

NELSON

QCE Chemistry

UNITS

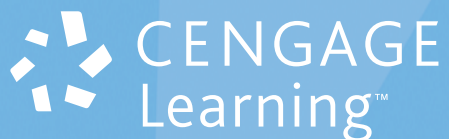
1

2



Debra Smith
Michelle Brown
Rachel Bond





COPYRIGHT NOTICE

Copyright in this work is owned by Cengage Learning Australia ("the work"). A condition of purchase of this electronic version of the work is that you agree to respect the copyright in the work, abide by the Copyright Act 1968 and specifically agree not to transfer, sell, assign, misuse, copy or transmit an electronic or other version of the work to any third party.

Please note: This product is accompanied by a licence (single user, network or adoption) governing the terms and conditions of its use.

This is a legal agreement between the you, (the “Customer”) and Cengage Learning Australia Pty Limited (ABN 14 058 280 149) (the “Licensor”) which provides the terms and conditions of this non-exclusive licence and the limited warranty for the Product. Use of the Product indicates an acknowledgement that the Customer has read and agreed to be bound by the terms and conditions of this Agreement. If you do not agree to these terms and conditions, return the Product to the place of purchase within 15 days of the date of purchase (with proof of purchase) for a full refund

1. Licence Grant

You do not receive title to the Product. Copyright in the Product (which includes all images, photographs, video, animations, audio, music and text incorporated in the Product, including all of the accompanying printed material) is owned by the Licensor and/or its suppliers and is protected by Australian copyright laws. The Licensor grants you a non-exclusive licence to use the Product subject to the restrictions and terms set out in this Agreement.

2. A Licence allows you to:

Use the Product on your computer. The Customer represents that they shall in no way place the Product in the public domain or in any way compromise our copyright in the Material. You agree to take reasonable steps to protect our copyright.

3. You may not:

Alter, modify, translate, reverse engineer, decompile, or adapt the software or create derivative works based on the Product. Make further copies by any means technological, electronic, digital whatsoever without the written permission of the Licensor. Rent or transfer all or any part of your rights under this Agreement. Remove or alter any copyright or other proprietary notice or label attached to the software.

4. Termination

Any failure to comply with the terms and conditions of this agreement will result in the automatic termination of this licence. Upon termination of this licence for any reason, the Customer must destroy or return to the Licensor all copies of the software and accompanying documentation.

5. Warranties

To the extent permitted by law, the Licensor’s liability for any breach of the warranty or any term implied by law into this licence is limited to the lowest cost of replacing the goods, acquiring equivalent goods or having the goods repaired.

NELSON

QCE Chemistry

UNITS

1

2

LEARNING DISCOVERY

EXOTHERMIC REACTIONS AND ENTHALPY



In a world of atoms and molecules, some reactions release energy, warming our world, while others absorb energy, resulting in pockets of cool. These reactions help to drive life itself and power inventions that propel us forward.

Debra Smith
Michelle Brown
Rachel Bond

Nelson QCE Chemistry Units 1 & 2

1st Edition

Debra Smith

Michelle Brown

Rachel Bond

Product manager: Katherine Roan

Content developer: Dr Trisha Downing

Editor: Catherine Greenwood

Proofreader: Helen Koehne

Series designer: Danielle Maccarone

Series cover designer: Regine Abos (Studio Regina)

Series text design: Rina Gargano (Alba Design)

Cover image: stock.adobe.com/oksamedvedeva

Permissions researcher: Lumina Datamatics

Content manager: Alice Kane

Typeset by MPS Limited

Any URLs contained in this publication were checked for currency during the production process. Note, however, that the publisher cannot vouch for the ongoing currency of URLs.

Acknowledgements

Selected VCE Examination questions from the VCE Study Design are copyright Victorian Curriculum and Assessment Authority (VCAA), reproduced by permission. VCE ® is a registered trademark of the VCAA. The VCAA does not endorse this product and makes no warranties regarding the correctness or accuracy of this study resource. To the extent permitted by law, the VCAA excludes all liability for any loss or damage suffered or incurred as a result of accessing, using or relying on the content. Current VCE Study Designs, past VCE exams and related content can be accessed directly at www.vcaa.edu.au

© 2024 Cengage Learning Australia Pty Limited

Copyright Notice

This Work is copyright. No part of this Work may be reproduced, stored in a retrieval system, or transmitted in any form or by any means without prior written permission of the Publisher. Except as permitted under the *Copyright Act 1968*, for example any fair dealing for the purposes of private study, research, criticism or review, subject to certain limitations. These limitations include: Restricting the copying to a maximum of one chapter or 10% of this book, whichever is greater; providing an appropriate notice and warning with the copies of the Work disseminated; taking all reasonable steps to limit access to these copies to people authorised to receive these copies; ensuring you hold the appropriate Licences issued by the Copyright Agency Limited ("CAL"), supply a remuneration notice to CAL and pay any required fees. For details of CAL licences and remuneration notices please contact CAL at Level 11, 66 Goulburn Street, Sydney NSW 2000, Tel: (02) 9394 7600, Fax: (02) 9394 7601
Email: info@copyright.com.au
Website: www.copyright.com.au

For product information and technology assistance,
in Australia call **1300 790 853**;
in New Zealand call **0800 449 725**

For permission to use material from this text or product, please email
aust.permissions@cengage.com

ISBN 978 0 17 048346 9

Cengage Learning Australia

Level 5, 80 Dorcas Street
Southbank VIC 3006 Australia

For learning solutions, visit cengage.com.au

Printed in China by 1010 Printing International Limited.

