_2S \rightarrow FeS + 2HCl$
  b $FeCl_2$
  c 62.42 g
5 **a** 696 g
  b 76.1 g

Lesson 7.7 Yield

Worked example 7.7A

96.02%

Check your learning 7.7

1 Theoretical yield is the amount of product predicted from stoichiometric calculations. Experimental yield is the actual amount of product formed.
2 There cannot be more product produced than that predicted by the molar ratio of the equation.
3 **a** 20 g
  b 90%
4 **a** 30.7 g
  b 81%

Lesson 7.9 Review: Mole concept and law of conservation of mass

Multiple choice

1 D **2** D **3** A **4** A **5** B
6 C **7** A **8** B **9** D **10** D

Short response

11 $n(CO_2)$

12 Student answers may vary.

13 6.020×10^{23} atoms

14 **a** 194.22 g mol^{-1}
   b 92.02 g mol^{-1}
   c 74.10 g mol^{-1}
   d 84.01 g mol^{-1}

15 1.76×10^{24} atoms

16 Student answers will vary.

17 **a** 3.5 mol
   b 14 mol
   c 21 mol
   d 14 mol
   e 21 g
   f 220 g

18 **a** 1,420 g
   b 400 g

19 **a** 798.5 g
   b 402.5 g

20 1,497 g

21 Empirical formula is the simplest ratio of the elements present in a molecule (e.g. CH_2O). A molecular formula shows the number of each element present in the molecule (e.g. $C_6H_{12}O_6$).

22 **a** $C_3H_7O_3N$
   b $C_3H_7O_3N$

23 **a** 0.0406 mol
   b 5.76 g

24 **a** CHO
   b $C_8H_8O_2$
   c CH_2O
   d %C = 40.66%; %N = 23.71%; %O = 27.08%

25 87.0%

26 93.7%

27 **a** Sb_2S_3
   b 51.0 g
   c 107 g
   d 93.0%

Data drill

1 A: HCl; B: Zn

2 As the volume of HCl increases, the mass of H_2 gas produced increases.

3 Less gas is produced than expected after 80 mL is added. This could be due to limitations or uncertainty in the measurement of the acid added or in the volume of gas produced.

4 Adding extra HCl will not produce more gas as after the addition of 80 mL of HCl, there was no extra gas produced. This indicates that the zinc has been used up (limiting reactant). The final mass of H_2 gas after the addition of 200 mL of HCl will be 0.15 g.

5 Using 2.5 g of zinc would not affect the reaction rate and the slope of the graph in section A would be the same. However, the smaller mass of zinc would be fully reacted after the addition of a smaller volume of HCl, i.e. the plateau in the curve will occur earlier and at a lower volume of H_2. Go to Oxford Digital.

Module 8: Chemical reactions

Lesson 8.1 Physical and chemical changes

Worked example 8.1A

$2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$

Worked example 8.1B

$2KI(aq) + Br_2(aq) \rightarrow 2KBr(aq) + I_2(aq)$

Worked example 8.1C

$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl$

Worked example 8.1D

$2HCl\ (aq) + Ca(OH)_2\ (aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$

Worked example 8.1E

$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

Worked example 8.1F

$2C_8H_{18}(l) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(g)$

Worked example 8.1G

$2Al(s) + 6HCl \rightarrow 2AlCl_3(aq) + 3H_2(g)$

Real-world chemistry

1 $4C_{12}H_{23}(l) + 71O_2(g) \rightarrow 48CO_2(g) + 46H_2O(l)$

2 $2C_{17}H_{34}O_2(l) + 49O_2(g) \rightarrow 34CO_2(g) + 34H_2O(l)$

3 339.41 g for diesel; 289.82 g for biodiesel

Check your learning 8.1

1 The law of conservation of mass states that matter cannot be created or destroyed.

2 If a substance is dissolved in water, it is assigned the (aq) state. Pure substances in their melted state are liquids (l).

3 Addition or removal of heat energy causes the motion of particles to change significantly, but without any increase or decrease in temperature.

4 **a** $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$
   b $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
   c $N_2(g) + O_2(g) \rightarrow 2NO(g)$
   d $2C_4H_6(g) + 11O_2(g) \rightarrow 8CO_2(g) + 6H_2O(l)$
   e $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$
   f $3KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$
   g $2Zn(s) + Pb(NO_3)_4(aq) \rightarrow 2Zn(NO_3)_2(aq) + Pb(s)$
   h $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow 2NaCl(aq) + BaSO_4(s)$
   i $2NaOH(aq) + Cu(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + Cu(OH)_2(s)$
   j $H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g)$

5 $2Ni_2O_3(s) \rightarrow 4Ni(s) + 3O_2(g)$

6 $CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(g)$

7 Phase changes involve a transition between different states of matter without a change in chemical composition. Chemical reactions involve the breaking and forming of chemical bonds, resulting in the formation of different substances with different chemical compositions.

8 **a** AgBr(s)
   b $KNO_3(s)$
   c $BaSO_4(s)$
   d $O_2(g)$
   e $H_2O(l)$
   f S(s)
   g CO(g)
   h $CaCO_3(s)$
   i Mg(s)
   j $Cl_2(g)$
   k $Br_2(l)$
   l $N_2(g)$

Lesson 8.2 Chemical energy and thermochemistry

Worked example 8.2A

4,411 kJ mol^{-1}

Worked example 8.2B

a –1,263 kJ mol^{-1}

b Exothermic

Worked example 8.2C

$CH_3OH(l) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
$\Delta H = -890\ kJ$

Worked example 8.2D

$16CO_2(g) + 18H_2O(g) \rightarrow 2C_8H_{18}(l) + 25O_2(g)$

$\Delta H = 10{,}920\ kJ$

Worked example 8.2E

a $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

$\Delta H = -1{,}360\,kJ$

b $-2.33\times10^4\,kJ$

Real-world chemistry

1 2.24 times

2 1.03 times

3 The rapid production of a large volume of high temperature gas involves rapid expansion. This exerts a powerful force on the surrounding materials, which is essential for breaking rocks or demolishing structures.

Check your learning 8.2

1 The amount of energy stored within a substance is called its enthalpy, or heat content.

2 Balanced chemical reaction equation and ΔH value

3 Energy cannot be created or destroyed.

4 Heat is a form of energy referred to as "thermal energy". Temperature is a measure of the average kinetic energy of the particles.

5 Endothermic: energy of reactants > energy of products; endothermic: energy of reactants < energy of products

6 Go to Oxford Digital.

7 **a** $-1{,}904.5\,kJ\,mol^{-1}$

b Go to Oxford Digital.

c Exothermic

d $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

$\Delta H = -3{,}809\,kJ$

8 **a** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

$\Delta H = -890\,kJ$

b $2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$

$\Delta H = -650\,kJ$

c $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$

$\Delta H = -363\,kJ$

9 Any energy absorbed or released during a reaction cannot disappear; it is conserved. In exothermic reactions, this energy is usually released to the surroundings, making them hotter.

Lesson 8.4 Specific heat capacity

Worked example 8.4A

2.2 kJ

Worked example 8.4B

$1.6\times10^2\,kJ\,mol^{-1}$

Worked example 8.4C

a $-57\,kJ\,mol^{-1}$

b $NaOH(aq) + CH_3COOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l)$

$\Delta H = -57\,kJ$

Worked example 8.4D

$1.7\times10^3\,kJ\,mol^{-1}$

Skill drill

1 **a** Student answers may vary.

b Student answers will vary.

c Student answers may vary.

d Student answers will vary.

Real-world chemistry

1 2

2 14.3 mol

3 Cellular respiration is needed to convert lactic acid back to glucose in the liver requires a great deal of energy; a large amount of oxygen is required.

Check your learning 8.4

1 A measure of the amount of heat energy, in joules, that it takes to increase the temperature of 1 g of water by 1°C

2 q = energy released or absorbed in joules; m = mass of substance in grams; c = specific heat capacity in $J\,g^{-1}\,°C^{-1}$ or $J\,g^{-1}\,K^{-1}$; ΔT = change in temperature in °C or K

3 ΔH can be measured using calorimetry. For aqueous solutions, a simple solution calorimeter is used. For combustion reactions, a spirit burner is used.

4 $5.1\times10^3\,J$

5 $-90\,kJ\,mol^{-1}$

6 $-4.1\times10^2\,kJ\,mol^{-1}$

7 Student answers may vary.

8 Student answers will vary.

Lesson 8.7 Review: Chemical reactions

Multiple choice

1	C	**2**	A	**3**	B
4	D	**5**	D	**6**	C

Short response

7 Energy is the capacity to do work or produce change. Heat is a form of energy transferred between systems or objects with different temperatures.

8 Energy is transferred between the chemical system and its surroundings, but the total energy remains constant as the reaction proceeds.

9 Go to Oxford Digital.

10 **a** Different molecules are made up of different atoms bonded in different ways, so the reactants and products in chemical equations have different amounts of potential energy.

b If the total bond energy of the products is lower than that of the reactants, the reaction is exothermic. If the total bond energy of the products is higher than that of the reactants, the reaction is endothermic.

11 Go to Oxford Digital.

12 A change of state does not result in a change in temperature because the energy added during the process to change the state of a substance is used to overcome intermolecular forces between substance particles rather than increasing the kinetic energy of the particles.

13 Student answers will vary.

14 **a** $C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$; combustion

b $2AgNO_3(aq) + MgCl_2(aq) \rightarrow 2AgCl(aq) + Mg(NO_3)_2(aq)$; double displacement

c $2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2\,(aq) + 2H_2O(l)$; acid–base reaction

d $Cl_2(aq) + 2KBr(aq) \rightarrow Br_2(aq) + 2KCl(aq)$; single displacement reaction

15 Go to Oxford Digital.

16 $4{,}000\,kJ\,mol^{-1}$

17 $3{,}234\,kJ\,mol^{-1}$

18 **a** $-1{,}255\,kJ\,mol^{-1}$

b The value in Table 3 is more valid as it would have been obtained under very controlled conditions, whereas when we calculate the heat of combustion using bond enthalpies, we use average values that do not consider that the same bond can have different energies depending on their reaction environment.

19 **a** $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

$\Delta H = -780\,kJ$

b $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

$\Delta H = -2{,}220\,kJ$

20 $-130\,kJ\,mol^{-1}$

Data drill

1 $-900\,kJ\,mol^{-1}$

2 $700\,kJ\,mol^{-1}$

3 As the number of carbons increases, the molar heat of combustion increases.

4 The heat of combustion of alcohols is slightly lower than that of alkanes.

5 $-6{,}700\,kJ\,mol^{-1}$

Unit 1 Topic 3 review

Multiple choice

1 B **2** B **3** B **4** D **5** C
6 C **7** C **8** D **9** B **10** D
11 C **12** D **13** A **14** B **15** C

Short response

16 a 2.92 mol
  b 1.76×10^{24} molecules
17 $N(Cu^{2+}) = 4.48 \times 10^{21}$ ions; $N(Cl^{-}) = 8.96 \times 10^{21}$ ions
18 a 55.89 kJ
  b 34.9°C
19 a 2.61 g
  b 70.9%
20 a 5.0 kJ
  b $-1{,}214\ kJ\ mol^{-1}$
21 a $Pb(NO_3)_2(aq) + 2KI\ (aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$
  b As the reaction progresses, the mass of lead(II) iodide formed increases until a mass of 3.5 g of lead nitrate is reacted, then plateaus.
  c The potassium iodide was the limiting reactant when larger masses of lead nitrate were used.
22 a 11,000 g
  b 7.31 kg

Unit 1 Review

Multiple choice

1 C **2** A **3** D **4** D **5** A
6 D **7** D **8** B **9** C **10** A
11 D **12** C **13** C **14** C **15** D
16 D **17** C **18** D **19** A

Short response

20 0.5000 mol
21 Homogeneous mixtures have a uniform composition (e.g. cordial), while heterogenous mixtures have an inconsistent composition (e.g. muesli).
22 Go to Oxford Digital.
23 39
24 As you move across the period, electronegativity increases. As you move down a group, electronegativity decreases.

25 a $[Ar]\ 3d^6\ 4s^2$
  b $[Ne]\ 3s^2\ 3p^6$
26 a i N_2O_4
  ii $Ca_3(PO_4)_2$
  b i Sulfur trioxide
  ii Sodium carbonate
27 All allotropes are composed only of carbon atoms covalently bonded together in large structures and giant covalent networks. The carbon atoms are arranged differently in each allotrope.
28 26 protons, 30 neutrons, 23 electrons
29 $C_6H_{12}O_6$
30 a Go to Oxford Digital.
  b Go to Oxford Digital.
31 a As molar mass increases, the boiling points of both alkanes and alcohols increase.
  b C–C = 0.0; C–H = 0.4; C–O = 0.8; O–H = 1.2
  c The greater the difference, the more polar the bond. The more polar the bond, the stronger the intermolecular forces.
32 a 8.4 kJ
  b $2{,}900\ kJ\ mol^{-1}$
  c 30%
33 a Substance A = decanoic acid; substance B = silicon carbide; substance C = lead; substance D = sodium chloride
  b Solid: unable to conduct electricity; aqueous: will be able to conduct electricity
  c Substance B; has a very high melting point and would retain its structural integrity at high temperatures
34 a $2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$
  b KI
  c 2.77 g
35 a $140\ \mu g\ mL^{-1}$
  b Absorbance reading is beyond the range of the calibration curve; extrapolating could introduce inaccurate and unreliable results
  c Dilute the original sample again

Module 9: Intermolecular forces

Lesson 9.1 Shapes of covalent molecules

Challenge

Not all molecules with the same shape will have the same bond angles. Bond angle is determined by the total number of valence shell electron pairs as well as the number of bonded electron pairs in a molecule.

Worked example 9.1A

Bent; 104.5°

Challenge

Three atoms; first carbon has a tetrahedral shape, the second carbon has a trigonal planar shape and the oxygen has a bent shape

Skill drill

1 IV: molecules examined; DV: bond angle
2 Go to Oxford Digital.
3 Bond angle deviates from 109° as molar mass increases.
4 Internal angles may differ due to differences in atomic radii.

Real-world chemistry

1 Student answers will vary.
2 Student answers will vary.

Check your learning 9.1

1 Tetrahedral molecules consist of a central atom with four covalent bonds and no non-bonding electrons. Pyramidal molecules consist of a central atom forming three covalent bonds and one lone pair. Bent molecules have a central atom with two covalent bonds and two lone pairs.
2 Single bonds: electron pair shared between two atoms; double bonds: sharing of two electron pairs; triple bonds: sharing of three electron pairs
3 a Linear. Go to Oxford Digital.
  b Linear. Go to Oxford Digital.
  c Pyramidal. Go to Oxford Digital.
  d Trigonal planar. Go to Oxford Digital.
  e Bent. Go to Oxford Digital.
  f Tetrahedral. Go to Oxford Digital.
  g Trigonal planar. Go to Oxford Digital.
  h Linear. Go to Oxford Digital.
  i Bent. Go to Oxford Digital.

Lesson 9.3 Polarity and intermolecular forces

Challenge

They are highly unreactive due to their full outer electron shells and do not form any known compounds. Electronegativity values are experimentally obtained by analysing compounds.

Worked example 9.3A

a Polar
b Non-polar

Skill drill

1 Methanol = 5.1; acetonitrile = 5.7; toluene = 3.8
2 Go to Oxford Digital.
3 Methanol: high accuracy and precision; acetonitrile: high accuracy, low precision; toluene: low accuracy, high precision

Worked example 9.3B

Dispersion forces

Real-world chemistry

1 DNA strands need to separate when a copy of the DNA is required to make a protein.
2 The "rungs" are unequal and held together with different strengths.
3 Student answers will vary.

Check your learning 9.3

1 Covalent bonds are classified as polar depending on the electronegativity differences between the atoms involved in the bond. If this is large enough, the molecule will have permanent partially positive and negative ends.
2 Both are intermolecular forces. Hydrogen bonding is a specific type of dipole–dipole attraction, where the partially positive hydrogen atom attached to the most partially negative atoms from the periodic table (F, O or N) on one molecule is attracted to the lone pairs of the F, O or N that is bonded to H in another molecule.
3 **a** Dispersion forces
  b Dispersion forces
  c Dispersion forces, dipole–dipole attractions, hydrogen bonding
  d Dispersion forces, dipole–dipole attractions, hydrogen bonding
  e Dispersion forces, dipole–dipole attractions, hydrogen bonding
  f Dispersion forces, dipole–dipole attractions
4 Go to Oxford Digital.
5 Hydrogen bonding between base pairs of two different strands; covalent bonding within base pairs of the same strand
6 Student answers will vary.

Lesson 9.4 Intermolecular forces and physical properties

Worked example 9.4A

As ammonia has much stronger intermolecular forces than hydrogen sulfide, the boiling point of ammonia will be higher than that of hydrogen sulfide.

Real-world chemistry

1 As temperature increases, the vapour pressure increases.
2 Molecules with stronger molecular forces will have lower vapour pressure and a decreased likelihood of evaporation.
3 Ammonia will have a lower boiling point than water, and a higher vapour pressure.

Worked example 9.4B

Water (polar) and octane (non-polar) can only form weak dispersion forces, so octane will not dissolve in water.

Check your learning 9.4

1 When heat energy is applied, kinetic energy is increased and molecules begin to move faster and will soon have enough energy to start moving away from neighbouring molecules. If the temperature is high enough, the molecules can change state.
2 The boiling point of non-polar molecules increases as the number of carbon atoms increases. The boiling points of polar covalent molecules are higher than alkanes of similar mass.
3 Vapour pressure decreases as boiling point increases.
4 F_2 has only dispersion forces holding its molecules together, which require a small amount of energy to break. HF can form strong hydrogen bonds, which require more energy to break.
5 **a** Will dissolve
  b Will not dissolve
6 Ammonia has a greater solubility than methane in water since it can form hydrogen bonds with water.
7 HF interacts through hydrogen bonds which require more energy to break apart. HCl, HBr, and HI form dipole–dipole attractions, and have a lower boiling point than HF. All compounds also form dispersion forces and the strength of these forces increases with molar mass. HCl has the lowest molar mass and will boil more easily. HI has the highest molar mass and needs more energy to break apart.

Lesson 9.5 Review: Intermolecular forces

Multiple choice

1	C	2	A	3	C	4	B
5	A	6	A	7	A		

Short response

8 **a** Non-polar. Go to Oxford Digital.
  b Polar. Go to Oxford Digital.
  c Polar. Go to Oxford Digital.
  d Non-polar. Go to Oxford Digital.
  e Polar. Go to Oxford Digital.
9 Covalent bonds occur between atoms when they share their electrons. An intermolecular force is an attraction between molecules, and is much weaker than a covalent bond.
10 The greater the kinetic energy of a particle, the faster it moves. When heat energy is applied, molecules begin to move faster, and soon have enough energy to start moving away from neighbouring molecules. The temperature at which the heat energy is sufficient for molecules to break free of their rigid structure, or turn a solid into a liquid, is called the melting point. The boiling point is the temperature at which the heat energy is sufficient to break molecules free or turn a liquid into a gas.
11 Hydrogen bonding because it involves the most electronegative (F, O and N) and electropositive (H) atoms on the periodic table
12 At any given time, there may be more valence electrons on one side of the molecule than on the other, randomly creating a dipole that is not permanent. Dipoles on neighbouring molecules can interact and form an intermolecular interaction.
13 NaCl forms very strong ionic bonds, which require much higher temperatures to break. Water is a polar molecule that forms strong hydrogen bonds, which are weaker than ionic bonds. Methane is non-polar, and therefore, forms dispersion forces, which are the weakest form of intermolecular forces and require the least energy to break.
14 Boiling points are higher when intermolecular forces are stronger.
15 As the number of carbon atoms increases, the boiling point of the alkane also increases due to the increase in dispersion forces.
16 DNA contains two molecules with nitrogen-containing base pairs with polar N–H and O–H bonds. The two long strand-like molecules are joined by hydrogen bonds, which are weak enough to be easily broken without damaging the structure during DNA replication.
17 Ethanol interacts with other ethanol molecules through a small amount of hydrogen bonding and dispersion forces. Strong hydrogen bonding occurs between water molecules. Therefore, ethanol vaporises more readily than water, giving a higher vapour pressure.
18 Room temperature is a higher temperature than the boiling point at –161.6°C for CH_4, so methane is a gas at room temperature. As the number of carbon atoms increases (e.g. in octane), the boiling point of the molecule also increases because of the increase in dispersion forces.

19 The electronegativity differences between atoms of each molecule are 0; they are non-polar. They do have slightly different molecular sizes, numbers of electrons, electron densities and bond numbers. Generally, molecules which have more electrons would have stronger dispersion forces, but this is somewhat counteracted by the smaller size of these molecules across the same row of the periodic table. All these molecules will thus have similar dispersion forces, and similar melting points.

20 **a** A saturated fatty acid contains only single bonds, whereas an unsaturated fatty acid contains one or more double bonds. The presence of a double bond means that unsaturated fatty acid molecules can form one of two molecular geometries that include a *trans* or *cis* double bond. Fatty acids with *cis* double bonds experience less intermolecular forces. Computational chemistry can predict the geometric structure of larger molecules (like fatty acids), their interactions with other molecules and therefore their behaviour.

b Student answers will vary.

21 Go to Oxford Digital.

22 Linear: polar hydrogen cyanide; trigonal planar: polar CH_2O; pyramidal: polar ammonia; bent: polar water; tetrahedral: polar fluoromethane

23 **a** Dispersion forces

b Dispersion forces

c Dispersion forces, dipole–dipole attractions, hydrogen bonding

d Dispersion forces, dipole–dipole attractions, hydrogen bonding

24 **a** Methanol

a Butane

c Ethanoic acid

25 Incorrect

26 Fluorine is in the second period of the periodic table. There is a very low degree of nuclear shielding and the p orbital electrons are pulled very close to the nucleus. Fluorine needs only one electron to fill the p orbital (octet rule) to create stability.

27 1-propanol can form hydrogen bonds, whereas the strongest forces formed by ethyl methyl ether are dipole–dipole attractions. Less energy is required to break the dipole–dipole attractions than the hydrogen bonds, and the boiling point will be lower for ethyl methyl ether.

28 The glucose molecule is structured as a much smaller ring-shaped molecule, allowing much better contact with water molecules by the exposed O–H groups. The smaller size and ease of forming intermolecular hydrogen bonds allows it to dissolve.

29 When water freezes to become solid ice, the hydrogen bonding between the water molecules causes them to form an open three-dimensional lattice and leave spaces within its structure. This spreading out of the water molecules creates a stable lattice structure on freezing. When ice melts, this kinetic energy allows the molecules of water to move closer to each other and fill in the previously open spaces in the solid lattice. The increase in intermolecular spacing during crystal formation of the ice causes a decrease in density when compared to water.

30 Solid carbon dioxide contains weak dispersion forces that are easily broken and the solid has a very high vapour pressure. Thus, at room temperature and standard pressure, there is enough heat energy to increase the kinetic energy of the molecules and break the weak dispersion forces to convert dry ice straight from the solid phase state to the gaseous phase.

Data drill

1 Melting point and boiling point increase as the number of carbon atoms increase.

2 Decane; dispersion forces

3 Go to Oxford Digital.

4 The melting points of alkanes are all below room temperature, so they are all solids at room temperature. They appear to somewhat "see-saw" as they rise, so that those with an even number of carbons have higher melting points than those with odd numbers of carbons. This could be due to solid alkanes with even numbers of carbon atoms packing together more closely in the solid state, and more energy being required to move them apart.

5 Melting and boiling points increase as molar mass of straight-chain alkanes increase. Increasing chain length in straight alkanes corresponds with increasing dispersion forces between molecules.

Module 10: Chromatography techniques

Lesson 10.1 Principles of chromatography

Skill drill

1 Does not account for the fact that not all polar substances dissolve in water and not all non-polar substances form layers with water

2 Qualitative as it is categorical and non-numerical

Worked example 10.1A

Highest affinity: hexanoic acid because it forms strong intermolecular forces due to the combination of the long non-polar carbon chain (dispersion forces) and the polar -COOH group (hydrogen bonding)

Lowest affinity: ethanol because it has weaker intermolecular forces compared to the other substances, with only a small -OH group contributing to its hydrogen bonding

Check your learning 10.1

1 **a** The phase that flows in chromatography, moving the components of a sample at different rates over the stationary phase

b The phase to which the components of a chromatographic sample are adsorbed; remaining fixed in place, it interacts differently with each component, causing them to move at different rates.

c The interaction of a substance within a sample with the mobile or stationary phases; the different affinities that each component has for the phases makes them move at different rates, separating them.

d Components of the sample are more attracted to either the mobile or the stationary phase, depending on their intermolecular forces. This makes the components move at different rates and separates them.

2 Polar components are more attracted to a polar mobile phase. According to the principle "like dissolves like", a mixture containing water (polar) and hexane (non-polar), if water is used as the mobile phase, the water component of the mixture will be more attracted to the mobile phase. Water will interact with hydrogen bonding and dipole–dipole attractions which are stronger than the dispersion forces of non-polar hexane.

3 Both involve the interaction between molecules and the stationary phase. Adsorption is the process where molecules adhere to the stationary phase, whereas desorption is the reverse process of molecules detaching from the surface and returning to the mobile phase.

4 Ethanol because it contains a hydroxyl (-OH) group that can form strong hydrogen bonds with the mobile phase compared to ethyl ethanoate, which can only engage with weaker dipole–dipole attractions

5 **a** Component B because it is polar and will have a higher affinity for the mobile phase due to dipole–dipole attractions

b Component C because it is the least polar and will therefore interact more strongly with the non-polar stationary phase

c Component B because it is the most polar and will form stronger dipole–dipole attractions with the polar stationary phase, leading to greater adsorption compared to the less polar components

d Component C (which will interact the least with the polar stationary phase and desorb quickly into the polar mobile phase) and component A (which will also desorb, but not as quickly as component C)

6 The dipole–dipole attractions between molecules and either the mobile or stationary phase can be used to separate a mixture of substances. Dispersion forces are used to separate larger molecules. The greater the dispersion force, the higher the interaction with the mobile or stationary phases. Substances capable of hydrogen bonding will have a strong affinity for polar stationary or mobile phases, allowing for polar mixtures capable of hydrogen bonds to be effectively separated.

7 If the mobile phase is non-polar (i.e. has dispersion forces), then the stationary phase should have dipole–dipole attractions. The stationary phase needs to have dipole–dipole attractions to enable the components in the sample to be separated, i.e. components of a sample are separated according to differences in the relative competition for them by the mobile phase and the stationary phase.

Lesson 10.2 Paper and thin layer chromatography

Worked example 10.2A

a R_F(A) = 0.24; R_F(B) = 0.45; R_F(C) = 0.86

b R_F = 0.45 and 0.70

c Dye A

d Dye C

e Dye B and an unidentified dye not in the standards

f Valid as all R_F values are less than 1

Real-world chemistry

1 Vitamin D must be a non-polar molecule.

2 If the mobile phase was completely polar, it would have not effectively interacted with non-polar compounds, leading to inefficient separation. The slightly polar mobile phase provides a suitable balance of interactions with both non-polar and slightly polar compounds.

3 Run a known standard of vitamin D through the chromatography process being used for the blood sample and compare the retention times and peaks to those of the sample to identify its components

4 During summer, there is more sunlight, which increases the exposure to the UVB rays that are necessary for the synthesis of vitamin D. This could cause a higher concentration of vitamin D. In winter, however, there is significantly less sunlight and therefore, reduced exposure to UVB rays. This would result in lower concentrations of vitamin D.

Check your learning 10.2

1 Measure the distance the solute moves from the origin and the distance the solvent moves from the origin. Divide the first by the second.

2 Student answers will vary.

3 **a** R_F(yellow spot) = 0.19

R_F(red spot) = 0.33

R_F(blue spot) = 0.71

b Blue spot

c Yellow spot

4 All components experience dispersion forces. The more polar substances would stay near the line of origin with a polar stationary phase due to the higher affinity of dipole–dipole attractions. The less polar components would have less affinity, likely only experiencing weaker forces (such as dispersion) and would move with the mobile phase more readily towards the solvent line.

5 **a** Sample A had R_F values of 0.25 and 0.46, which indicate the food colourings Indigotine and Erythrosine, respectively. Sample B had R_F values of 0.44 and 0.88, which indicate Erythrosine and Patent blue V, respectively. Sample C had R_F values of 0.24 and 0.69, which indicate Indigotine and Tartrazine or Green S, respectively.

b The components are not the same as the R_F values are different, meaning that they have differing affinities for either the mobile and stationary phases.

6 **a**

Team	Component 1 (cm)	Component 2 (cm)	Component 3 (cm)
1	0.56	0.79	0.21
2	0.32	0.59	0.87
3	0.56	0.79	0.21
4	0.56	0.79	0.21
5	0.32	0.59	0.87
Crime scene note	0.32	0.59	0.87

b Teams 1, 3 and 4 analysed the same pen, and teams 2 and 5 analysed the other pen.

c Teams 2 and 5

d No because the R_F values are different for component 1

e Yes, the sample contains all three components 1, 2 and 3

Lesson 10.3 Review: Chromatography techniques

Multiple choice

1 B **2** C **3** B **4** A

5 A **6** D **7** D

Short response

8 The stationary phase is either a solid with a high surface area or a liquid coated onto a solid support. The mobile phase moves through and carries a sample (solute) over the stationary phase. Components of the sample are more attracted to either the mobile or the stationary phase, depending on their intermolecular forces. This makes the components move at different rates and separates them. Two key terms are used to describe this attraction: adsorption (the components of the sample adsorb onto the stationary phase from the mobile phase if they are attracted to, or have an affinity for, the stationary phase) and desorption (the components of the sample desorb off the stationary phase into the mobile phase if they are attracted to, or have an affinity for, the mobile phase).

9 Chromatography involves separating the components of a mixture, then comparing the R_F values to those of known pure standards. The differences in the chromatograms (different spots or peaks, or R_F) can reveal the presence of impurities.

10 They can interact through dispersion forces which are the weakest intermolecular force resulting from temporary dipoles.

11 The standards have known R_F values which can be compared to those of the unknown substances. This enables accurate identification of the sample components.

12 In a valid separation, substances typically exhibit some degree of interaction with both phases, resulting in R_F values between 0 and 1.

13 a R_F(yellow dot) = 0.15; R_F(blue dot) = 0.52

b Dye C

c Dye B

d Dye B

e The chromatograph shows Dye B, which is claimed by the manufacturer to be present in the food. The chromatograph indicates that the other claimed dyes (A and C) are not present in the food sample. The R_F value of the other dye in the sample does not match any of the claimed dyes.

14 a R_F(yellow dot) = 0.30; R_F(red dot) = 0.45; R_F(purple dot) = 0.85

b Mobile phase = polar; stationary phase = polar

c Dye A

d Dye B

e Dye B

f More standard dyes need to be run to accurately identify the dyes using their R_F values.

15 a The chromatograph was produced by two-dimensional paper or thin layer chromatography. The original chromatograph paper is rotated 90°, providing a new origin line and separated further using a different mobile phase with different properties.

b Components are not well separated in the first separation due to similar interactions with the stationary phase and first mobile phase can be better separated in the second dimension with a different mobile phase. This is achieved by a difference in composition of mobile phases, as shown in the second separation.

c Occasionally, a chromatography technique will not separate some components sufficiently within the sample. This is because the components interact the same amount with the mobile and stationary phases. Rather than repeating the analysis, chemists will rotate the chromatograph by 90°, so that the sample is on the bottom. They rule a new origin line and run the analysis again, using a different mobile phase with different properties.

d i The positions of the dyes on the chromatograph can be different each time due to the use of two different solvent systems. This makes it hard to compare the spots directly with known dyes. Instead, we use R_F values, which are calculated based on the distance each dye moves relative to the solvent front. R_F values are consistent and allow for more accurate identification by comparing these values with the R_F values of known dyes.

ii The retardation factor (R_F) of a substance is the ratio of the distance moved by a substance to the distance moved by the solvent or mobile phase (i.e. the solvent front).

e R_F(grey spot) = 0.29; R_F(red spot) = 0.43; R_F(blue spot) = 0.79; R_F(yellow spot) = 0.65; R_F(purple spot) = 0.76

f i Blue spot

ii Purple spot

iii Grey spot

iv Yellow spot

v Blue spot

vi Purple spot

vii Grey spot

16 a The component with the R_F value of 0.89

b The component with the R_F value of 0.27

c The mobile phase is a mixture of water and ethanol being mostly a polar solution. Therefore, the polar molecules with the R_F value of 0.89 have the higher affinity for the polar mobile phase, which means that this component of the sample substance is polar. The less polar (or non-polar) molecules with the R_F value of 0.27 have higher affinity for the stationary phase being paper, which means that this component of the sample substance is non-polar.

17 Water is polar and will have affinity for the polar end of the ethanol molecule, allowing it to be carried through the column.

18 a Distance (A) = 14.11 cm; distance (B) = 12.58 cm; distance (C) = 9.18 cm; distance (D) = 5.44 cm

b Component A has the highest R_F value and is described as the most polar. This implies that the mobile phase must be polar because component A (being polar), has a strong affinity for the mobile phase and travels the greatest distance. Given that the mobile phase is polar, and component A has travelled the greatest distance, the stationary phase must be less polar (or non-polar). This explains why the polar components interact less with the stationary phase and more with the polar mobile phase, leading to higher R_F values for the more polar substances.

c Component A

d Component D

Data drill

1 Component A

2 The types of intermolecular forces experienced are dispersion forces, dipole–dipole attractions and hydrogen bonding

3 A < B + C < D

4 Both Components A and D experience dispersion forces (the weakest intermolecular force). Component A, which is less polar, experiences weaker dispersion forces and possibly weaker dipole–dipole attractions with the slightly polar stationary phase. Component D, which is more polar, experiences stronger dipole–dipole attractions (and possibly hydrogen bonding) with the polar mobile phase, leading to a higher R_F value.