1 2 3 4 5 6 7 28 27 26 25 24



CONTENTS

Authors and reviewer team	vi
About this book	vii

DIGITAL CHAPTER

DC	SCIENCE RESEARCH	
	DC.1 Student experiment	DC-6
	DC.2 Research investigation	DC-28

Chapter summary	DC-34
Chapter exam	DC-39
Answers	DC-43

UNIT ONE CHEMICAL FUNDAMENTALS: STRUCTURE, PROPERTIES AND REACTIONS

2

TOPIC 1: PROPERTIES AND STRUCTURE OF ATOMS

1	ATOMIC STRUCTURE	4
	1.1 The atom	6
	1.2 Atoms, elements and the periodic table	9
	1.3 The electron configuration of atoms	13

Chapter summary	23
Chapter exam	25

2	ISOTOPES		28
	2.1	Different forms of the same element	30
	2.2	Relative atomic mass	32
	Chapter summary		35
	Chapter exam		36

3	THE PERIODIC TABLE AND TRENDS		40
	3.1	The periodic table of elements	42
	3.2	Trends in the periodic table	45
	3.3	Metallic and non-metallic character in the periodic table	54
	3.4	Properties of the oxides of the elements in period 3	57
	Chapter summary		59
Chapter exam		61	

4	ANALYTICAL TECHNIQUES		66
	4.1	Mass spectrometry	68
	4.2	Emission spectroscopy	71
	4.3	Absorption spectroscopy	77
	Chapter summary		83
	Chapter exam		85

5	INTRODUCTION TO BONDING		90
	5.1	Chemical bonding	92
	5.2	Ionic bonding	94
	5.3	Metallic bonding	101
	5.4	Covalent bonding	102
	Chapter summary		108
	Chapter exam		110

SCIENCE AS A HUMAN ENDEAVOUR: EXPLORING RADIOISOTOPES AND RADIOACTIVITY	114

TOPIC 2: PROPERTIES AND STRUCTURE OF MATERIALS

6	COMPOUNDS AND MIXTURES		116
	6.1	Classifying substances	118
	6.2	Heterogeneous and homogeneous mixtures	120
	6.3	Properties of pure substances and mixtures	126
	Chapter summary		128
	Chapter exam		130

7	BONDING AND PROPERTIES		134
	7.1	Ionic compounds	136
	7.2	Metals	139
	7.3	Covalent compounds	143
	7.4	Hydrocarbons	148
	7.5	Comparing different substances	157
	Chapter summary		161
	Chapter exam		163

**SCIENCE AS A HUMAN ENDEAVOUR:
NANOTECHNOLOGY IN COSMETICS 166**

**TOPIC 3: CHEMICAL REACTIONS –
REACTANTS, PRODUCTS AND
ENERGY CHANGES**

8 CHEMICAL REACTIONS 168

- 8.1 Physical and chemical change 170
- 8.2 Balanced chemical equations 173
- 8.3 Types of chemical reactions 177
- Chapter summary 184
- Chapter exam 186

**9 THE MOLE CONCEPT AND THE
LAW OF CONSERVATION OF MASS 190**

- 9.1 The mole and Avogadro's number
of particles 192
- 9.2 The mole concept 196
- 9.3 The law of conservation of mass
in chemical equations 207

- 9.4 Stoichiometry 209
- Chapter summary 221
- Chapter exam 223

**10 EXOTHERMIC AND ENDOTHERMIC
REACTIONS 226**

- 10.1 Energy and temperature 228
- 10.2 The law of conservation
of energy 230
- 10.3 Relationship between
temperature and enthalpy 238
- 10.4 Calculating energy change 240
- 10.5 Hess's law and enthalpy
change calculations 247
- Chapter summary 252
- Chapter exam 254

**SCIENCE AS A HUMAN ENDEAVOUR:
FOSSIL FUELS VERSUS BIOFUELS 258**

UNIT TWO MOLECULAR INTERACTIONS AND REACTIONS 260

**TOPIC 1: INTERMOLECULAR
FORCES AND GASES**

11 INTERMOLECULAR FORCES 262

- 11.1 Molecular structure: valence
shell electron pair repulsion
(VSEPR) theory 264
- 11.2 The polarity of molecules 268
- 11.3 Observable properties and
intermolecular forces 274
- Chapter summary 282
- Chapter exam 284

12 CHROMATOGRAPHY 288

- 12.1 Chromatography techniques 290
- 12.2 Separating components 297
- Chapter summary 303
- Chapter exam 305

13 GASES 312

- 13.1 Gases 314
- 13.2 Kinetic theory 317
- Chapter summary 332
- Chapter exam 334

**SCIENCE AS A HUMAN ENDEAVOUR:
MAKING SEA EXPLORATION POSSIBLE 337**

**TOPIC 2: AQUEOUS SOLUTIONS
AND ACIDITY**

14 AQUEOUS SOLUTIONS AND MOLARITY 340

- 14.1 The structure and properties of water 342
- 14.2 Solutes, solvents and solutions 349
- 14.3 Concentration 350
- 14.4 Dilutions and reacting quantities 354
- 14.5 Saturated solutions 361
- Chapter summary 363
- Chapter exam 365

15	IDENTIFYING IONS IN SOLUTION	368
	15.1 Solubility rules and writing precipitation equations	370
	15.2 Determination of ions in solution	378
	Chapter summary	383
	Chapter exam	384

16	SOLUBILITY	388
	16.1 Solubility and intermolecular bonding	390
	16.2 The effect of temperature on solubility	398
	16.3 Interpreting and analysing solubility curves	402
	Chapter summary	408
	Chapter exam	409

17	REACTIONS OF ACIDS AND BASES	414
	17.1 Reactions of acids	416
	17.2 Other acid reactions	420
	Chapter summary	426
	Chapter exam	427

18	pH	430
	18.1 pH and the pH scale	432
	18.2 The Arrhenius model	436
	18.3 Recognising strong and weak acids and bases	438
	Chapter summary	443
	Chapter exam	444

SCIENCE AS A HUMAN ENDEAVOUR:	
BLOOD PLASMA – ‘LIQUID GOLD’	447

TOPIC 3: RATES OF CHEMICAL REACTIONS

19	RATES OF REACTIONS	450
	19.1 The rate of a chemical reaction	452
	19.2 What is needed for a chemical reaction to occur?	457
	19.3 Factors that affect the rate of a chemical reaction	462
	Chapter summary	475
	Chapter exam	477

SCIENCE AS A HUMAN ENDEAVOUR:	
DRIVERS OF CHEMICAL REACTIONS	483

Answers	486
Glossary	522
Index	530

AUTHOR AND REVIEWER TEAM

Nelson QCE Chemistry is adapted from *Nelson QScience Chemistry Units 1 & 2* and *Nelson QScience Chemistry Units 3 & 4*, by Nicholas Stansbie, Brett Steeples and Dr Sarah Windsor.

DEBRA SMITH

Debra Smith is an experienced Chemistry teacher and author. Debra has been Head of Science at a number of Queensland state high schools and has been involved in writing science curriculum at the local, state and national levels. She has been President and Treasurer of the Australian Science Teachers Association and is a life member of STAQ. Debra has been the recipient of state and national teaching awards, including the 2010 Prime Minister's Prize for Excellence in Science Teaching – Secondary.

MICHELLE BROWN

Michelle Brown studied Applied Science at Queensland University of Technology before working at a private laboratory in Brisbane as a research assistant for two years mapping the chromosomes of trees. After completing a Graduate Diploma in Education, she taught Chemistry, Biology, Science, Mathematics, and Computer studies. She has been teaching Chemistry for more than 25 years in both state and independent schools.

RACHEL BOND

Rachel Bond is an experienced Chemistry teacher in Queensland, having worked both in the Queensland state and Catholic educational systems. She is currently Curriculum Leader of Science at the Secondary Campus of a Queensland Catholic college. She has led Chemistry communities of practice in her local area, has presented at the Australian Science Teachers Association Conference, and is editor for a company producing external Chemistry exam papers for Queensland audiences. She has also been awarded an academic scholarship for her Masters Studies of Educational Leadership at the Australian Catholic University.

REVIEWERS

The following people have contributed to the review of the QCE Chemistry series (in no particular order): Florence Coghlan, Erica McLean, Colette Anderson and Wendy Cook.

ACKNOWLEDGEMENTS

The publisher would like to acknowledge the authors for their expertise and devotion to supporting teachers and students engaging in Chemistry around Queensland.

ABOUT THIS BOOK

Nelson QCE Chemistry Units 1 & 2 is a comprehensive textbook specifically tailored to align with the 2025 QCAA Senior Secondary Science Syllabus – Chemistry. It has been thoughtfully developed to empower students by providing a strong foundation in essential concepts and equipping them with the necessary skills to excel in their studies. Emphasising the importance of making connections between topics and practising exam techniques, this edition is designed to support students in unlocking their full potential and achieving success in their chemistry journey.

AT THE BEGINNING OF THE BOOK

- An online chapter to support the development of students' science inquiry skills for application throughout the syllabus.

AT THE BEGINNING OF EACH UNIT

- Unit introductions are an overview of the key content in the unit.



AT THE BEGINNING OF EACH CHAPTER

- Chapter introduction to set the context of the upcoming key content
- List of syllabus dot points being covered in the chapter
- List of student resources available on Nelson MindTap



IN EACH CHAPTER

- **Assumed knowledge** – knowledge and skills students are expected to know coming into the chapter that relate to the chapter content
- **Learning outcomes** – highlights the key outcomes from chapter
- **Key terms** – defined in situ to help students deconstruct scientific language
- **Learning check** – written to the developmental levels highlighted in the syllabus objectives
- **Key formulas** – important formulas to remember
- **Syllabus links** – highlighting links to other areas in the syllabus from different chapters to help students make connections
- **Practicals** – syllabus-aligned practicals with guided instructions on the materials, procedure, collection and analysis of results, and discussion

The collage displays five pages from the IB Chemistry textbook, illustrating the components of each chapter:

- Page 1 (Left):** Shows 'Assumed knowledge' (e.g., understanding of atoms, moles, and chemical equations) and 'Learning outcomes' (e.g., understanding the structure of the atom, the periodic table, and chemical bonding).
- Page 2 (Middle-Left):** Shows 'The atom' section, including a diagram of an atom and text explaining the structure of matter.
- Page 3 (Middle-Right):** Shows 'Concentration' section, including a diagram of a solution and text explaining the concept of concentration.
- Page 4 (Right):** Shows 'Practical activity 14.3', including a diagram of a titration setup and text explaining the procedure.
- Page 5 (Far Right):** Shows a 'Chapter summary' section, including a diagram of a chapter structure and text summarizing the key concepts.

AT THE END OF EACH CHAPTER

- **Chapter summary** – visual summaries to help summarise key concepts
- **Chapter exam** – exam-style questions to help students develop exam skills, including deliberate practice in data analysis and making connections across content

The collage displays two pages from the IB Chemistry textbook, illustrating the components at the end of each chapter:

- Page 1 (Left):** Shows the 'Chapter summary' section, including a diagram of an atom and text summarizing the key concepts of atomic structure.
- Page 2 (Right):** Shows the 'Chapter exam' section, including a table of exam questions and a diagram of a chapter structure.

Science as a Human Endeavour – a deep dive on the evolution of science and how it has contributed to and influenced society



Glossary provides explanations of all terms introduced in the text.
Answers provide answers for student reference.



An online learning space that provides students with tailored learning experiences.

- Access tools and content that make learning simpler yet smarter.
- **Flexible formats:** choose how you navigate using either the online eText or offline PDFs.
- Margin links in the student book signpost multimedia student resources found on Nelson MindTap.



Weblink
Error bars

For students:

- **Engage** with the **online eText** by adding notes, highlights and bookmarks and using the **Search** and **Read Aloud** (in Australian voice) functions.
- **Check** your understanding with assumed knowledge reviews.**
- **Explore** key concepts with worksheets and additional online-only practicals.
- Access **supporting materials** such as the chapter summary pack and weblinks.
- **Revise** with chapter tests to practise your skills and build confidence.**
- **Navigate** your own learning path, accessing the content and support resources as you need it.