5 The assumption is not valid. The R_F value (0.5) is not sufficient to confirm the presence of both dyes as the single spot could represent a mixture of dyes that have a similar R_F value.

6 This method is not valid for the separation of this dye. It fails to separate the orange into its distinct components of red and yellow dyes, preventing accurate separation and identification.

Module 11: Gases

Lesson 11.1 Properties of gases

Worked example 11.1A

a 139 kPA

b 106 bar

c 56,200 Pa

Worked example 11.1B

a 291 K

b 318 K

c 228 K

d 323°C

e −273°C

f −149°C

Worked example 11.1C

a 4,750 mL

b 0.0015 m^3

c 500 L

Check your learning 11.1

1 In the gaseous state, particles are in constant, random motion and have kinetic energy which is associated with the motion of particles. They have a high level of energy, so they separate and move freely.

2 The force per unit area exerted by gas particles of one substance upon another

3 **a** Gas particles have very little attraction to one another because the energy of their movement is greater than the energy of attractions between particles.

   b In the gaseous state, particles are in constant, random motion. The particles in liquids and solids have less movement and greater attraction forces.

   c The space between gas particles is much larger than the space that the particles themselves occupy. Therefore, gases expand to occupy the whole space of a container, unlike solids and liquids.

4 **a** 42°C
   b 347 K
   c 273,000 Pa
   d 0.50 bar
   e $2.5\times10^{-5}\ \text{m}^3$
   f 450 L

5 **a** The kinetic molecular theory states that gas particles have more kinetic energy when the temperature is higher. If the temperature decreases, the kinetic energy will also decrease.

   b The kinetic molecular theory states that gas particles do not lose energy when they collide. If pressure is the result of collisions that have increased, then kinetic energy has increased.

   c Kinetic molecular theory states that gases expand to fill a container. If the volume is unchanged, then the kinetic energy is unchanged.

6 $223.61\ \text{m s}^{-1}$

7 Temperature is a measure of the average heat energy of the particles within a system. This can be represented by a Maxwell–Boltzmann distribution curve. All gas particles in a reaction vessel have different amounts of energy because of these temperature differences.

8 The kinetic energy of the molecules in the scented gas increases from the burning candle. The scented gas particles move more freely and expand across the room, filling the room with the scent. It takes time for the particles to move through the space of the room, so the scent diffuses slowly.

Lesson 11.3 Gas laws

Worked example 11.3A

190 K

Worked example 11.3B

64.8°C

Worked example 11.3C

130 bar

Worked example 11.3D

0.485 mol

Worked example 11.3E

17 g

Worked example 11.3F

Cl_2

Real-world chemistry

1 To dive underwater, you need to be able to breathe a constant supply of air. To stay underwater for a longer time, the air must be under high pressure so that a larger number of gaseous particles can be contained in the small volume of the tank.

2 The tanks need to be able to hold gases at very high pressure when diving. To be able to contain these gases safely, the tanks need thick metal walls to withstand the pressure.

Real-world chemistry

1 As you dive deeper, the surrounding water exerts pressure on your body at the rate of 1 atm every 10 m you dive. This would mean the gases within your body occupy less volume than at the surface, so you need to breathe in gases at a higher pressure to maintain a constant volume of gas within the body.

2 Boyle’s law states that as pressure decreases, the volume of a gas increases. At a high altitude where pressure is lower, the volume of your airways would increase, but the concentration of air particles will be smaller. This would mean your body might not have enough air in your body to meet its needs.

Check your learning 11.3

1 **a** As the volume decreases, the pressure increases at constant temperature.

   b As the temperature increases, the pressure increases at constant volume.

   c As the temperature increases, the volume increases at constant pressure.

2 **a** 73 L
   b 2.1 L
   c 353 mL

3 **a** 0.233 g
   b 5.57×10^{23} atoms

4 **a** 248 K
   b 286 K
   c 546 K

5 15.98 kPa

6 **a** $0.0023\ \text{L K}^{-1}$
   b 1.09 L

7 Heat causes an increase in the kinetic energy of the air particles, which causes them to move more freely and with higher amounts of energy. The hot air soon occupies the balloon space and due to its heat becomes less dense than the surrounding cooler air, which causes the balloon to rise.

8 **a** Theoretically (mathematically), for the volume could be 0 L, the temperature would also be −273 K or 0°C.

   b Matter is neither created nor destroyed; therefore, the volume of gas would still exist, and the volume is 0 L at 0 K only in the mathematical or theoretical sense.

9 **a** An increased density at a fixed volume would require an increased mass of nitrogen gas. This adds more nitrogen molecules to the system, which would increase the pressure at constant volume.

   b If the density of the gas was increased, the volume of a variable volume container would expand. This adds more nitrogen molecules to the system, and the particles would push against the walls of the container to try and maintain the same pressure.

Lesson 11.5 Ideal gases

Worked example 11.5A

84.7 L

Worked example 11.5B

0.882 g

Worked example 11.5C

a $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

b 75.6 L

Skill drill

1 IV: whether participants vaped or not; DV: participants' lung capacity; controlled variable: apparatus for measuring lung capacity

2 Vaping causes a decrease in lung capacity.

3 The accuracy of the data cannot be determined since no theoretical values are available. However, because temperature and altitude were uncontrolled variables in the experiment (and since we know that temperature and pressure affect the volume of gases), the reliability of the data is reduced.

4 Student answers will vary.

Check your learning 11.5

1 Pressure = 100 kPa, temperature = 0°C or 273 K

2 Particles do not interact with one another, occupy no volume, collisions with each other are elastic and particles are in continuous, random motion.

3 **a** 19 kPa
b 0.47 L
c 8.1 g

4 Nitrogen gas

5 Go to Oxford Digital.

6 **a** $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$
b 0.756 L

7 **a** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
b H_2
c 14,133 L
d 12.03%

Lesson 11.6 Review: Gases

Multiple choice

1 D **2** C **3** D **4** D **5** A
6 A **7** D **8** B **9** A **10** C

Short response

11 **a** If the variable volume container is increased in volume at constant temperature, then the gas molecules will have more available space for greater movement occupying more space thereby reducing the pressure.
b At a decreased temperature, the gas particles in the variable volume container would decrease in energy and movement thereby occupying less space. The variable volume container would reduce in volume.

12 A pressure cooker uses pressure and steam to help cook the food. The food and water are placed inside the cooker and once closed heat is applied. The increased temperature helps to heat the water helps to convert the water to steam (gas), which increases the pressure inside the cooker. As the steam pressure increases under constant volume (closed cooker), the steam cooks the food at a faster rate.

13 Increasing the number of molecules of gas will increase the collisions with the container, which will then cause the volume to increase.

14 An increase in pressure causes an increase in the temperature of the system.

15 **a** 6.7 L
b 101 L
c 21.4 L
d 1.705 L
e 1.7 L
f 3.0 L

16 2,720 L

17 **a** 0.12 bar
b +3°C

18 **a** 53 g
b 10 L
c 38,700 kPa

19 Molar volume is 24.8 L mol^{-1} at SLC, and 22.7 L mol^{-1} at STP. The temperature at SLC is greater than at STP. This means the particles of gas will have more kinetic energy and cause the volume of the gas to increase.

20 **a** Radon
b Argon
c Argon

21 As the balloon rises into the atmosphere, the surrounding atmospheric pressure decreases. This causes the particles inside the balloon to equalise with the pressure outside. To do this, the balloon's volume will need to increase, as the pressure is inversely proportional to volume. Eventually, the volume of the balloon will increase so much that the balloon will burst.

22 According to the ideal gas equation, if pressure is increased and volume decreased, then temperature will increase. If the temperature of the gas increases enough, there will be sufficient energy to ignite the fuel without needing a spark.

23 0.0500 g

24 On the way to school, the temperature in the car allows the water to heat by the increased temperature in the car. The increase in temperature creates added pressure in the water bottle. As the student opens the bottle to take a drink, a small amount of pressure is released in the form of gas. The outside air temperature is colder than the original temperature inside the closed water bottle. Therefore, the temperature change creates a difference in pressure, and the sides of the bottle crumple inwards because the pressure in the bottle decreases.

25 Apply some heat to the outside of the ping pong ball near the dent. Wait for the pressure inside the ball to increase and slowly force out the dent.

26 The turkey timer works with increased temperature. When the turkey starts to heat in the oven, the black plug at position D melts, creating a fixed end. Once the black plug is fixed, the spring (C) is released and the red tag (A) is popped out due to the spring releasing inside the tube (B). Once the red tag has popped out and is showing, the turkey is cooked.

27 **a** Recommended tyre pressure is usually between 30 and 35 psi. As the temperature of the outside air cools, the pressure inside the tyre reduces which can accelerate tyre wear and damage, as more of the tyre's surface area comes into contact with the road.
b If the tyre pressure is too high above the nominal range 30 to 35 psi, then the tyre will not make sufficient contact with the road leading to slippage and decreased traction. An increased pressure will increase the volume of the tyre, causing it to become rounder and less of its surface area will make contact with the road.
c If the air in the tyre is under filled, then the tyre pressure is reduced which accelerates tyre wear and damage, as more of the tyre's surface area then comes into contact with the road. This increased contact with the tyre on the road can cause additional friction.

28 In theory, an ideal gas particle does not occupy any volume or collide/interact with other particles. However, a gas is matter consisting of atoms and they must have mass and some volume, no matter how small. Although there are large distances between gas particles, collisions between them can still occur.

Data drill

1 454 mL

2 480.3 mL

3 The theoretical value was 454 mL, while the average experimental value was higher, at 480.3 mL.

4 Go to Oxford Digital.

5 Data is not accurate. Temperature should be proportional to volume according to Charles' law, so the graph should be linear.

6 As the temperature increases, the volume of hydrogen gas produced increases.

Unit 2 Topic 1 review

Multiple choice

1	C	**2**	A	**3**	A	**4**	D	**5**	B
6	A	**7**	D	**8**	A	**9**	A	**10**	B
11	B	**12**	D	**13**	A	**14**	D	**15**	C

Short response

16 Octane is non-polar and can only interact with other octane molecules through dispersion forces. However, as it is a long chain, it can form more temporary dipoles, giving it a boiling point over 100°C. Water is a much smaller polar molecule that can form strong hydrogen bonds, in addition to dispersion forces and dipole–dipole attractions. Therefore, it has a boiling point of 100°C.

17 Hydrogen bonding is the strongest of the intermolecular forces, followed by dipole–dipole attractions, then dispersion forces.

18 a Go to Oxford Digital.

b H_2S: non-polar; CH_4: non-polar; NH_3: polar

c H_2S: dipole–dipole attractions and dispersion forces; CH_4: dispersion forces only; NH_3: dispersion forces, dipole–dipole attractions and hydrogen bonding

d NH_3; can undergo much stronger intermolecular forces of attraction between molecules than H_2S and CH_4, which require more energy to overcome

19 a The pressure of a gas above a liquid in a sealed container; when intermolecular forces are strong within a liquid, it is more difficult for the liquid to evaporate and the vapour pressure will be low

b i Diethyl ether has the lowest vapour pressure. It would therefore be expected to experience strong intermolecular forces, like dipole–dipole attractions. We would therefore expect it to be polar.

ii The vapour pressure of diethyl ether is not what we expected. From its formula $CH_3CH_2OCH_2CH_3$, it looks to be symmetrical and therefore, non-polar, so we would expect it to form weak intermolecular forces and therefore, have a high vapour pressure.

c Water has a slightly higher vapour pressure than octane. Despite experiencing hydrogen bonding compared to dispersion forces only between octane molecules, it is a very small molecule when compared to octane so will be able to evaporate at a slightly higher rate than octane.

20 a i Bent

ii Both are polar.

iii Water: dispersion forces, dipole–dipole attractions and hydrogen bonding; hydrogen sulfide: dispersion forces and dipole–dipole attractions

iv Hydrogen sulfide has a hydrogen atom bonded to sulfur (period 3), which is less electronegative and larger than oxygen (period 2). Water, consisting of hydrogen and oxygen (period 2), will contain highly polar bonds due to the greater electronegativity difference between atoms compared to sulfur. This will cause water to have a higher boiling point than hydrogen sulfide as it can interact through hydrogen bonding, a stronger intermolecular force than the dipole–dipole attractions of hydrogen sulfide. Stronger intermolecular forces require more energy to overcome.

b i Water: bent; hydrogen fluoride: linear

ii Both are polar.

iii Both interact through dispersion forces, dipole–dipole attractions and hydrogen bonding

iv Both contain highly polar bonds due to the great electronegativity difference between atoms compared to hydrogen. They can form hydrogen bonds, which requires a large amount of energy to overcome. However, due to the slightly larger size of the water, it will form more dispersion forces and have a higher boiling point (100°C compared to 30°C for HF).

c i Ammonia: pyramidal; methane: tetrahedral

ii Ammonia: polar; methane: non-polar

iii Ammonia: dispersion forces, dipole–dipole attractions and hydrogen bonding; methane: dispersion forces

iv Both methane and ammonia contain atoms from period 2. However, ammonia has a nitrogen atom from group 15 which is more electronegative than the carbon from group 14, allowing it to form hydrogen bonds, which require a larger amount of energy to overcome than the dispersion forces experienced by methane. This results in a higher boiling point for ammonia (–30°C) compared to methane (–160°C).

21 Chromatography uses two phases to separate mixtures, a mobile phase (a liquid or gas) and a stationary phase (usually a solid). Components of the sample are more attracted to either the mobile or stationary phase based on their affinity to each phase. Since like dissolves like, if the mobile phase is polar the more polar substances will be attracted to this phase and move along with it. The strength of the intermolecular forces between the components of the mixture and the phases causes the components to separate. The stronger the forces between the components and the stationary phase, the less the component moves along the chromatograph. The distance travelled by the components have a specific calculated R_F value which allows them to be identified if compared to standards. The purity of a substance can be determined by the number of components that are separated on the chromatograph.

22 a Go to Oxford Digital.

b Red component ($R_F = 0.91$)

c Both green and orange components ($R_F = 0.23$)

d They are not the same dye because they have different R_F values.

e The separation is mostly valid except for the green and orange components having the same R_F value of 0.23. To improve the separation, the mobile phase should be changed to see if greater separation can occur between the green and orange components.

23 a Neither the red nor yellow components of the dyes of sample A match sample B, so the dyes are not the same. For sample A, R_F(red) = 0.22 and R_F(yellow) = 0.47, whereas for sample B, R_F(red) = 0.11 and R_F(yellow) = 0.51.

b Green component

c Yellow component

d Yes, the components can be individually determined because each component has a different R_F value. However, they would need to be compared against standard dyes.

e Sample A and C are pure as they only contain two components. Sample B is not pure as it contains three different components.

24 a As temperature increases, the kinetic energy of the gas particles will increase. This causes the volume of the gas to increase as they push against the walls of a variable volume container with more force.

b As temperature increases, the kinetic energy of the gas particles will increase. In a fixed volume system, this causes the pressure of the gas to increase due to a greater proportion of higher energy collisions.

c If the number of particles is increased, then the volume would increase. The pressure would remain constant due to the compensatory increase in volume.

d If the number of particles is increased, then the pressure would increase. This is because the extra particles will cause more collisions with the container, increasing the pressure.

e If the pressure is increased, then the volume would increase. This is because the extra pressure will cause more collisions with the container, increasing its volume.

25 **a** 0.11 mol

b 0.11 mol

c There are no differences in the values. Either equation can be used to calculate the moles of gas.

d Since temperature and pressure are constant at STP, there is no need to use the ideal gas equation at STP. It is easier and quicker to use the molar volume formula as the same number of moles of any gas will have the same volume at STP.

26 **a** 250 mL

b 125 mL

27 **a** 50,177 g

b $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

$\Delta H = -890.6\,kJ$

c 152,875 g

d 3,094 kJ

28 **a** Temperature

b Pressure

c As the temperature of a gas increases, it is predicted the pressure of the gas will increase if the volume remains constant.

The results of the experiment are summarised in the table.

Temperature	Pressure
Room	Low
Medium	Medium
Hot	High

d Qualitative as no specific numbers were measured

e If volume is kept constant, then as the temperature is increased the pressure will increase. This is because, at higher temperatures, more particles have greater kinetic energy and collide with each other and the walls of the container with a greater force per unit area.

f Student answers will vary.

g The temperatures and pressures should be accurately recorded. The experiment could also be conducted in triplicate to assess precision.

Module 12: Aqueous solutions and molarity

Lesson 12.1 Properties of water

Worked example 12.1A

a Instant coffee powder

b Water

c Coffee

Real-world chemistry

1 In hot air, all particles have more kinetic energy. At the same time, the nearby water will also have more kinetic energy, and water molecules with more energy are much more likely to overcome intermolecular forces in the liquid and evaporate. Once water molecules have evaporated into warm air, they retain the high kinetic energy, and are much less likely to condense. Water molecules become the solute, and air the solvent.