** Available if assigned by your teacher in Cognero Assess.

For teachers*:

- **Tailor** content to different learning needs – assign directly to the student, or the whole class.
- Use **teaching plans** and curriculum guides to support easy program and assessment planning.
- **Prepare** students for assessment with practice versions of data tests, unit-level exams and the external assessment.
- **Answers** to student book questions help you grade efficiently.
- **Risk assessments** and technical notes are provided for each practical activity to ensure the safety of students.
- **Chapter summary** presentations provided in Powerpoint format let you edit and build on them.
- Help build your students' exam readiness with **Cognero Assess** – a test bank containing hundreds of questions and answers to create, assign or export formative and summative tests.

* Complimentary access to these resources is only available to teachers who use this book as part of a class set, book hire or booklist. Not available for single purchases. Alternatively, teacher access can be purchased as a separate subscription for schools. Contact your Cengage learning consultant for information about access and conditions

Security & privacy:

Nelson MindTap joined the Safer Technologies 4 Schools (ST4S) Product Badge Program in 2024. The annual ST4S assessment is part of our commitment to supporting the online security and safety of students and schools. Learn more at st4s.edu.au



Shutterstock.com/Phattana Stock

SYLLABUS
DOT POINTS**SCIENCE INQUIRY SKILLS**

Throughout the course of study, students will:

- identify, research and construct questions for investigation
- propose hypotheses and/or predict possible outcomes
- 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, sample size, 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
- 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
 - material safety data sheets
 - workplace health and safety guidelines



- appropriate disposal methods
- standard operating procedures
- acknowledgement of sources and referencing
- 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
- use scientific language and representations to systematically record information, observations, data and measurement error, e.g.
 - symbols, units and prefixes
 - tables, graphs and diagrams
 - indicators of measurement uncertainty and state measurement uncertainties as a range ($\pm$) 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.
 - logbooks
- translate information between graphical, numerical and/or algebraic forms, e.g.
 - units and measurement conversions
 - ratios and percentages
 - symbols and notation
- 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 in processed data, including the use of absolute uncertainties of the mean $\left(\text{Formula: } \Delta \bar{x} = \frac{\pm(x_{\max} - x_{\min})}{2} \right)$ and percentage uncertainties $\left(\text{Formula: percentage uncertainty (\%)} = \frac{\text{absolute uncertainty}}{\text{measurement}} \times \frac{100}{1} \right)$
 - calculate the percentage error, when the experimental result can be compared with a theoretical or accepted result (value) $\left(\text{Formula: percentage error (\%)} = \left| \frac{\text{measured value} - \text{true value}}{\text{true value}} \right| \times \frac{100}{1} \right)$
 - 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
 - discriminate between precision and accuracy
 - identify that all measurements have limits to their precision and accuracy that must be considered when evaluating experimental results
 - interpret graphs in terms of the relationship between dependent and independent variables; 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
- select, synthesise and use evidence to construct scientific arguments and draw conclusions
- extrapolate findings to determine unknown values, predict outcomes and evaluate claims
- 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 acceptance, bias, status, appropriateness and reasonableness
 - compare findings to theoretical models or expected values
 - discriminate between validity and reliability
- 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
- 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
- appreciate the role of peer review in scientific research.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Conducting structured experiments is an important part of science because it allows for the gathering of information to help develop a greater appreciation and understanding of the world.

Performing research on scientific topics and the process of developing and implementing experimental methods form a large part of the scientific studies and formal internal assessment throughout this course. In fact, the Student Experiment (IA2) and Research Investigation (IA3) make up a significant portion of the internal assessments, incorporating both primary and secondary data. Furthermore, the scientific thinking acquired through these processes is regularly examined in the External Assessment. As such, it is important to develop these skills not just for this course, but also to improve critical thinking.

ASSUMED KNOWLEDGE

- ✓ The purpose of experiments is to collect information about a key idea or to answer a question.
- ✓ Controlled experiments have a general structure.
- ✓ Variables are factors or conditions that can be changed, controlled or measured and which can influence the result of an investigation.
- ✓ Variables include independent, dependent and controlled variables
- ✓ The data collected from an experiment needs to be related to the question being investigated.
- ✓ Data collected from an experiment can be presented in different ways depending on the nature of the data.
- ✓ Data can be classified as primary or secondary.

LEARNING OUTCOMES

- ✓ Develop research questions.
- ✓ Identify the importance of peer review in scientific research and compare different ideas and information from scientific texts.
- ✓ Plan and modify investigations, including the materials and methods needed to collect valid and reliable data (both primary and secondary data).
- ✓ Consider safety, ethics and the environment when conducting scientific investigations.
- ✓ Determine the best method to present data; for example, tables and graphs.
- ✓ Use scientific language and visual representations to organise and present information accurately.
- ✓ Identify and minimise errors in measurements.
- ✓ Calculate uncertainties and other measures of data accuracy and describe their impact on data.
- ✓ Select and construct the most appropriate data presentation technique.
- ✓ Make predictions based on trends observed in the data.
- ✓ Use mathematical techniques to analyse data to find patterns, trends and relationships, taking into account any limitations or sources of error or uncertainty.
- ✓ Draw conclusions based on evidence, comparing findings to expected results.
- ✓ Communicate scientific information in a clear and appropriate manner for different audiences, while acknowledging sources and using proper referencing.
- ✓ Reflect on investigations and suggest ways to improve the quality and accuracy of data and findings.

DC.1 Student experiment

research question a question that directs the scientific inquiry activity; it focuses the research investigation or student experiment, informing the direction of the research, and guiding all stages of inquiry, analysis, interpretation and evaluation

Forming

The research question

For your Student Experiment (IA2), you will be required to design an experiment to answer a **research question** related to a topic in the syllabus. In science, the design of experiments is guided by the scientific method (**Figure DC.1.1**) – a systematic and structured approach that ensures that the results are objective, accurate and reliable.

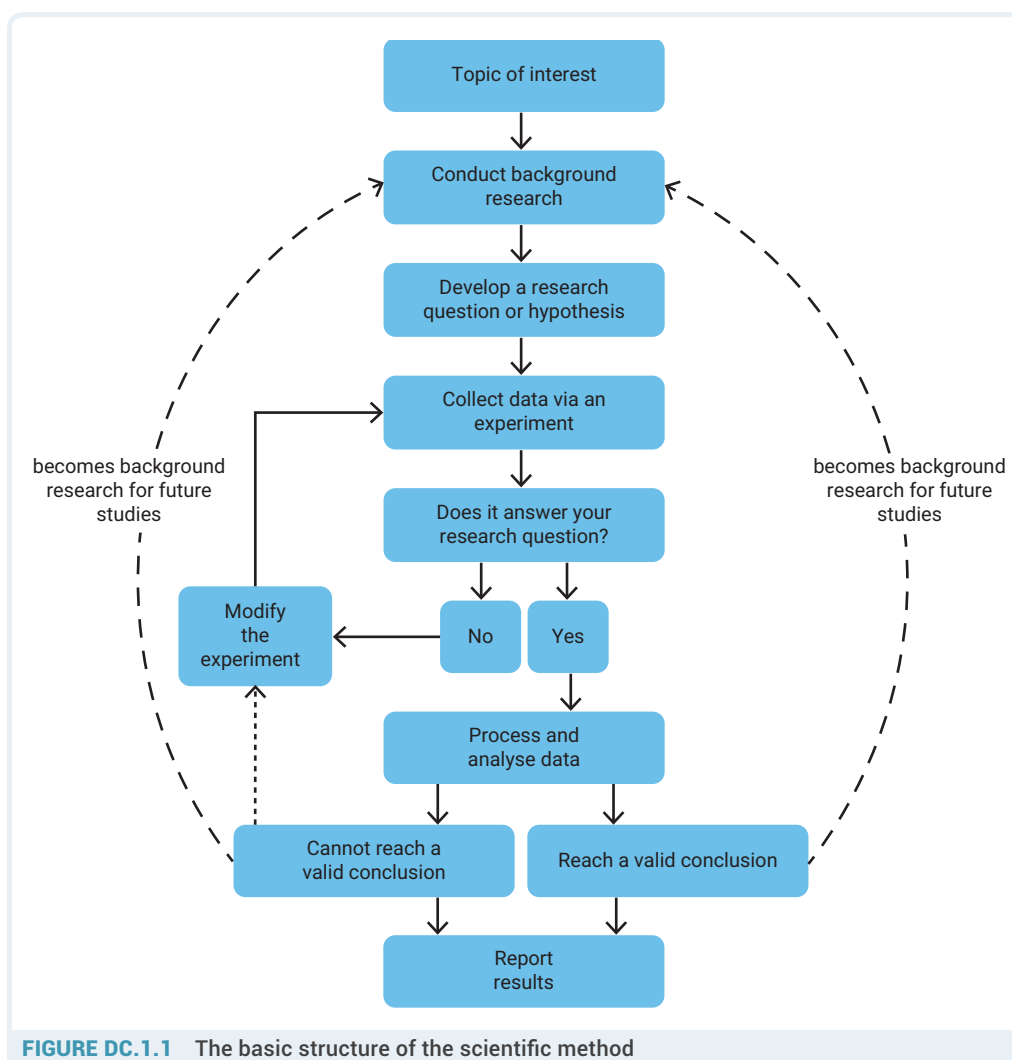


FIGURE DC.1.1 The basic structure of the scientific method

A research question is the question you are trying to answer with your research, and by doing so helps to guide and refine the research and experimental method. For example, a research question could be ‘How does a fever of more than 38°C affect the function of enzymes in the body?’

Rationale

When developing a research question, it is important to demonstrate an understanding of the underlying theory related to the topic. This is described in the rationale of your experiment and is also implied through your research question. In the example above, the research question explores the relationship between enzymes and temperature. Given that it is known that the function of enzymes can be affected by changes in temperature, exploring the impact of an increase in temperature (because of a fever) expands our understanding of enzymes.

With an understanding of the topic, it is likely that you have developed a possible answer to the research question. As you know, this is the hypothesis. The hypothesis highlights the relationship between the **independent variable** and the **dependent variable**, showing the directional impact that one would have on the other. A possible hypothesis to the example above is ‘As the temperature of the fever increases, enzyme function would be reduced’. In this case, the independent variable is the change in temperature and the dependent variable is the level of enzyme function, which can be measured by the amount of product produced over time.

independent variable the variable that is purposely changed or manipulated in an experiment

dependent variable the variable that changes due to changes to the independent variable

Methodology

For your experiment, you will need to modify an existing method from previous experiments. During your research, you may have encountered various studies conducted by scientists who were interested in investigating a similar topic. These studies can serve as a valuable foundation for you to build on and refine your own approach. How and what you modify in the experiment will depend on:

- the variables you are testing
- sources of error and bias in the previous method
- the type of data being collected (**quantitative** or **qualitative**)
- how many data points you will need to collect to ensure that there is sufficient data for analysis
- access to resources.

Once you have your base experiment (which could be one that you completed in class), you will be required to make some modifications to design your own student experiment. A modification may be one of three types (**Table DC.1.1**).