2 When air is holding the maximum amount of water for a particular temperature, it is said to be "saturated" (100% humidity). Hotter air can hold more water vapour than cold air. As air cools, the water vapour saturates the air. The kinetic energy of the gas particles decreases. Once the water vapour in air is saturated due to cooler air being able to hold less water vapour than hotter air, further cooling will lead to condensation of the lower energy water vapour molecules. The water precipitates as rain.

3 The hydrogen bonding between water molecules is very strong due to its polarity and bent shape, and must be overcome to move them to a gaseous phase. This requires a relatively large amount of energy, giving water a high boiling point.

Check your learning 12.1

1 Bent structure with an angle of 104.5°

2 Oxygen is more electronegative than hydrogen, so it pulls on the electrons in the covalent bonds. This makes the oxygen atom partially negative, and the two hydrogen atoms partially positive. Together with the bent shape of water, the partial charges make the molecule polar. This allows it to form hydrogen bonds.

3 When water freezes to become solid ice, it forms hydrogen bonds in a lattice structure and the water molecules align to leave spaces within its structure. These increased spaces within the structure of solid ice means it is expanded compared to liquid water and is less dense.

4 Osmium has a higher density than beryllium and will occupy less volume for the same amount of mass.

5 Student answers will vary.

Lesson 12.2 Concentration

Skill drill

1 IV: water temperature; DV: mass of $CuSO_4$ dissolved in the water

2 Go to Oxford Digital.

3 The mass of $CuSO_4$ that dissolves in water increases as water temperature rises.

4 The experiment is valid as it measures what it had set out to measure.

Worked example 12.2A

3.6 M

Worked example 12.2B

2.49 M

Worked example 12.2C

200 mL

Worked example 12.2D

0.21 M

Worked example 12.2E

17.3 g L^{-1}

Worked example 12.2F

0.125 g

Real-world chemistry

1 For solid foods, concentrations are expressed in grams per 100 grams as this shows a comparison in actual mass. For liquid foods, concentrations are expressed as grams per 100 mL as they contain a majority of water, so g 100 mL^{-1} gives a better indication of the actual mass of a food in the liquid.

2 Concentration expressed in terms of moles is more scientifically accurate and represents the amount of substance actually present. To effectively convey nutritional information to the general

public, who may not be scientifically trained, a different measure is needed. Showing concentration as g 100 g^{-1}, or g 100 mL^{-1}, delivers that information in an easier to understand format.

Worked example 12.2G

20 ppm

Worked example 12.2H

0.25 mg

Worked example 12.2I

a 1,100 ppm
b 0.0032 $g\,L^{-1}$
c 150 $g\,L^{-1}$
d 3.1×10^{-4} M

Worked example 12.2J

1.19×10^{-4} M

Worked example 12.2K

0.175 M

Challenge

1 0.0024 M
2 The paint is unsafe for use.

Real-world chemistry

1 0.0027 M
2 N_2O increased the least (58 ppb).
3 Student answers will vary.

Check your learning 12.2

1 In a non-laboratory situation, it is more realistic to weigh the mass of a chemical dissolved in water.
2 **a** 16,000 ppm
  b 0.025 $g\,L^{-1}$
  c 0.84 M
3 **a** 6.0 M
  b 0.3 M
  c 6.55×10^{-4} M
4 **a** 580 ppm
  b 930 ppm
  c 4.69×10^{-6} ppm
5 Student answers will vary.
6 **a** The reliability is questionable because the concentration of the sample was extrapolated beyond the concentrations of the standards. To ensure that the sample concentration is reliable, the sample would need to be diluted so that the concentration "fitted" within the concentration range of the standards or the concentration of the standards could have been increased.
  b Student answers will vary.
7 Student answers will vary.

Lesson 12.3 Review: Aqueous solutions and molarity

Multiple choice

1 A **2** C **3** D **4** C **5** D **6** B
7 A **8** D **9** D **10** A **11** D **12** C

Short response

13 Density is measured as mass per volume. It is a unit of concentration because it indicates the amount of a substance in a certain volume.
14 Water molecules are polar. Because of its partial charges, water can become attracted to the charges present on the rod. The partially positive H atoms of the water molecules will become attracted to the negative charges from the rod, or the partially negative O atoms of the water molecules will become attracted to the positive charges from the rod (depending on the type of charge on the rod). This causes the water to become attracted to the rod and curve sideways towards it.
15 Water is a useful solvent because it can form strong hydrogen bonds or dipole–dipole attractions to other polar molecules. It is considered a "universal" solvent because it dissolves more substances than any other solvent.
16 24.7 mL
17 0.08 M
18 13 g
19 0.121 M
20 15.6 $g\,L^{-1}$
21 0.018 $g\,L^{-1}$
22 0.148 g
23 17 ppm
24 0.22 mg
25 2 L
26 Dilution involves the addition of solvent to reduce the concentration of a solution.
27 **a** 14 mL
  b 17 M
  c 100 mL
28 **a** 1.0 M means that each solution will contain 1 mol of particles in 1 L. Two of the solutes are ionic (NaCl and $Cu(NO_3)_2$) and will ionise to form a greater number of particles. For every 1 mol of NaCl, 2 mol of particles are formed producing a total concentration of all ions of 2.0 M. For every 1 mol of $Cu(NO_3)_2$, 3 mol of particles are formed, producing a total concentration of all ions of 3.0 M. Therefore, $Cu(NO_3)_2$ would be expected to have a higher boiling point than NaCl. As sucrose is a covalent molecule, it will have the lowest boiling point of the three substances, as it does not dissociate, and has a concentration of 1.0 M.
  b $Cu(NO_3)_2$
29 Sugar could be considered a solute as it is a dry powder of polar glucose molecules. Water could be considered a solvent as it is a polar liquid. Adding water to the glucose results in a sugar solution.
30 40 g per 100 g of water
31 When pure water is boiled, the bulk liquid consists only of water molecules and so it is easier to form bubbles of water vapour. When a solution is boiled, the solution contains some water and some solute. The presence of the solute means that there is less water present in a given volume and this makes it more difficult to form vapour, leading to a lowering of the vapour pressure. At the boiling point of pure water, the vapour pressure of the solution is too low for vapour to form under the surface of the liquid and so boiling does not occur. A higher temperature must be reached, to reach the vapour pressure necessary for boiling to occur.
32 Student answers will vary.

Data drill

1 IV: temperature of the liquid; DV: surface tension
2 Minimum temperature = 274.05 K ± 0.30%
  Maximum temperature = 304.96 K ± 0.86%
3 Go to Oxford Digital.
4 As temperature increases, surface tension decreases. Water has a much higher surface tension than methanol and ethanol at all temperatures.
5 56.6 $mM\,m^{-1}$

Module 13: Solubility and identifying ions in solution

Lesson 13.1 Identifying ions in solution

Worked example 13.1A

$Mn_3(PO_4)_2(aq) + 3Ba(CH_3COO)_2(aq)$
$\rightarrow Ba_3(PO_4)_2(s) + 3Mn(CH_3COO)_2(aq)$

Worked example 13.1B

a $Cu(NO_3)_2(aq)\ 2LiOH(aq)$
  $\rightarrow Cu(OH)_2(s) + 2LiNO_3(aq)$
b $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

Real-world chemistry

1 $Hg^{2}+(aq) + 2I^{-}(aq) \rightarrow HgI_2(s)$

2 A precipitate HgI_2 is formed and can be removed by filtration.

3 Student answers may vary.

Challenge

$$CACO_3(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Challenge

Student answers will vary.

Skill drill

1 The solubility of lead will vary in the presence of different ions, with the greatest solubility in ethanoates, nitrates and iodides.

2 To test the solubility of lead in solutions of different ions

3 Student answers will vary.

Check your learning 13.1

1 For slightly soluble substances, some of the chemical dissolves in the solvent, but most remains as a solid. The ions are so strongly attracted to one another that they are less likely to dissociate or dissolve.

2 Ionic substances are substances that consist of ions held together by electrostatic forces. These opposite charges are the reason why ionic bonds are so strong. An ionic bond is an electrostatic attraction between a cation and an anion.

3 **a** Soluble; all nitrate salts are water soluble

b Insoluble; most hydroxides are water insoluble

c Insoluble; most chloride salts are water soluble but $HgCl_2$ is an exception

d Soluble; most chloride salts are water soluble

e Insoluble; most sulfate salts are water soluble but $BaSO_4$ is an exception

f Soluble; all group 1 elements are water soluble

g Soluble; all substances containing NH_4^+ are water soluble

h Soluble; all group 1 elements are water soluble

i Insoluble; most hydroxides are water insoluble

j Insoluble; most hydroxides are water insoluble

k Soluble; most phosphates are water insoluble but $(NH_4)_3PO_4$ is an exception

l Soluble; all group 1 elements are water soluble

4 **a** Precipitation reaction as solid $PbCl_2$ precipitate is formed

b Not precipitation reaction as all products are water soluble

c Not precipitation reaction as all products are water soluble

5 **a** $Co^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CoCO_3(s)$

b $3BaO(aq) + 2Al(NO_3)_3(aq) \rightarrow Al_2O_3(s) + 3Ba(NO_3)_2(aq)$

$2Al^{3+}(aq) + 3O^{2-}(aq) \rightarrow Al_2O_3(s)$

c $3Fe(NO_3)_2(aq) + 2Na_3PO_4\,(aq) \rightarrow 6NaNO_3(aq) + Fe_3(PO_4)_2(s)$

$3Fe^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$

d $3Pb(NO_3)_2(aq) + Al_2(SO_4)_3(aq) \rightarrow 3PbSO_4(s) + 2Al(NO_3)_3(aq)$

$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

e $2AgNO_3(aq) + MgI_2(aq) \rightarrow Mg(NO_3)_2(aq) + 2AgI(s)$

$Ag^{+}(aq) + I^{-}(aq) \rightarrow AgI(s)$

Knowledge utilisation

6 **a** Because of the visual appearance of the precipitate

b Weigh the precipitate after the completion of the reaction and filtration of the solution

c Student answers will vary.

7 No, ions can be spectator ions.

Lesson 13.3 Factors affecting solubility

Worked example 13.3A

Not all of the solute will dissolve (10 g undissolved).

Worked example 13.3B

31 g

Worked example 13.3C

65 g

Worked example 13.3D

9.67 g 100 g^{-1}

Challenge

An increase in temperature causes an increase in kinetic energy, causing substances to become more or less soluble in water, depending on whether they are polar or non-polar. The solubility of Li_2SO_4 in water decreases with increasing temperature, as its dissolution is an endothermic process.

Real-world chemistry

1 Plasma is made of up >90% water, allowing it to act as a solvent by forming intermolecular forces with solutes.

2 Small amounts of gases dissolve in the blood plasma and are transferred around the body. Carbon dioxide is more soluble than oxygen. Other components of blood plasma affect its ability to act as a solvent. These include the presence of ionic substances and a pH that is maintained at 7.4.

Check your learning 13.3

1 An increase in temperature causes an increase in kinetic energy of the molecules. Increased kinetic energy alters the intermolecular forces, enabling the salts to dissolve more in water.

2 The attraction between the substance and the polar water molecules may be weak and the attractive forces within the substance holding one particle to another may be very strong.

3 Ionic substance, polar molecular substance, non-polar molecular substance

4 **a** **i** 62 g 100 g^{-1} water

ii 31 g 100 g^{-1} water

iii 20.1 g 30 mL^{-1}

iv 1.56 g

b Student answers will vary.

5 The line on solubility diagrams represents the saturation point of a solution at a given temperature. A line in the upwards direction indicates that solubility increases as temperature increases.

6 **a** 50 g

b 62 g

c 75 g

7 The solubilities would be considerably lower or non-existent compared to that of the polar gases.

8 Honey at the bottom, milk, then baby oil

Lesson 13.4 Review: Solubility and identifying ions in solution

Multiple choice

1 C **2** D **3** D **4** A **5** A

6 D **7** B **8** C **9** A

Short response

10 If a molecular substance has the same properties as water, it will dissolve better in water. Often molecular substances are non-polar and contain only dispersion forces, which allows them to interact with neighbouring molecules. Because dispersion forces create a temporary dipole, these molecules are still able to interact with the solvent water molecules, but not to the same extent as a polar molecule, making them much less soluble in water.

11 The cations and anions in the ionic reactions swap partners to become different substances. Sometimes, this

produces a salt that is not water soluble, i.e. solid. An example is the formation of silver chloride: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$.

12 Measurements in $g\,100\,g^{-1}$ water can be easily converted to percentages.

13 Sucrose has a much higher solubility than that of glucose at room temperature. More OH groups on sucrose are accessible for bonding to polar water molecules, thereby increasing its solubility in water.

14 Because salts consist of ions, they contain whole positive and negative charges, which are electrostatically attracted to one another. This is a strong type of bonding because it involves whole charges, as opposed to partial charges. The interaction between molecular substances (like covalently bonded SO_4^{2-} molecules) and water is weaker than the interaction between ionic substances and water as they involve partial positive and partial negative charges.

15 **a** $2NaNO_3(aq) + BaSO_4(s)$
   b $2NaNO_3(aq) + Pb(OH)_2(s)$
   c $3KCl(aq) + Al(OH)_3(s)$

16 **a** 187.5 g 100 g⁻¹ water
   b 23 g 100 g⁻¹ water
   c 30°C
   d 85 g 40 mL⁻¹ water
   e 10 g 9 mL⁻¹ water

17 Weigh out the required mass of NH_4NO_3. Then, add it to water that has been heated to a temperature at which >187.5 g of KBr is soluble, such as 30°C. The solution would then need to be cooled back down to 20°C.

18 While the use of a hot plate for heating is a common and effective method to create a supersaturated solution, it is not the only method available. Therefore, the claim that a supersaturated solution cannot be made in a laboratory without a hot plate is not accurate.

19 Student answers will vary.

20 Student answers will vary.

21 Student answers will vary.

22 Solution B is the higher density sucrose and solution A is the lower density salt water.

Data drill

1 0.8 g

2 18°C

3 Go to Oxford Digital.

4 As temperature increases, the solubility of carbon dioxide in water decreases.

5 1.2 g

6 The concentration of the solute in the solution will vary, which will affect the measured solubility. This can lead to inaccurate determinations of solubility limits.

Module 14: Acids and bases

Lesson 14.1 The Arrhenius model of acids and bases

Challenge

$H_3PO_4(aq) \rightarrow H^+(aq) + H_2PO_4^-(aq)$
$H_2PO_4^-(aq) \rightarrow H^+(aq) + HPO_4^{2-}(aq)$
$HPO_4^{2-}(aq) \rightarrow H^+(aq) + PO_4^{3-}(aq)$

Real-world chemistry

1 Most covalent compounds are unlikely to ionise due to the stable and very strong electron-sharing bonds. Because these bonds are so stable, they are unlikely to gain or lose electrons to become ions.

2 Only ionic compounds that can accept hydrogen ions or donate hydroxide ions in solution act as bases.

Check your learning 14.1

1 **a** An acid is a chemical compound that donates a hydrogen ion. Acids are sour tasting, can be corrosive, can conduct electricity, and can create solutions with a high H^+ concentration.
   b A base is a chemical compound that accepts a hydrogen ion. Bases are bitter to taste and feel slippery, are corrosive and can produce chemical burns on the skin, and can create solutions with a lower concentration of H^+ ions than acids.

2 **a** The group of chemicals that can produce a high concentration of hydrogen ions in an aqueous solution
   b The group of chemicals that can produce a high concentration of hydroxide ions in an aqueous solution

3 Go to Oxford Digital.

4 A dilute strong acid will completely dissociate its H^+ ions into solution but will have fewer acid molecules present in the given volume of solution than a concentrated acid. A concentrated weak acid has only partially dissociated its H^+ ions into solution but has more acid molecules present in the given volume of solution than a dilute acid.

5 The Arrhenius model requires all acids and bases to be able to produce hydrogen ions or hydroxide ions respectively. However, some substances (like NH_3) can still act as a base despite not containing hydroxide ions. Brønsted and Lowry independently suggested that acids release hydrogen ions and bases bond with hydrogen ions.

Lesson 14.2 The pH scale

Worked example 14.2A

2

Worked example 14.2B

0.84

Challenge

The "free" aluminium ions react with pigments in the petals to produce pink-red flowers. Therefore, aluminium ions need to be present in the soil to produce pink-red flowers on the hydrangea. Even if the soil is acidic, if there are no aluminium ions to form complexes that turn the flowers pink-red, we would expect the petals to remain blue.

Skill drill

1 Qualitative

2 The type (strength) of acid/base

3 To determine the relative strength of acids and bases

4 A is a strong acid, B is weak base, C is a weak acid and D is a strong base.

Real-world chemistry

1 1

2 Antacids are substances that neutralise stomach acid by reacting with stomach acids to produce a neutral ionic salt and water, which prevents them from damaging the cells of the stomach or oesophagus.

Check your learning 14.2

1 **a** Acidic
   b Neutral
   c Basic
   d Acidic

2 Indicators are dyes that have different coloured forms when they are protonated (have a H^+) and unprotonated (do not have a H^+), and they change colour in the presence of high and low hydrogen ion concentrations.

3 **a** Bromophenol blue is yellow, methyl red is red and phenolphthalein is colourless.
   b Bromophenol blue is blue, methyl red is yellow and phenolphthalein is red.