TABLE DC.1.1 Types of modifications that can be implemented for the Student Experiment (IA2)

Type of modification	Explanation	Instruction	Example
Refine	To improve by making subtle changes to the accuracy or precision of the data	• Make improvements without changing the independent or dependent variables.	• Use equipment with a higher level of precision. • Improve the methodology or way of measuring the independent variable. • Change the sample size.
Redirect	To gain further insight by changing the course or direction of the data	• Change the independent variable.	• Measure pH instead of temperature. • Use a different species. • Use different chemicals.
Extend	To change or extend the scope of the current data range	• Change the range of the independent variable. • Extend the range of independent variables.	• Use more concentrations of solution. • Use more sample categories or data ranges.

Although you will not need to write your entire methodology in your experiment, it is important that you can justify the modifications that you made. For example, if you decided to refine an experiment by using a digital thermometer instead of an analogue thermometer, you could justify this by saying that your refinement will improve the accuracy of your data collection because the digital thermometer has a smaller uncertainty and removes human error and/or that it has greater precision.

Having the right equipment is important; however, misreading the measuring instruments leads to inaccurate results. For example, when reading the volume of a liquid in a measuring cylinder, the measurement must be taken at eye level and measured from the bottom of the meniscus (**Figure DC.1.2**).

qualitative data
information that is not numerical in nature

quantitative data
numerical information

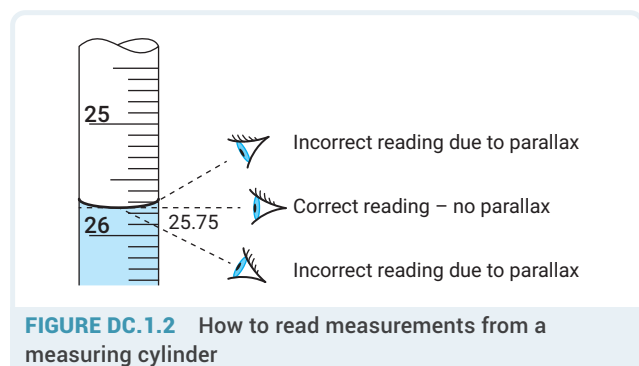


FIGURE DC.1.2 How to read measurements from a measuring cylinder

When collecting and recording data, ensure that it is measured in the appropriate units. For example, the concentration of solutions can be measured in different units such as g L^{-1} , mol L^{-1} or ppm (parts per million). However, this will depend on the nature of the experiment. Since there are different units to express the same measurement, you will often need to convert between units, especially when analysing results. **Table DC.1.2** shows common unit conversions.

TABLE DC.1.2 Common unit conversions

Measurement	Common conversions
Distance	1 km = 1000 m = 100 000 cm = 1 000 000 mm
Mass	1 kg = 1000 g = 1 000 000 mg
Volume	1 L = 1000 mL = 1000 cm ³

Essentially, any modifications to the methodology are done to improve the reliability of data and validity of the experiment. Although you will not need to show your full methodology, you will need to explain and justify any modifications and refinements in your final presentation of your experiment.

If the proposed experiment involves exploring knowledges of First Nations peoples, it is extremely important to understand all the cultural guidelines and protocols involved in conducting such research. For example, the Australian Institute of Aboriginal and Torres Strait Islander Studies (AIATSIS) published the *Code of Ethics for Aboriginal and Torres Strait Islander Research* outlining the principles that underpin ethical Australian Indigenous research to help guide research that impacts on or is of significance to Aboriginal and Torres Strait Islander peoples.



Weblink
Ethical research

Resource
Risk assessment

Finding

Health and safety

Health and safety are important considerations for practical exercises in all sciences. When undertaking your own practical research investigations, you must consider any relevant workplace health and safety guidelines. In Queensland, this includes the *Work Health and Safety Act 2011*. As the researcher, you must ensure safe laboratory practices when planning and conducting investigations by using risk assessments, supported by material safety data sheets (MSDSs), and accounting for risks. MSDSs are important when you are using chemicals as part of your investigation. This includes both the use and the disposal of any potentially harmful materials used and produced in your experiment. Even if your research does not use chemicals but requires participants to take some actions that may cause harm, you will still need to complete a risk assessment form (**Figure DC.1.3**). Your school is likely to have one of these documents for you to complete when you conduct your experiment. If you are unsure of the ethical, environmental, or health and safety aspects of your experiment, check with your teacher. Not only are the risks inherent in the equipment to be identified, but the steps to reduce or manage the risk need to be stated.

Science investigation risk assessment for
Nelson Science 10 Nelson MindTap

Chapter 1

School			
Name of teacher/technician		Year level/class	
Date			

Name of investigation/activity	Karyograms
Book reference	Nelson Science 10, Chapter 1, Module 1.2, downloadable/PDF science
Activity type	☐ Demonstration ☒ Student activity
Description of activity	Students sort out chromosomes into an ordered karyogram.

Equipment		
Equipment to be used	Potential hazards	Control measures/safe handling procedures
- A printout of the karyograms investigation sheet - Scissors - Glue - Blank sheet of paper	☐ Electrical ⚡ ☐ Radiation ⚡ ☐ Thermal ⚡ ☒ Sharp ☐ Projectile ☐ Glass ☐ Gravity – Weights or magnets ☐ Other –	☐ Safety glasses ☐ Sharps container ☐ Thermally insulated gloves ☐ Signage ☐ Safety shield ☐ Other –

Chemicals		
Chemicals to be used	Potential hazards	Control measures/safe handling procedures
None	☐ Explosive ⚡ ☐ Flammable ⚡ ☐ Oxidising ⚡ ☐ Gases under pressure ⚡ ☐ Corrosive ⚡ ☐ Acute toxicity ⚡ ☐ Chronic health hazards ⚡ ☐ Health hazards ⚡ ☐ Environmental ⚡ ☐ Other –	☐ Ventilation ☐ Fume cupboard ☐ Safety shield ☐ Safety glasses ☐ Lab coat ☐ Gloves ☐ Limit concentration/quantity ☐ Other –

FIGURE DC.1.3 A section of a risk assessment form

Any risks that you identify need to be highlighted in your experiment, including the steps to mitigate these risks (**Figure DC.1.4**).