4 Capsicum, onions, avocado

5 If the concentration of hydrogen (H^+) ions increases, then the pH decreases to the lower range of acidic pH values, that is, pH<7.

6 Neutral substances have a pH of 7; anything slightly higher than 7 is considered basic. Therefore, the pH of 7.4 would be considered basic, not neutral.

Lesson 14.3 Review: Acids and bases

Multiple choice

1	C	**2**	C	**3**	D	**4**	D	**5**	B
6	D	**7**	C	**8**	D	**9**	B	**10**	A

Short response

11 Ionisation is the process by which an atom or molecule gains or loses electrons, resulting in the formation of positively charged cations or negatively charged anions.

12 Completely dissociate into H^+ cations and corresponding anions in the presence of water

13 Completely dissociate into their OH^- anions and corresponding cations in the presence of water

14 The Arrhenius model states that all acids produce hydrogen ions and all bases produce hydroxide ions.

15 H_2SO_4 is a strong acid because it completely dissociates into SO_4^{2-} ions and H^+ ions in an aqueous solution and therefore produces many H^+ ions in solution. The concentration of an acid depends on the number of acid molecules in solution. A concentrated solution of H_2SO_4 has greater molarity and concentration of H^+ ions than that of a dilute acid.

16 HCl completely dissociates into solution to produce H^+ and Cl^- ions, which can carry charge.

17 Go to Oxford Digital.

18 Both are weak acids, meaning that they partially ionise to form hydrogen ions (H^+) in water. However, a concentrated weak acid solution contains more molecules of acid per volume of water than a dilute weak acid solution.

19 Neutralisation results in the formation of water and a neutral salt, which have a pH of 7, which is considered neutral. For example:
$HCl + NaOH \rightarrow NaCl + H_2O$.

20 **a** 4
b 7
c 11

21 A strong base completely dissociates to produce high concentrations of OH^- ions; the solution will be considered a strongly basic solution and will therefore have a higher pH value.

22 Varying indicator types give differing ranges of results when pH is reported. Therefore, the choice of indicator needs to be applicable for the range of pH being tested. For confirmation of an acid or base, two or more indicators can be used to verify the pH of a solution.

23 As the solution is opaque, it may be difficult to observe a colour change from an indicator, especially if the change is subtle. Therefore, an indicator or method of measuring pH needs to provide a clear and distinct colour change visible through the cloudy solution.

24 The addition of water will dilute the acid concentration, but the degree of dissociation of the ions will remain unchanged.

25 In the Arrhenius model of acids and bases, bases have high concentrations of hydroxide ions (OH^-). If $Ca(OH)_2$ could dissociate in water and be soluble, then the concentration of OH^- ions would align with the Arrhenius model of increased concentration of OH^- ions in solution being associated with bases.

26 As strong acids dissociate completely into its ions, whereas a weak acid partially dissociates, equal concentrations and volumes of a strong acid would contain more hydrogen ions than a weak acid. More base would be needed to neutralise it than the same volume of 0.1 M weak acid.

27 Bulimia involves recurrent episodes of binge eating followed by purging, often through self-induced vomiting. Vomiting brings stomach acid into the mouth and throat, which can cause damage to the soft tissue of the mouth and gums. These areas of damaged cells can then become infected and cause ulceration. The acid weakens and wears down tooth enamel, leading to dental erosion and an increased risk of cavities.

28 pH is calculated from the concentration of hydrogen ions, using the formula: $pH = -\log_{10}[H^+]$ where $[H^+]$ is the concentration of H^+ (M). Therefore, a $[H^+]$ of $0.001 = 10^{-3}$ M gives a pH of 3, and a hydrogen concentration of $0.0001 = 10^{-4}$ M gives a pH of 4.

29 Go to Oxford Digital.

Data drill

1 $[H^+]_A = 1.00\times10^{-7}$ M
$[H^+]_B = 1.91\times10^{-9}$ M

2 Acid A: bromothymol blue
Acid B: phenolphthalein

3 As the equivalence point is approached, the pH of the reaction system increases.

4 The initial pH of Acid A is lower than that of Acid B. As the reaction proceeds and approaches the equivalence point, the pH of the reaction system increases more gradually for Acid A and the base than it does for Acid B and the base.

5 Acid A is a strong acid and Acid B is a weak acid.

6 Assuming the same volume of acid and concentration of each acid was used, Acid A would be a stronger acid as it has a low pH of 1 measured at the beginning of the titration before any base is added. A low pH indicates a high H^+ ion concentration. In contrast, Acid B has an initial pH of about 3, indicating it is a weaker acid.

Module 15: Reactions of acids

Lesson 15.1 Acid–base neutralisation reactions

Worked example 15.1A

$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Check your learning 15.1

1 **a** The starting chemicals or substances in a chemical reaction
b The final chemicals formed from reactants during a chemical reaction
c The reaction between an acid and a base to produce water and a metal salt
d A solution that is neither acidic nor basic and has a pH of 7

2 An indicator is added to the acid, then the acid is slowly neutralised by dripping (or titrating) a base of known concentration into it. The equivalence point is the point in the titration when the acid and base are present in equivalent amounts, and the reaction is complete. The point at which the indicator changes colour is the end point.

3 To determine the concentration of an acid or a base

4 **a** $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
b $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
$2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$
c $2CHOOH(aq) + Ba(OH)_2(aq) \rightarrow Ba(CHOO)_2(aq) + 2HO(l)$
$2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$

5 **a** $Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(a) + 2H_2O(l)$
$Ba^{2+}(aq) + 2OH^-(aq) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2H_2O(l)$
b The conductivity of the mixture decreases as the reaction approaches equivalence point as the ions are consumed to form products.

6 No. The solution will pull towards the dissociation of acidic ions and remain acidic, therefore not being able to neutralise the reaction with a weak base.

7 Student answers will vary.

Lesson 15.2 Other acid reactions

Worked example 15.2A

$2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$

$2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

Real-world chemistry

1 $2H^+(aq) + Fe(s) \rightarrow Fe+(aq) + H_2(g)$
$Ca^{2+}(aq) + CO_3^{2-}(aq) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$

2 $2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$
$2NO_2(g) + H_2O(l) \rightarrow 2H^+(aq) + NO^{3-}(aq) + NO^{2-}(aq)$

Challenge

Soda ash neutralises the strong acid sulfuric acid by producing the neutral salt sodium sulfate, carbon dioxide and water, which are all neutral substances. The carbonate ions from sodium carbonate react with the hydrogen ions from sulfuric acid to form water and carbon dioxide, while the sodium ions combine with sulfate ions to form the sodium sulfate. This reaction effectively neutralises the acidic properties of sulfuric acid.

$Na_2CO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + CO_2(g) + 2H_2O(l)$

Skill drill

1 Calcium carbonate: gas bubbles and calcium carbonate solid breaking down
Sodium hydroxide: sodium hydroxide solid will dissolve into solution
Zinc metal: zinc solid will become pitted and form gas bubbles
Magnesium hydroxide: magnesium hydroxide will dissolve into solution

2 $CaCO_3(s) + 2HCl(aq) \rightarrow CaaCl_2(aq) + H_2O(l) + CO_2(g)$
$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
$Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

3 Student answers will vary.

Check your learning 15.2

1 **a** acid + base → metal salt + water
b acid + metal → metal salt + hydrogen gas
c acid + carbonate → water + metal salt + carbon dioxide gas

2 hydrogen ion + hydroxide ion → water
hydrogen ion + metal → metal ion + hydrogen
hydrogen ion + carbonate ion → carbon dioxide + water

3 **a** $H_2SO_4(aq) + Mg(OH)_2(s) \rightarrow MgSO_4(aq) + 2H_2O(l)$
$2H^+(aq) + Mg(OH)_2(s) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
b $H_2SO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + H_2(g)$
$2H^+(aq) + Mg(s) \rightarrow Mg_2+(aq) + H_2(g)$
c $H_2SO_4(aq) + MgCO_3(s) \rightarrow MgSO_4(aq) + CO_2(g) + H_2O(l)$
$2H^+(aq) + MgCO_3(s) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$

4 The pop test is a test to determine if a gas produced is hydrogen. The lit splint test is used to determine if a gas produced is carbon dioxide.

5 Carbon dioxide reacts with water in the ocean to form carbonic acid.
$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
Carbonic acid then reacts with calcium carbonate found in shells of sea organisms to form calcium bicarbonate salt, carbon dioxide and water.
$2H_2CO_3(aq) + CaCO_3(s) \rightarrow H_2O(l) + Ca(HCO_3)_2(aq) + CO_2(g)$
This means that CO_2 eventually dissolves the calcium carbonate-containing shells of sea organisms, making their shells weaker and negatively affecting their survival.

Lesson 15.3 Review: Reactions of acids

Multiple choice

1	B	**2**	C	**3**	B	**4**	D	**5**	C
6	D	**7**	B	**8**	D	**9**	C		

Short response

10 A neutralisation reaction is the reaction between an acid and a base (i.e. the reactants) to produce water and a metal salt (i.e. the products).

11 Salt and hydrogen gas

12 Metal salt, carbon dioxide and water

13 **a** Student answers will vary.
b Student answers will vary.

14 $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
$2OH^-(aq) + 2H^+(aq) \rightarrow 2H_2O(l)$

15 Go to Oxford Digital.

16 **a** $Ca(OH)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l)$
$Ca^{2+}(aq) + 2OH^-(aq) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s) + H_2O(l)$
b $LiOH(aq) + HCl(aq) \rightarrow LiCl(aq) + H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
c $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$
d $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
$2OH^-(aq) + 2H^+(aq) \rightarrow 2H_2O(l)$
e $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca_2+(aq) + CO_2(g) + H_2O$
f $Pb(s) + 2HI(aq) \rightarrow PbI_2(s) + H_2(g)$
$Pb(s) + 2H^+(aq) + I^-(aq) \rightarrow PbI_2(s) + H_2(g)$
g $3Zn(s) + 2H_3PO_4(aq) \rightarrow Zn_3(PO4)_2(aq) + 3H2(g)$
$3Zn(s) + 6H^+(aq) \rightarrow 3Zn^+(aq) + 3H_2(g)$

17 Add sodium bicarbonate to raise the pH of the pool to a much safer and slightly basic pH of 7.5, since it will react with hydrogen ions in the pool water to form water and carbon dioxide

18 Sodium bicarbonate could be used as a mild base to neutralise the acids from bee stings and ant bites. It will react with the acid to produce a sodium salt, carbon dioxide and water.

19 A lit splint cannot be used to detect oxygen as it would continue to burn, giving no obvious sign of a gas being present, i.e. a lit splint requires oxygen to keep burning.

20 **a** Mostly iron and limited amounts of the other nutrients
b Magnesium and lower the amounts of nitrogen

Data drill

1 $H_2O(l) + LiX(aq)$

2 Phenol red would be an appropriate for the reaction between Acid 1 and LiOH, but not for Acid 2 and LiOH.

3 Acid 1 is a strong acid. Acid 2 is a weak acid.

4 Acid 2 has a greater concentration of hydrogen ions because it required more LiOH to neutralise Acid 2 during titration. The exact concentration of the original acid can be calculated from the information given.

5 If a weak base was used instead of the strong base LiOH, the vertical section of the graph would be shorter for the titration with Acid 1 and the equivalence point would be in the acidic pH range, and the vertical section of the graph would be very small for titration with Acid 2 and the equivalence point would occur at a neutral pH.

Unit 2 Topic 2 review

Multiple choice

1	C	**2**	B	**3**	C	**4**	C	**5**	B
6	D	**7**	B	**8**	B	**9**	C	**10**	C
11	D	**12**	C	**13**	D	**14**	C	**15**	D

Short response

16 You would need to first weigh out the required mass of NaCl, then, add it to water that has been heated to temperature at which >37.5 g of NaCl is soluble. The solution could then need to be cooled back down to 70°C.

17 **a** Strong acid: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Weak acid: $CH3COOH(aq) \rightarrow CH3COO^-(aq) + H^+(aq)$

b A weak base partially ionises in water, producing fewer OH^- ions. A dilute base has a low concentration of the base in solution, regardless of its ionisation properties.

18 Calcium chloride forms strong ionic bonds in its lattice structure and a large amount of energy is required to break these bonds and convert the solid into a liquid. However, it readily dissolves in water because it is a polar substance. The charged ions can form ion–dipole interactions with the water molecules and the ions can separate from the ionic lattice.

19 Ammonia is highly soluble in water at lower temperatures as it is a gas at room temperature. However, as temperature increases, its solubility decreases. Ammonium chloride is highly soluble in water at higher temperatures, with solubility increasing with temperature due to its ionic nature.

20 **a** NO is polar and forms dipole–dipole attractions with water molecules, allowing it to more readily dissolve in water than oxygen at any given temperature. Oxygen is a non-polar diatomic molecule that does not readily interact with water due to only being able to form weak intermolecular dispersion forces.

b CH_4 has a slightly higher solubility in water than N_2 at all temperatures. Both gases are non-polar and can only form weak dispersion forces with water. However, methane has a slightly higher solubility as it has more electrons available to form temporary dipoles, and therefore dispersion forces, than nitrogen.

21 **a** 97.5 g

b 45°C

22 **a** $2Li_3PO_4(aq) + 3ZnI_2(aq) \rightarrow 6LiI(aq) + Zn_3(PO_4)_2(s)$

$3Zn^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Zn_3(PO_4)_2(s)$

b $MnCl_2(aq) + 2NaOH(aq) \rightarrow 2NaCl(aq) + Mn(OH)_2(s)$

$Mn^{2+}(aq) + 2OH^-(aq) \rightarrow Mn(OH)_2(s)$

23 **a** Student answers will vary.

b Student answers will vary.

24 **a** 3.33

b 8.67

25 **a** $MgCO_3(s) + H_2SO_4(aq) \rightarrow H_2O(l) + MgSO_4(aq) + CO_2(g)$

b $HNO_3(aq) + Ca(OH)_2(s) \rightarrow H_2O(l) + Ca(NO_3)_2(aq)$

26 0.148 M

27 1.43 M

28 0.06 L

29 **a** In general, the solubility curves of strong and weak acids would show an increase in solubility with an increase in temperature. The strong base would show a higher solubility than the weak base at any given temperature due to strong bases more readily dissociating in water than weak bases.

b Generally, strong bases have a much higher solubility in water compared to weak bases due to their complete dissociation into ions and consequently stronger attraction to the polar water molecules. In contrast, a weak base only partially ionises and have fewer ions available to form attractions with water molecules, giving rise to a lower solubility.

c When the temperature of the water increases, there is more heat energy available to overcome the intermolecular forces holding the molecules of base together. This allows them to instead form attractive forces with water molecules, increasing their solubility.

30 **a** 2.81 g

b If a reactant is in excess, it can cause an excess of hydrogen or hydroxide ions which will impact the pH of the solution and move it away from neutral.

c The pH of the solution will increase (become basic).

d To test if the solution has a reactant in excess, a universal indicator or pH meter could be used. If neutralisation has been achieved, then the test should show a neutral pH, but if a reactant is in excess, it will be acidic (if excess acid) or basic (if excess base).

Module 16: Rates of chemical reactions

Lesson 16.1 Factors affecting reaction rates

Worked example 16.1A

Increasing the concentration of HCl will increase the number of particles in the set volume, reducing the space between them and allowing a greater frequency of collisions. A larger surface area of solid Mg would also provide more sites for collision to occur, increasing the frequency of collisions. Increasing temperature will increase the kinetic energy of all particles, providing more particles with sufficient energy to react. All of these will increase the frequency of collisions, the proportion of successful collisions and therefore, the rate of reaction.

Worked example 16.1B

a 0.7 g min^{-1}

b 0.19 g min^{-1}

c 0.2 g min^{-1}

d 0.05 g min^{-1}

Worked example 16.1C

450 mL

Skill drill

1 Carbon dioxide gas is formed in the reaction. The gas escapes the reaction vessel and is lost to the surroundings as it is produced, so the mass recorded decreases.

2 0.73 g

3 The last repeat of the experiment was the most accurate as it is within 0.06 g of the theoretical value (0.726 g). The other two experiments are less accurate than this with a difference of 0.016 g and 0.026 g.

4 High precision as the values gathered are within 0.02 g or 2.9% of each other

Real-world chemistry

1 The increased surface area of sugar dust provides more surface for collisions with oxygen molecules, increasing the frequency of collisions, proportion of successful collisions and thus, the rate of the reaction.

2 Regular cleaning of vents or upgrading the ventilation system to remove dust, remove any ignition sources

3 Student answers will vary.

Check your learning 16.1

1 Reactants must collide with the correct orientation, or in the correct position to successful react. Reactants must also collide with sufficient energy (activation energy) to rearrange and form products.

2 Increased temperature, pressure, concentration, surface area and catalysts

3 **a** 1.23 L

b Go to Oxford Digital.

c 0.135 L min^{-1}

d The 0.540 L volume of CO_2 gas measured by the students is less than half of the true value at 1.23 L.

e Student answers will vary.

f Increasing the volume of HCl used, as HCl is already in excess.