1.4 Management of risks

The overall experiment was given a low–medium risk due to several safety hazards. An over-heating power supply may cause melting to outer-plastic and can shut down – affecting connected outlets to the supply (Hill 2021). Consequently, the power supply was shut off every 2 minutes and was placed over a heat-resistant mat to eliminate heat-transfer and to allow a risk-free 8V supply. Furthermore, many power cables were connected to walls, computers and other equipment throughout the procedure, hence a safety hazard for a potential “trips and falls” in the laboratory-safety-procedure section (Safety 2013). Thereby, chairs were placed over all wires – to caution to anyone in near premises.

FIGURE DC.1.4 An example of the inclusion of risks in an experiment

Apart from highlighting any potential dangers, another way to reduce the risk of injury and improve safety is to clearly outline the procedures in the experiment. This also includes the proper use of any materials involved in the experiment. This can be referred to as the standard operating procedures of an experiment.

Ethics

Ethics is a guiding framework that all research investigations must follow. Ethical concepts provide moral guidance for making decisions about the design and implementation of a research investigation. Examples of ethical concepts are shown in **Table DC.1.3**.

TABLE DC.1.3 Descriptions of different ethical concepts

Concept	Description
Beneficence	Having a commitment to do good (and minimise risk and harm)
Integrity	Acting with honesty and transparency
Justice	Ensuring fair distribution of benefits, risks, costs and resources
Non-maleficence	Avoiding harm or ensuring that potential harm is outweighed by benefits
Respect	Respecting individual differences and ensuring the right to autonomy and choice

You must apply your ethical understanding throughout your study of science, particularly for your own research.

Analysing data

primary data data collected directly by a person or group

logbook a complete, permanent record of how an experiment or research project was conducted; it shows what was done at every step along the way

The **primary data** collected in the experiment should be first organised into a raw data table. When constructing these tables, the independent variable is usually expressed in the first column and the dependent variables from the trials in the experiment are placed in the subsequent columns. For example, when measuring the time taken for a current to pass through a solution of different concentrations, the different concentrations (the independent variable) are presented in the first column, and the time measurements (dependent variable) are presented in subsequent columns (**Table DC.1.4**). It is often suggested that you record everything in your **logbook**; however, this is not a mandatory component of your investigation.

TABLE DC.1.4 An example of a table of raw data from an experiment

Concentration (mol L ⁻¹)	Time (± 0.5 s)		
	Trial 1	Trial 2	Trial 3
0.2	359.5	368.5	364.5
0.4	360.0	345.5	327.5
0.6	325.5	339.5	333.5
0.8	343.5	307.0	327.5
1.0	307.0	339.5	326.5

QCAA Chemistry 2019 v1.3 IA2 high-level annotated sample response August 2018

As shown in Table DC.1.4, the units for each measurement are included in the column headings.

Once the data has been collected, the next step is to analyse it. As part of this, we need to make a judgement on the quality of the data in terms of:

- accuracy
- precision
- reliability
- validity
- sources of error.

It is important to note that each senior subject has different and specific forms of mathematical analysis. What is appropriate for one type of data in one subject may not be appropriate for another. For example, calculated means and uncertainties might be appropriate for a Chemistry or Physics experiment, but mean and standard deviations might be more appropriate for Psychology or Biology experimental data. The following information is a general overview of the types of analysis you could undertake, but it is best to check that the type of analysis you choose is appropriate for your data.

Accuracy and precision

In science, the **accuracy** of a measurement is how close it is to the true value of the quantity being measured. Even when the true value is unknown, scientists can rely on the best available **accepted value** to compare with the experimental measurement result to determine its accuracy. Often the accepted value is the theoretical value calculated for the measurement.

A way to help indicate the accuracy of a measurement is to calculate **percentage error**. Percentage error shows us how close the measured value is to the true or accepted value:

$$\text{Percentage error (\%)} = \left| \frac{\text{measured value} - \text{true value}}{\text{true value}} \right| \times \frac{100}{1}$$

A low percentage error indicates a high degree of accuracy, whereas a high percentage error indicates a low degree of accuracy.

accuracy the degree to which the result of a measurement, calculation or specification conforms to the correct value or a standard

accepted value the value of a substance or quantity that is universally agreed as being a best estimate due to multiple and highly accurate measurements

percentage error the difference between a measurement result and an accepted value, expressed as a percentage of the accepted value

KEY FORMULA

Percentage error

$$\text{Percentage error (\%)} = \left| \frac{\text{measured value} - \text{true value}}{\text{true value}} \right| \times \frac{100}{1}$$

WORKED EXAMPLE DC.1.1

A student used a ruler to measure the height of a 100 mL beaker. These beakers are known to have a height of 7.2 cm. The measured value was 6.8 cm. Calculate the percentage error of the measurement.

ANSWER

1 Determine the measured and true values.

Measured value: 6.8 cm

True value: 7.2 cm

2 Substitute and calculate the percentage error.

$$\begin{aligned}\text{Percentage error} &= \left| \frac{6.8 - 7.2}{7.2} \right| \times 100 \\ &= 5.6\%\end{aligned}$$

This suggests that the measurement is slightly lower than the true value.

precision the closeness of several independent measurements of the same quantity to each other

In contrast, **precision** describes how close a set of measured values are to each other. For single measurements, precision is about the level of detail given by the measurement. For example, 0.3 g is less precise than 0.312 g. As you can see, some measuring instruments are more precise than others. This can be due to the:

- technology used in the device
- quality of components
- resolution
- scale.

The ability to measure precise results is important because it can affect the reliability and uncertainty of data. Uncertainty will be discussed in further detail later in this chapter. It is important to note that measurements that are precise are not necessarily accurate (**Figure DC.1.5**).