4. Have the car parts wrapped in suitable plastic to shield the parts from the rain or take the car parts out of the rain and place them inside the building for protection
5. Student answers will vary.
6. Reactions occur at the surface of a solid reactant. Therefore, the greater the surface area, the greater the number of collisions. Smaller solids have larger surface areas. Therefore, the rate of dissolving the granulated sugar in a hot tea or coffee is higher than that for sugar cubes because of the increased surface area of the granulated sugar compared to that of the sugar cubes.

Lesson 16.3 Energy profile diagrams

Worked example 16.3A

Go to Oxford Digital.

Skill drill

1. 0.93 L
2. There is a difference of 0.37 L or 40% error between the theoretical and actual yields. The data would be considered inaccurate.
3. Use methods that maximise the collection of the hydrogen gas and prevent its escape to the surroundings, consider the purity of reactants and controlling laboratory conditions, particularly temperature
4. Conducting the same experimental procedures multiple times, ideally under identical conditions, to see if similar results can be consistently achieved; helps to assess and improve precision, and identify random errors and minimise their impact
5. Repeated experiments can help identify systematic errors because the errors will consistently appear across all trials. If every repeated measurement deviates from the true value by the same amount, this indicates a systematic error. Once identified, systematic errors can often be corrected. It also reduces the impact of random errors on the overall dataset. By averaging the results of multiple trials, the random variations tend to cancel out, leading to a more accurate estimate of the true value.

Real-world chemistry

1. $2Fe_2O_3(s) \rightarrow 4Fe(s) + 3O_2(g)$
   $CH_3COONa{\bullet}3H_2O(aq)$
   $\rightarrow CH_3COONa(s) + 3H_2O(l)$
2. Go to Oxford Digital.
3. They absorb heat energy from their surroundings, leading to a decrease in the temperature of the immediate environment.

Check your learning 16.3

1. The minimum amount of energy required in a collision for a reaction to occur; reaction rates depend on the size of the activation energy. If the energy contained within the bonds of the reactants is relatively small (i.e., the bonds are weak), a smaller activation energy is required. Strong bonds contain larger amounts of energy, are harder to break apart and require a larger activation energy.
2. The peak of the curve on the energy profile diagram where energy is at its greatest and the arrangement of atoms is the most unstable; at this point, bonds are both breaking and forming into products.
3. Go to Oxford Digital.
4. Go to Oxford Digital.
5. **a** Go to Oxford Digital.
   b $E_a = 2{,}632\,kJ\,mol^{-1}$
   $\Delta H = +802.3\,kJ\,mol^{-1}$
   Original reaction: exothermic
   Reverse rection: endothermic

Lesson 16.4 Catalysts

Worked example 16.4A

Go to Oxford Digital.

Worked example 16.4B

Go to Oxford Digital.

Real-world chemistry

1. A catalysis reaction involves intermolecular forces (dispersion forces, dipole–dipole attractions and hydrogen bonding).
2. The active site is the region of the enzyme in which substrates bind via intermolecular interactions to form a more stable transition state. It allows the formation of a new product or breaks down the single substrate into multiple products.
3. Without enzymes, the digestive system would be unable to break down food into its absorbable components, leading to severe nutritional deficiencies, energy shortages, compromised immune function, and significant gastrointestinal discomfort.

Check your learning 16.4

1. A catalyst provides a reaction pathway with a lower activation energy and increases the rate of reaction without participating in the chemical reaction.
2. Go to Oxford Digital.
3. A homogeneous catalyst is a catalyst that is in the same state as the reactants, and a heterogeneous catalyst is a catalyst that is in a different state from the reactants.
4. Go to Oxford Digital.
5. **a** $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$
   $CH_2CH_2 + H_2O \rightarrow CH_3CH_2OH$
   b Reaction 1: enzymes of the yeast (e.g. zymase) and bacteria
   Reaction 2: phosphoric acid
   c Reaction 1: uses a renewable source, less energy and lower temperatures but produces CO_2 as waste
   Reaction 2: forms no waste products but uses a non-renewable source, and requires a lot of energy for high heat and high pressure conditions
6. Student answers will vary.

Lesson 16.5 Review: Rates of chemical reactions

Multiple choice

1	C	**2**	A	**3**	D	**4**	A	**5**	B
6	A	**7**	C	**8**	A	**9**	A	**10**	B

Short response

11. A type of protein that catalyses or accelerates biochemical reactions
12. When a reaction is heated, energy is transferred to the reactants, making them move faster. More particles have energy greater than the activation energy, so there are more successful collisions and an increased rate of reaction.
13. Go to Oxford Digital.
14. The activation energy of the catalysed pathway is lower than the activation energy of the uncatalysed pathway.
15. **a** A timer could be used to record the time taken for the cloudy precipitate to form and cover the black cross.
    b Student answers will vary, but could suggest increasing temperature and increasing the concentration of the reactants.
16. **a** Record the temperature increase over time.
    b Student answers will vary, but could suggest adding a catalyst, increasing the concentration of reactants and increasing the temperature.
    c Cooler temperatures reduce the kinetic energy of molecules, which reduces the number of particles with sufficient energy to react. This reduces the proportion of successful collisions and slows down decomposition.
19. They make it more likely that collisions will result in a successful reaction by providing an alternative pathway with a lower activation energy, increasing the rate of reaction.

20 **a** Go to Oxford Digital.
  b Go to Oxford Digital.
  c Go to Oxford Digital.

21 **a** Record the mass lost from the reactant (Mg(s)) over time
  b Increased surface area of the cube and increased HCl concentration
  c There are too many variables.
  d It is possible; the rate of reaction for the increased surface area (cube of Mg), and for the increased concentration of 2 M HCl, should give a higher rate of reaction than that of the Mg ribbon and 1 M HCl, respectively.

22 Incorrect; if reactants do not collide with sufficient energy, even a correct orientation does not result in a successful collision.

23 **a** Endothermic; the products have more energy than reactants, and energy is absorbed from the environment as heat
  b Quantity I
  c When the reaction is switched from endothermic to exothermic, I < II. When the reaction is switched from exothermic to endothermic, I > II.
  d Go to Oxford Digital.

Data drill

1 1 to 2 minutes: 0.01 L min^{-1}
4 to 5 minutes: 0.006 L min^{-1}
9 to 10 minutes: 0.005 L min^{-1}

2 Approximately 11 minutes

3 The graph for change in mass of magnesium would show a decrease in mass over time. This is because it is a reactant being consumed in the reaction.

4 The rate of reaction decreases as the reaction proceeds. The highest rates occurred in the first 20% of the reaction before slowing until the reaction was complete.

5 0.07 L

Unit 2 Topic 3 review

Multiple choice

1 A	**2** B	**3** C	**4** B	**5** D
6 A	**7** C	**8** A	**9** C	**10** C
11 D	**12** A	**13** D	**14** C	**15** D

Short response

16 **a** Go to Oxford Digital.
  b –184 kJ mol^{-1}
  c Go to Oxford Digital.

17 **a** The curve is skewed to the right as more particles have an energy greater than the activation energy.
  b The curve is skewed to the left as less particles have an energy greater than the activation energy.
  c The curve will be approximately the same shape but have a greater area under the curve (including a higher peak) as more particles are present, including the quantity of particles with energy greater than the activation energy.
  d The curve will remain the same as the energy of the particles will remain the same.
  e The curve will remain the same, but the activation energy will be lower and thus, there will be more particles with energy greater than the activation energy.

18 **a** The kinetic energy of the particles will increase, and the frequency and proportion of successful collisions will increase, which will increase the rate of the reaction.
  b If the pressure is increased, the space between molecules will decrease and the concentration of reactant molecules in the given volume will increase. This will lead to an increased frequency and proportion of successful collisions which will increase the rate of the reaction.
  c If a catalyst is added, this will provide a pathway with a lower activation energy. This will increase the rate of reaction.

19 **a** $CaCO_3$
  b 2.27 L
  c Green; there has been no change to the concentration of the reactants, so the same amount of product will be formed. Only the steepness of the curve will be greater.
  d An increase in the amount of both reactants; this will not change the rate of the reaction and will just result in more product being formed.

20 **a** Student answers will vary.
  b Measuring the loss of reactants (glucose) over time or measuring the amount of products (CO_2) formed
  c Yeast contains enzymes that act as a catalyst for the fermentation reaction, providing an alternative pathway for the reaction to occur in which a lower activation energy is required. As a result, the reaction can occur more quickly.

21 **a** 0.118 g
  b 0 to 2 seconds
  c **i** Go to Oxford Digital.
    ii The rate of reaction will increase for a reaction because it increases the surface area of the iron, providing more surfaces for collisions to occur. This will increase the frequency of collisions, and thus increase the proportion of successful collisions.
  d **i** Go to Oxford Digital.
    ii The number of nitric acid particles will decrease, reducing the frequency of collisions. The proportion of successful collisions will decrease and the rate of reaction will decrease.

22 **a** The number of particles in the reaction system with specific amounts of energy
  b Go to Oxford Digital.
  c A decrease in temperature will cause particles in the system to have less energy. Fewer particles will have the required activation energy for successful collisions and this will lead to a decrease in rate of reaction.
  d Go to Oxford Digital.
  e Halving the volume of a reaction system will increase the concentration of the reactants. The reduced space between the molecules will increase the frequency of collisions, leading to an increased proportion of successful collisions and rate of reaction.

23 **a** $C_3H_8(l) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
$\Delta H = -2{,}220\ \text{kJ mol}^{-1}$
  b 3,675 J
  c Go to Oxford Digital.
  d Go to Oxford Digital.
  e The complete combustion of propane produces CO_2 and H_2O, which are more stable and have lower enthalpies compared to CO, whereas incomplete combustion produces CO which is less stable than CO_2, leading to a less exothermic reaction (smaller ΔH).

24 **a** Temperature
  b Time taken for precipitate to form
  c As temperature increases, the time taken for the precipitate (sulfur) to form decreases in the reaction between thiosulfate and hydrochloric acid.
  d The reaction is both quantitative and qualitative. It gathers numerical data from the observations made for time taken, but this is also qualitative as subjective judgement is required to record this time.
  e As the temperature increases, the time taken to form enough precipitate to cover the black cross decreases. This could indicate an increase in the rate of reaction. This is because the reactant particles have more kinetic energy and collide with other reactant particles more frequently, increasing the proportion of successful collisions.
  f Student answers will vary.
  g Student answers may vary.

25 **a** Go to Oxford Digital.

b Ethanol has a lower molar enthalpy of combustion than octane due to its simpler molecular structure. Although it does contain a carbon–oxygen bond, it has far fewer carbon–carbon bonds than octane. This means that less energy is required to break the bonds in ethanol, and thus, the difference in enthalpy of reactants and products is smaller.

c Due to its larger size and more complex structure, octane would be predicted to have a higher activation energy for combustion compared to ethanol. This is because more energy is required to break the stronger and greater quantity of C–C and C–H bonds present in octane.

d Ethanol's smaller size and simpler structure results in fewer bonds to be broken and a lower activation energy for combustion. The lower activation energy for ethanol means that the reaction rate will be faster, similar to what we observe in the presence of catalysts.

Unit 2 Review

Multiple choice

1	A	**2**	C	**3**	A	**4**	C	**5**	A
6	D	**7**	B	**8**	C	**9**	C	**10**	C
11	D	**12**	C	**13**	B	**14**	C	**15**	C
16	D	**17**	C	**18**	C	**19**	B	**20**	A

Short response

21 Ammonia is pyramidal instead of a trigonal planar shape because three bonding electron pairs and one lone pair arrange themselves in a tetrahedral geometry to maximise repulsion, the lone pair repulsion pushes the hydrogen atoms closer together (from flat trigonal planar to pyramidal) as they have a greater repulsive force than the bonding pairs, and the lone pair repulsion also reduces the bond angles to approximately 107°, consistent with a pyramidal shape.

22 HF contains a highly polar H–F bond due to fluorine being the most electronegative element. This, combined with the linear shape of HF, allows it to form strong hydrogen bonds. However, these are not as extensive as in water, which can form up to four hydrogen bonds per molecule due to its bent shape. Therefore, water has a higher boiling point. The strong hydrogen bonds of HF give it a higher boiling point than NH_3, however, because nitrogen is less electronegative than fluorine, resulting in a less polar bond and therefore, weaker hydrogen bonds.

23 The increased number of carbon atoms in octane means there are more points of contact between molecules, and more opportunities for temporary dipoles to form, increasing the dispersion forces. Therefore, more energy is required to weaken the intermolecular forces between octane molecules, causing it to be a liquid at room temperature.

24 95.4 kPa

25 Increasing the temperature will increase the kinetic energy of the gas particles, causing them to move faster and collide more frequently with the container walls. The more frequent and forceful collisions result in an increase in pressure.

26 To increase the time that a diver can remain underwater, they must have a gas tank that can store a large volume of air. At constant temperature, volume is inversely proportional to pressure. Compressing air to fit into a tank that is a manageable size for the diver requires significant pressure, which increases the pressure exerted on the tank walls. The gas tank must have very strong walls to withstand this high internal pressure without bursting.

27 0.170 M

28 0.886 M

29 Student answers will vary.

30 Solution A: H_3PO_4
Solution B: CH_3NH_2
Solution C: NaOH
Solution D: HNO_3
Solution E: HF

31 **a** Both molecules are tetrahedral.

b Methane is non-polar and dichloromethane is polar.

c Methane: dispersion forces; dichloromethane: dispersion forces and dipole–dipole attractions

d Dichloromethane therefore requires more energy to overcome its stronger dipole–dipole attractions, giving it a higher boiling point than methane which only needs to overcome dispersion forces.

32 **a** Go to Oxford Digital.

b The yellow component in the sample has the same R_F value (0.33) as the yellow standard (0.33). Therefore, they have the same high affinity to the stationary phase as the standard. The red component in the sample, however, has a much greater R_F value (0.89) than the red standard (0.55). This indicates that the component had a greater affinity for the mobile phase than the standard did.

c The chromatogram can only accurately identify the yellow component of the sample, with the component having the same R_F value as the corresponding standard (0.33). The other two components, although appearing red and blue, do not have match the R_F values of the corresponding standards and therefore, remain unidentified.

33 **a** $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

b NaOH is in excess by 0.0025 mol.

c 12.6

34 **a** Nitrogen; the solubility decreases with temperature, which is characteristic of gases

b 450 g L^{-1}

c 25.0 g

35 **a** DV: volume of hydrogen (cm^3); IV: time (minutes)

b Student answers will vary.

c Curve 1: reaction C
Curve 2: reaction B
Curve 3: reaction A

d The higher concentration of HNO_3 in curves 1 and 2 provide more H^+ ions to completely react with the zinc, resulting in a greater volume of H_2 gas produced. In curve 3, the lower concentration of HNO_3 limits the amount of zinc that can react, resulting in a smaller volume of H_2 gas.

e Increasing the concentration of HNO_3 increases both the rate of reaction and the total volume of H_2 gas produced. However, at a concentration of 0.5 M, HNO_3 became a limiting reactant and did not react with all of the zinc available.

Glossary

%

Δ*H*
the difference in enthalpy between reactants and products in a chemical equation

A

absolute uncertainty
the exact magnitude of difference between the mean and the range of measurements; an indicator of the precision of measurements

absorbs
takes in; in this case, the electron "takes in" energy from a photon

absorption spectrum
the band of lines observed when an element absorbs light; viewed as a series of black lines on the visible light spectrum

accuracy
a measure of how close the measured value is to the true or accepted value

acid
a chemical compound that donates a hydrogen ion

acid–carbonate reaction
a chemical reaction between an acid and a carbonate to form a salt, carbon dioxide gas and water

activation energy (*Ea*)
the minimum amount of energy required in a collision for a reaction to occur

adsorption
the attraction of a substance within a sample to the stationary phase

affinity
the interaction of a substance within a sample with the mobile or stationary phases

alkali
a base that is soluble in water

alkane
a hydrocarbon that contains the maximum number of hydrogen atoms possible for the given number of carbon atoms; for linear or branched alkanes with n carbon atoms, the number of hydrogen atoms is 2n + 2; for a cycloalkane with one ring and n carbon atoms, the number of hydrogen atoms is 2n

alkene
a hydrocarbon that contains one or more double bonds between carbon atoms

alkyne
a hydrocarbon that contains one or more triple bonds between carbon atoms

allotrope
one of the different structural forms in which a particular element may exist in a given physical state; exist due to differences in bonding between atoms of that element

amount
the number of moles

analyte
the sample being analysed

anion
a non-metal atom which has gained one or more electrons and is negatively charged

aqueous solution
a mixture of a solute dissolved in water

aqueous
a substance that is dissolved in water or the state given to a substance that is dissolved in water

aromatic compound/arene
a cyclic unsaturated hydrocarbon compound that has delocalised clouds of electrons

atomic absorption spectroscopy (AAS)
the determination of the concentration of metal atoms in the gas phase by measuring the amount of light absorbed from a specific lamp

atomic radius
used to describe the size of an atom, which is measured by halving the distance between the nuclei of adjacent atoms

Avogadro's constant
the number of atoms or molecules in one mole of a substance; equivalent to 6.02 × 1023 particles per mol (mol^{-1})

B

balanced
having the same numbers of atoms of each element on either side of a reaction arrow

base
a chemical compound that accepts a hydrogen ion

bent
the molecular shape in which the central atom has two covalent bonds and two lone pairs of electrons, forming a V-shape

benzene
the simplest aromatic compound with the molecular formula C6H6, in which the carbon atoms are arranged in a flat (planar) six-membered ring, with one hydrogen atom attached to each carbon atom

best estimate
a value closest to the true value, usually found by taking repeated measurements and averaging

bibliography
a full list of all resources used in some research, provided at the end of a report or submitted along with another form of presentation

blank solution
a solution that does not contain the compound being measured, but contains all other compounds that would be present in the sample being tested; this allows the response due to those other compounds to be subtracted from other measurements

boiling point
the temperature at which a liquid becomes a gas

boiling
vaporisation at a substance's boiling point, which is the temperature at which a liquid forms bubbles of its own gas beneath the surface

bond angle
the angle between covalent bonds attached to the same central atom

bond energy
the amount of energy required to break a chemical bond in the gas phase, measure in $kJ\ mol^{-1}$

bond enthalpy
the amount of energy required to break 1 mol of a particular bond in the gaseous phase to give uncharged fragments

bonding electrons
valence electrons that are shared with another atom

bond stability
a measure of how difficult it is to break a bond; depends on the enthalpy of the bond

C

calibration curve
a graph of the peak area of a set of standards, obtained from a chromatogram, against their concentration

calibration
the configuration of an instrument against standard values to ensure that it measures a true value

catalyst
a substance that provides an alternative pathway for the reaction to occur that has a lower activation energy and increases the rate of reaction without participating in the chemical reaction

cation
a metal atom that has lost one or more electrons and is positively charged

causation
when a change in a single variable causes a change in a second variable

central tendency
the tendency for repeated measurements of the same value to be grouped around the mean, mode or median

centrifugation
the technique for separating mixtures where tubes are placed into a centrifuge that spins in a circular motion very fast to push solids to the bottom

chemical energy
a form of potential (stored) energy, stored within the chemical bonds of a substance

chemical reaction
when the bonds in a molecule break apart and atoms rearrange to form a new substance

chemistry
an experimental science concerned with the study of matter, and how substances can be combined or separated and how substances interact with energy

chromatograph
the pattern of bands, spots or peaks formed on the chromatography paper or TLC plate demonstrating the separation of a mixture

chromatography
an analytical technique that separates the components of a sample mixture based on the properties of the molecules

closed system
a system from which gas particles cannot escape

cognitive verb
a task word that requires you to perform a specific cognitive task

collision theory
the theory that states that reactants must collide with the correct orientation and sufficient energy for a reaction to occur

colloid
a heterogenous mixture in which the dispersed insoluble substance does not separate over time

combination reaction
a chemical reaction in which two or more elements and/or compounds react to form a single compound; also called a synthesis or addition reaction

combustion reaction
a chemical reaction in which a reactant is burnt with oxygen to form a metal oxide, covalent compound or carbon dioxide and water

complete combustion
when a hydrocarbon reacts with sufficient oxygen (an excess)

compound
a substance that is made up of atoms of different elements, joined together by chemical bonds

concentrated solution
a solution containing many particles per unit volume of solution

concentration
a measure of the number of particles in a specific volume of solution; units of concentration include mol L^{-1} (M), g L^{-1} and mg L^{-1} (ppm)

conclusion
a summary of the findings and results obtained from the research investigation

condensation
the process of turning a gas into a liquid by removing thermal energy; particles move slower and come close together

condense
to change from a gas to a liquid

condensed electron configuration
an electron configuration that only states electrons in the highest energy levels, based on the nearest noble gas

conduction
a process in which thermal energy is transferred between particles when they make direct contact with each other

confounding variable
an extraneous variable that changes systematically along with the variables being studied and offers an alternative explanation for the observations or measurements made

controlled variable
a condition that the experimenter tries to keep constant during an experiment, in order to minimise its effect on measurements or experiment outcomes

convection
a process in which thermal energy is transferred between particles via the movement of liquid or gas particles

coordinate covalent bond
a covalent bond in which two electrons come from only one atom

correlation
a link between a change in the independent variable and a change in the dependent variable; this does not mean that the changing independent variable caused the change in the dependent variable

covalent bonding
a sharing of electrons between atoms

covalent molecule
a chemical compound that is held together by covalent bonds

crystalline lattice
the ordered, repeating three-dimensional arrangement of atoms, molecules or ions in crystalline solids

D

decomposition
a chemical reaction in which a compound breaks down into simpler compounds or elements on the application of heat or electricity

delocalised electron
electron that is not retained in bonds between two atoms or associated with a single nucleus of an atom but is instead spread over several atoms or is highly mobile within a solid

density
how closely packed particles are; the mass of a substance per unit of volume

dependent variable (DV)
the variable that is measured or observed during an experiment, whose change is in response to changes made in the independent variable

deposition
the process of turning a gas into a solid by removing thermal energy; particles stop moving and become more attracted to one another

desorption
the release of a substance within a sample from the stationary phase into the mobile phase

diatomic molecule
a molecule containing two atoms of the same element or of different elements bonded together

dilute solution
a solution containing few particles per unit of volume

dilution
a method for lowering the concentration of a solution

dipole moment
when a bond gains a temporary negative charge at one end and a temporary positive charge at the other end; also known as a temporary dipole

dipole
the separation of opposite charges formed in a bond or molecule so that it has a partial positive charge at one end and a partial negative charge at the opposite end

dispersion force
the weak interaction between all molecules that occurs due to the formation of temporary dipoles or dipole moments

dissociate
break apart

dissolve
when a substance dissolves in a solvent

double displacement reaction
a chemical reaction in which both the cation and anion of one reactant swaps with the cation and anion of the second reactant, forming new products

ductile
able to be drawn into a wire

E

electromagnetic radiation
energy consisting of electromagnetic waves

electromagnetic spectrum
the range of frequencies of electromagnetic radiation and their energies and wavelengths; electromagnetic radiation, or light, is comprised of photons

electron configuration
a representation of the electrons in an atom

electron dot
represent the number of valence electrons in a Lewis structure

electron shielding effect
the shielding of outer electrons by inner electrons, causing a smaller force of attraction to the nucleus to be felt by the outer electrons

electron
a negatively charged subatomic particle that exists outside of the nucleus

electronegativity
the attraction of a positively charged nucleus to negatively charged electrons of a neighbouring atom

electrostatic force
an attraction between objects of opposing charges

element
a substance that is made up of one type of atom, defined by its atomic number

emission spectrum
the band of lines observed when an element emits light; viewed as a series of coloured lines that represent the wavelengths of electromagnetic radiation emitted by a source

emit
releases; in this case, the electron “releases” energy in the form of a photon

empirical formula
the simplest whole-number ratio of the elements in a compound

end point
the point in an acid–base titration at which the indicator changes colour

endothermic
describes a reaction in which the reactants have less energy than products; the required energy is absorbed from the environment as heat

energy profile diagram
a theoretical representation of the energy pathway that reactants must follow in order to form products

enthalpy
the total energy or heat content of chemical substances or a system

equivalence point
the point in a titration at which a reaction is complete; when the acid and base are mixed in exact stoichiometric ratios given by the balanced chemical equation

Eurocentrism
a worldview that focuses on or favours Western or European histories and thinking

evaporation
the phase change from a liquid to a gas at the surface of the liquid

excess reactant
the reactant that is not totally consumed by a chemical reaction

excited state
an energy state that is higher than the ground state; an electron is unstable in this state and emits energy as a photon of electromagnetic radiation to return to the ground state

exothermic
describes a reaction in which the reactants have more energy than the products; excess energy is released to the environment as heat

experimental yield
the actual amount of product that is produced in a chemical reaction

extraneous variable
a variable not being intentionally studied during an experiment that could have an unplanned or unintended effect on the results

extrapolation
the prediction of values beyond the range of data points by extending the trendline

F

First Nations peoples
past and current descendants of the original inhabitants and custodians of the Land we know today as Australia

flame test
the analysis of metal ions by heating them in a Bunsen burner flame and measuring the light emitted when excited electrons return to their ground state

flotation
the technique for separating different substances in a powdered ore depending on whether they sink in, or float on, a given liquid

formula unit
the formula of any compound that does not form discrete molecules; the smallest unit of a non-molecular substance, which applies to the formula of covalent network compounds and ionic compounds

free electron model of metallic bonding
a model of the bonding in metals that presents metals as ions surrounded by one or more valence electrons that are not associated with a particular atomic nucleus

freezing
the process of turning a liquid into a solid by removing thermal energy; particles stop moving around and stay in fixed positions where they can only vibrate

frequency of collisions
how often collisions occur in a reaction vessel in a period of time, i.e. the number of collisions in a unit of time

frequency
the number of waves (referring to wavelength) passing a point each second

full electron configuration
an electron configuration that states the location of every electron in the atom

G

gas laws
the laws that explain the properties of gases and how they behave at different temperatures, volumes and pressures

giant covalent network
an element or compound in which each atom is covalently bonded to other atoms to form an extended assembly of covalently bonded atoms

gradient
the slope of a graph

ground state
the lowest energy, and most stable, energy level that an electron can occupy

group
a vertical column of the periodic table; reflective of the number of outermost (valence) electrons

H

heat of dissolution
the enthalpy of a solution when a solid substance dissolved in water

heat of neutralisation
the enthalpy of a solution when an acid reacts with a base

heterogeneous catalyst
a catalyst that is in a different state from the reactants

heterogeneous mixture
a mixture with an inconsistent composition

homogeneous catalyst
a catalyst that is in the same state as the reactants

homogeneous mixture
a mixture with a uniform composition

hydrocarbon
a covalent molecule that only contains carbon and hydrogen atoms

hydrogen bonding
the strongest form of dipole–dipole attraction in which a partially negative atom (F, O or N) has an electrostatic attraction to a partially positive hydrogen atom that is bonded to an F, O or N in another molecule

hypothesis
a proposed explanation used as a starting point for further investigation

I

ideal gas
a gas in which particles do not interact with one another and occupy no volume

implication
potential consequence or effect of scientific results or conclusions

incomplete combustion
when a hydrocarbon reacts with a limited supply of oxygen

independent variable (IV)
the variable that is altered systematically in a controlled way by the experimenter to test the effect on a related dependent variable

indicator
a substance that changes colour in the presence of high and low hydrogen ion concentrations

insoluble
not able to dissolve in a solvent

intermolecular force
a force between molecules; inter = “between”

interpolation
the prediction of values between data points using a trendline

in-text reference
an acknowledgement of the source immediately after the research or information is referred to

intramolecular force
an attractive force between molecules; *inter* = “between”

ion
an atom containing an electric charge

ion–dipole interaction
an ion with a whole charge interacting with a partial charge caused by a polar bond or molecule

ionic bonding
the process by which atoms transfer valence electrons to each other

ionic charge
the difference between the positive charges from the protons and the negative charges from the electrons, represented as a superscript on the right-hand side of an elemental symbol

ionic compound
a chemical compound that is held together by ionic bonds

ionic equation
a chemical equation that shows only the ions that react during a chemical reaction, along with any neutral chemical species

ionic radius
the radius of an ion

ionisation energy
the energy (in $kJ\,mol^{-1}$) required to remove an electron from a gaseous atom

ionisation
the formation of ions from atoms by the addition or removal of electrons

ionise
form ions

isotope
atoms that have the same number of protons and electrons but different numbers of neutrons

K

kinetic energy
the energy involved with the motion of particles; can be increased by adding heat energy

kinetic molecular theory
the theory that states that all particles are in constant random motion

L

law of conservation of energy
states energy cannot be created or destroyed

law of conservation of mass
states that in an isolated system, mass cannot be created or destroyed

Lewis structure
visual diagrams to show the number of electrons in the valence shell

limiting reactant
the reactant that is totally consumed when a chemical reaction is complete; this limits the amount of product that can be produced

linear
the molecular shape in which the central atom has two covalent bonds and no lone pairs of electrons, forming a straight line

linearising
the process of transforming non-linear data by applying a mathematical function to one of the variables so that the relationship between the variables becomes closer to a straight line

lit splint test
a test to determine the presence of carbon dioxide gas

lone pair electrons
valence electrons that are not bonded to another atom; also known as non-bonding electrons

M

***m/z* ratio**
mass-to-charge ratio; the molar mass of the positive ion formed divided by its charge

malleable
able to be pressed into other shapes

mass spectrometer
an instrument that ionises elements and molecules and produces spectra showing mass-to-charge ratio

mass spectrometry
an analytical technique in which chemical elements are ionised and the ions are sorted according to their mass-to-charge ratio

mass spectrum
the column graph obtained from a mass spectrometer; displays m/z ratio on the x-axis and percentage abundance of each positively charged ion on the y-axis

matter
physical substances that have mass

maximum trendline
a line of best fit of maximum gradient within the bounds of the error bars

Maxwell–Boltzmann distribution
a representation of the various kinetic energies of each individual particle in a system

mean
the average of multiple values

median
the value that appears in the middle when the dataset is sorted from smallest to largest value

melting point
the temperature at which a solid becomes a liquid

melting
the process of turning a solid into a liquid by applying thermal energy; particles move away from fixed positions, but are still close together

metallic bond
the attractive force between a metal ion and delocalised electrons that hold metals together

metallic character
how readily an element loses valence electrons

metalloid
an element that displays both metallic and non-metallic properties

minimum trendline
a line of best fit of minimum gradient within the bounds of the error bars

mixture
a substance that is made up of two or more substances that are not chemically bonded

mobile phase
the phase that flows in chromatography, moving the components of a sample at different rates over the stationary phase

mode
the value that appears the most often in a dataset

molar mass
the mass of one mole of a substance

molar volume
the volume occupied by 1 mole of a gas at a specified temperature and pressure

molarity
molar concentration, units are moles per litre, $mol\,L^{-1}$ or M

mole ratio
the ratio of the amounts in moles of the reactants and products in a reaction

mole
the quantity of a substance containing the same number of particles as there are atoms in exactly 12.000 g of carbon-12, which is the same as Avogadro's number (6.02×10^{23})

molecular formula
the chemical formula of a compound that indicates the number of atoms of each element present

molecular shape
the shape formed between a central atom and bonded atoms when covalent bonds form

molecule
a chemical species with no overall charge that is held together by covalent bonds

molten
a melted substance

monatomic ion
an ion with only one atom; also called a monoatomic ion

N

negative control
a test to make sure that an experiment is working as expected, which should not cause any change to the DV

neutral
an aqueous solution that has a pH of 7

neutralisation reaction
the reaction between an acid and a base to produce water and a metal salt

neutron
a neutral subatomic particle that exists inside the nucleus

non-metallic character
how readily an element gains valence electrons

nuclear symbol
the $^{A}_{Z}\mathrm{X}$ notation used to represent atoms, where A is the mass number, Z is the atomic number and X is the element symbol

nucleus
the dense structure within the atom

O

octet rule
atoms gain and lose electrons to reach eight electrons in their valence shell; exceptions to this rule are hydrogen and helium

orbital
the three-dimensional region of space around an atomic nucleus where an electron is most likely to be found

origin
the line applied to a chromatogram to mark the point where the sample or standard was placed

outlier
a value that is much smaller or larger than most of the other values in a set of data; it is greater than three standard deviations away from the mean

oxidation state
the charge assigned to substances in a chemical reaction

P

paper chromatography
an analytical technique for separating and identifying mixtures; the stationary phase is a thin strip of absorbent paper

partially soluble
only partially dissolves in a solvent, forming a solution of low concentration

percentage abundance

the percentage of each isotope contained within a sample of element; must always add up to 100%

percentage composition

the percentage that each element contributes to the mass of a compound

percentage error

the percentage difference between the accepted (true or theoretical) value and the measured (experimental) value

percentage uncertainty

an indicator of uncertainty in which the range of values for a measurement result (the absolute uncertainty) is expressed as a percentage of the measurement

period

a horizontal row of the periodic table; reflective of the number of energy levels (shells) within an atom

periodic table

an organised presentation of elements by increasing atomic number

pH scale

a numerical scale that is used to determine the level of acid or base in an aqueous solution

pH

a qualitative measure of hydrogen ion concentration

photon

a particle and wave of light that carries energy; is absorbed or emitted from atoms

physical change

when the molecules remain the same but the substance is in a different state that has different properties

picometre

10^{-12} or 0.000000000001 metres

polar

having a dipole; for example, a bond or molecule

polarity

gives the molecule partially positive and partially negative ends; determined from the electronegativities of the atoms and the shape of a molecule

polyatomic ion

an ion with two or more atoms

pop test

a test to determine the presence of hydrogen gas

positive control

a test to make sure that an experiment is working as expected, which should give a positive result

precipitate

the solid product formed as a result of a precipitation reaction

precipitation reaction

a reaction between two soluble ionic substances that forms a solid product

precision

the consistency and reproducibility of a series of measurements

pressure

the force exerted, per unit area, by one substance upon another substance

primary data

first-hand data collected from an experiment designed to answer the specific research question

proportion of successful collisions

the percentage of all collisions which are successful, resulting in the breaking of reactant bonds to form products

proton

a positively charged subatomic particle that exists inside the nucleus

pure substance

a substance that is made up of a single type of element or compound

pyramidal

the molecular shape in which the central atom has three covalent bonds and one lone pair of electrons, forming a pyramid structure

Q

qualitative data

data about types or categories, which may be represented by names, symbols, codes, etc.

quantitative data

numerical data about a substance, object or phenomenon

R

random error

an unpredictable error in measurement or experimental procedure that has no detectable pattern

raw data

primary data that is collected but not yet processed or analysed

reaction rate

the change in concentration of a reactant or product per unit of time

reactivity

a measure of how readily a substance reacts with other substances

redox reaction

a chemical reaction involving the transfer of electrons from one reactant to another

refute

contradicts or does not support (in relation to the claim or hypothesis)

relative atomic mass (RAM)

the mass of an atom once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

relative atomic mass

the mass of an atom once the mass and percentage abundance of all isotopes have been considered, relative to the carbon-12 isotope

relative isotopic mass (RIM)

the mass of isotopes measured on a mass spectrum, relative to the carbon-12 isotope

reliability

the ability to be trusted to be accurate or correct, or to provide a correct result

replicate

a repeated measurement which gives an indication of the precision of measurements

retardation factor (R_f)

the ratio of the distance travelled by a component of a sample, from the origin, to the distance travelled by the mobile phase

risk assessment

an evaluation in a systematic way of the potential risks and their likelihood involved in doing an experiment or activity

S

safety data sheet

a document that lists information relates to hazardous and non-hazardous materials; also known as a product safety data sheet (PSDS)

saturated hydrocarbon

a compound of carbon and hydrogen that contains the maximum possible number of hydrogen atoms for the number of carbon atoms present

saturated

a solution in which no more solute can dissolve

saturation point

the point at which the maximum amount of solute has dissolved in the known quantity of solute at a given temperature

scientific method

systematic exploration of a phenomenon or topic using observation, measurement and experiment to support, disprove or modify hypotheses

secondary data

second-hand data obtained from other sources

significant figures

the number of numerical figures that can be used to express a measured or calculated quantity

single displacement reaction

a chemical reaction in which a more reactive element takes the place of a less reactive element in a compound

solubility

the maximum amount of solute that dissolves in a known quantity of solvent at a given temperature

solute

the minor components of a solution; a substance dissolved in a solvent to form a solution

solution calorimeter

a device that measures the change in temperature of water when an endothermic or exothermic reaction takes place

solution calorimetry

an analytical technique in which the energy generated by a sample is determined by measuring the change in the temperature of water

solution

a mixture of a solute dissolved in a solvent

solvent front

the point on a chromatogram where the mobile phase reaches before the analysis is terminated

solvent

the major component of a solution; a substance that can dissolve other substances

specific heat capacity
the amount of heat energy, in joules (J), that it takes to increase the temperature of a specific volume of water by 1 K or 1°C (remembering that 1 g of water is 1 mL)

spectator ion
an ion that does not participate in a chemical reaction

standard deviation
a statistical value which expresses how spread out a group of values are or by how much they differ from the mean value for the group

standard laboratory conditions (SLC)
standard conditions for temperature and pressure under laboratory conditions, when gases are measured at 100 kPa (1 bar) and 25°C (298.15 K)

standard operating procedure (SOP)
step-by-step guidance for how to operate a piece of equipment or instrument correctly and safely

standard temperature and pressure (STP)
standard conditions for temperature and pressure, when gases are measured at 100 kPa (1 bar) and 0°C (273 K)

standard
a sample with a known concentration of the chemical that is being tested for

state
the state of matter assigned to each chemical reactant or product – solid (s), liquid (l) gas (g), or aqueous (aq)

stationary phase
the phase to which the components of a chromatographic sample are adsorbed

stoichiometric ratio
the whole number ratio between each reactant and product in a reaction

stoichiometry
the numerical relationship between the amount of reactants used and the products produced in a reaction; the application of the numerical relationship in calculations

strong acid
an acid that completely ionises in water

strong base
a base that completely dissociates in water

subatomic particle
a component that makes up, and exists within, the atom

sublimation
the process of turning a solid into a gas by applying thermal energy; particles move and separate from one another

supersaturated
a solution that has been heated to dissolve more solute than can be dissolved under equilibrium conductions; on cooling, the solute remains in solution in a metastable state, until a disturbance (seed crystal, speck of dust, tapping the container) starts the crystallisation process

support
agrees with (in relation to the claim or hypothesis)

surface area
a measure of the total surface area available to react in a chemical reaction

surface energy
excess energy at the surface of a material

surface tension
the ability of the surface of a liquid to resist an external force

systematic error
a consistent, repeatable error that occurs every time a piece of equipment is used

T

temperature
a measure of the average heat energy of the particles within a system

tetrahedral
the molecular shape in which the central atom has four covalent bonds and no lone pairs of electrons, forming a tetrahedral structure

theoretical yield
the amount of product predicted from stoichiometric calculations

thermal energy
the energy any substance has at a particular temperature

thermal radiation
a process in which thermal energy is transferred in the form of electromagnetic waves

thermochemical equation
a balanced chemical equation, including states, with a ΔH value that is positive or negative

thin layer chromatography (TLC)
an analytical technique for separating and identifying mixtures; the stationary phase is typically a thin layer of silica gel, aluminium oxide or cellulose supported on a piece of glass or plastic

titrant
the solution being delivered in a titration

titration
a quantitative analytical technique used to find the unknown concentration of a solution

transition state
an intermediate state in which unstable, loosely arranged atoms reach the highest energy level (activation energy) of a reaction pathway, where bonds are breaking and re-forming; also called the activated complex

trendline (line of best fit)
a line drawn on a graph joining as many points as possible and showing the general direction of the data; should be drawn with, approximately, an equal number of points above and below the line

trigonal planar
the molecular shape in which the central atom has three covalent bonds and no lone pairs of electrons; sometimes called triangular planar

U

universal gas constant (R)
the proportionality constant that defines gas behaviour under ideal conditions

unsaturated hydrocarbon
a compound of carbon and hydrogen that contain double or triple bonds and therefore have fewer than the maximum number of hydrogen atoms for a particular number of carbon atoms

unsaturated
a solution that can dissolve more solute

V

valence electron
an electron in the outermost energy level of an atom

valence shell electron pair repulsion (VSEPR) theory
states that electron pairs in bonds and lone electron pairs around a central atom repel each other, forming a 3D shape

valency
the combining power of an element as determined by the number of electrons in its outer energy level (valence electrons)

validity
a measure of whether the investigation is sound and measures what it is intended to measure

vaporisation
the process of turning a liquid into a gas by applying thermal energy; particles move faster and separate from one another

vaporise
to change from a liquid to a gas

vapour pressure
the pressure of a vapour in contact with its solid or liquid form in an enclosed container

variable
a condition or parameter that is changed or changes during an experiment

volatile substance
a substance that easily evaporates from a liquid to a gas

volume
a measure of the space occupied by a substance

W

wavelength
the distance between two crests or troughs of a wave

weak acid
an acid that only partially dissociates in water

weak base
a base that only partially dissociates in water

Y

yield
the amount of product formed from a chemical reaction

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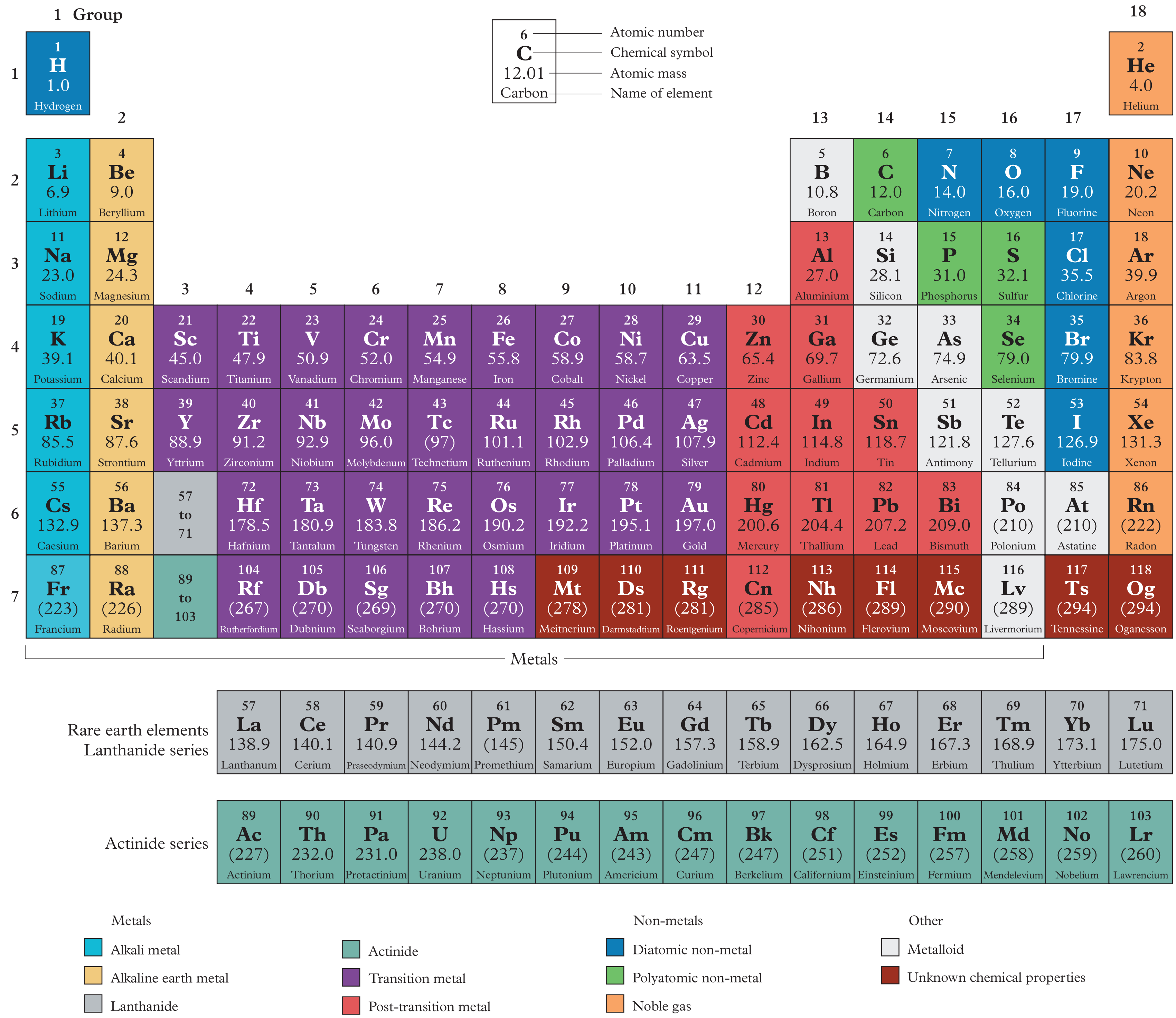
Appendix: Periodic table



6 — Atomic number
C — Chemical symbol
12.01 — Atomic mass
Carbon — Name of element

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 1.0 Hydrogen																	2 **He** 4.0 Helium
2	3 **Li** 6.9 Lithium	4 **Be** 9.0 Beryllium											5 **B** 10.8 Boron	6 **C** 12.0 Carbon	7 **N** 14.0 Nitrogen	8 **O** 16.0 Oxygen	9 **F** 19.0 Fluorine	10 **Ne** 20.2 Neon
3	11 **Na** 23.0 Sodium	12 **Mg** 24.3 Magnesium											13 **Al** 27.0 Aluminium	14 **Si** 28.1 Silicon	15 **P** 31.0 Phosphorus	16 **S** 32.1 Sulfur	17 **Cl** 35.5 Chlorine	18 **Ar** 39.9 Argon
4	19 **K** 39.1 Potassium	20 **Ca** 40.1 Calcium	21 **Sc** 45.0 Scandium	22 **Ti** 47.9 Titanium	23 **V** 50.9 Vanadium	24 **Cr** 52.0 Chromium	25 **Mn** 54.9 Manganese	26 **Fe** 55.8 Iron	27 **Co** 58.9 Cobalt	28 **Ni** 58.7 Nickel	29 **Cu** 63.5 Copper	30 **Zn** 65.4 Zinc	31 **Ga** 69.7 Gallium	32 **Ge** 72.6 Germanium	33 **As** 74.9 Arsenic	34 **Se** 79.0 Selenium	35 **Br** 79.9 Bromine	36 **Kr** 83.8 Krypton
5	37 **Rb** 85.5 Rubidium	38 **Sr** 87.6 Strontium	39 **Y** 88.9 Yttrium	40 **Zr** 91.2 Zirconium	41 **Nb** 92.9 Niobium	42 **Mo** 96.0 Molybdenum	43 **Tc** (97) Technetium	44 **Ru** 101.1 Ruthenium	45 **Rh** 102.9 Rhodium	46 **Pd** 106.4 Palladium	47 **Ag** 107.9 Silver	48 **Cd** 112.4 Cadmium	49 **In** 114.8 Indium	50 **Sn** 118.7 Tin	51 **Sb** 121.8 Antimony	52 **Te** 127.6 Tellurium	53 **I** 126.9 Iodine	54 **Xe** 131.3 Xenon
6	55 **Cs** 132.9 Caesium	56 **Ba** 137.3 Barium	57 to 71	72 **Hf** 178.5 Hafnium	73 **Ta** 180.9 Tantalum	74 **W** 183.8 Tungsten	75 **Re** 186.2 Rhenium	76 **Os** 190.2 Osmium	77 **Ir** 192.2 Iridium	78 **Pt** 195.1 Platinum	79 **Au** 197.0 Gold	80 **Hg** 200.6 Mercury	81 **Tl** 204.4 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 209.0 Bismuth	84 **Po** (210) Polonium	85 **At** (210) Astatine	86 **Rn** (222) Radon
7	87 **Fr** (223) Francium	88 **Ra** (226) Radium	89 to 103	104 **Rf** (267) Rutherfordium	105 **Db** (270) Dubnium	106 **Sg** (269) Seaborgium	107 **Bh** (270) Bohrium	108 **Hs** (270) Hassium	109 **Mt** (278) Meitnerium	110 **Ds** (281) Darmstadtium	111 **Rg** (281) Roentgenium	112 **Cn** (285) Copernicium	113 **Nh** (286) Nihonium	114 **Fl** (289) Flerovium	115 **Mc** (290) Moscovium	116 **Lv** (289) Livermorium	117 **Ts** (294) Tennessine	118 **Og** (294) Oganesson

Metals

Rare earth elements Lanthanide series	57 **La** 138.9 Lanthanum	58 **Ce** 140.1 Cerium	59 **Pr** 140.9 Praseodymium	60 **Nd** 144.2 Neodymium	61 **Pm** (145) Promethium	62 **Sm** 150.4 Samarium	63 **Eu** 152.0 Europium	64 **Gd** 157.3 Gadolinium	65 **Tb** 158.9 Terbium	66 **Dy** 162.5 Dysprosium	67 **Ho** 164.9 Holmium	68 **Er** 167.3 Erbium	69 **Tm** 168.9 Thulium	70 **Yb** 173.1 Ytterbium	71 **Lu** 175.0 Lutetium
Actinide series	89 **Ac** (227) Actinium	90 **Th** 232.0 Thorium	91 **Pa** 231.0 Protactinium	92 **U** 238.0 Uranium	93 **Np** (237) Neptunium	94 **Pu** (244) Plutonium	95 **Am** (243) Americium	96 **Cm** (247) Curium	97 **Bk** (247) Berkelium	98 **Cf** (251) Californium	99 **Es** (252) Einsteinium	100 **Fm** (257) Fermium	101 **Md** (258) Mendelevium	102 **No** (259) Nobelium	103 **Lr** (260) Lawrencium

Metals
- ■ Alkali metal
- ■ Alkaline earth metal
- ■ Lanthanide
- ■ Actinide
- ■ Transition metal
- ■ Post-transition metal

Non-metals
- ■ Diatomic non-metal
- ■ Polyatomic non-metal
- ■ Noble gas

Other
- □ Metalloid
- ■ Unknown chemical properties

Notes

Notes

Notes

OTHER TITLES
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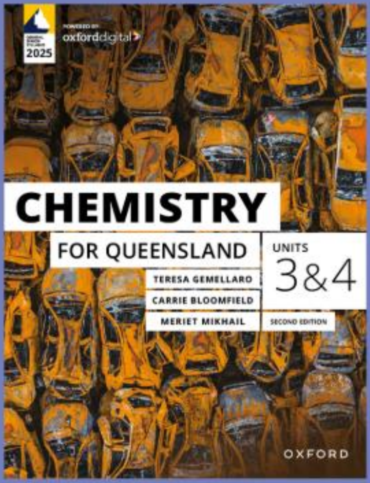

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