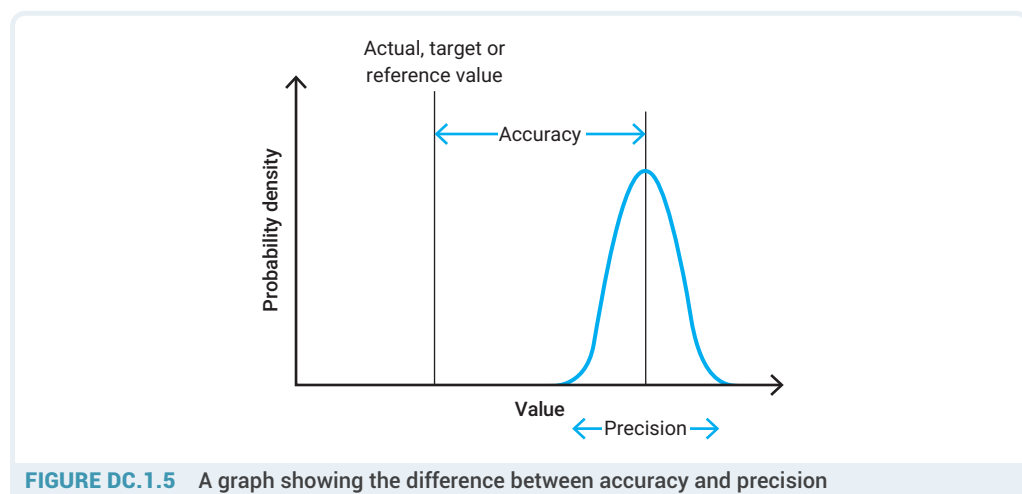
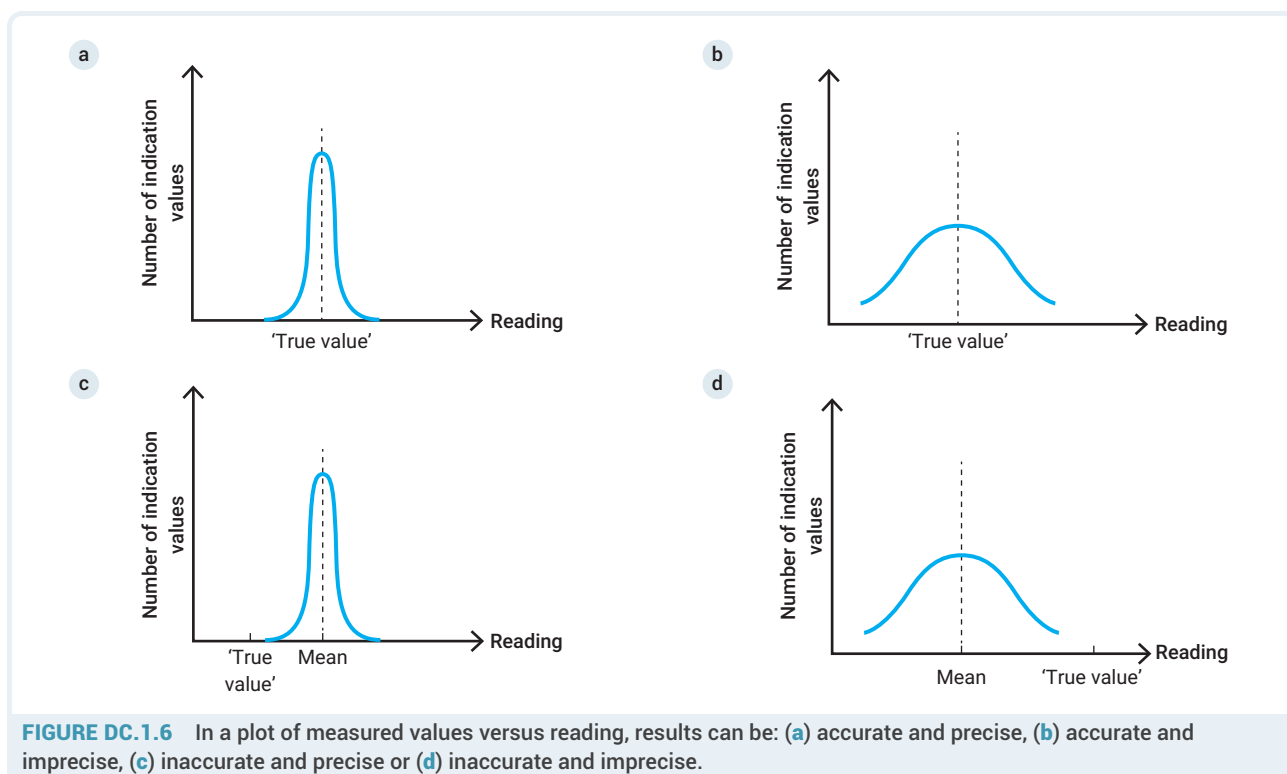


FIGURE DC.1.5 A graph showing the difference between accuracy and precision

Figure DC.1.6 helps further distinguish between accuracy and precision. In parts a and c, the individual indication values cluster closely around the mean, whereas parts b and d show imprecise measurement results because the individual measured values spread significantly around the mean.

For example, for an individual experiencing a fever, having precise measurements of body temperatures of 32.1°C, 33.2°C and 32.0°C does not mean that this is an accurate measure of their



body temperature. We know this because humans have a core body temperature of approximately 37.0°C and fevers cause body temperature to increase, not decrease.

To improve the accuracy of measured values, you could:

- conduct multiple trials and average the results
- ensure that all variables except for the independent variable are controlled (also referred to as fair testing)
- ensure that the measuring tools used in the experiment are appropriate for what is being measured.

This helps to minimise the impact of any errors in the experiment that could affect the accuracy of the measured results. Errors will be discussed in more detail in a later section.

Reliability

If an experiment is repeated, you would expect to obtain very similar results. When this happens, we say that the experimental results are reliable. However, this is not always the case because errors can affect the data collected. **Reliability** can be measured with uncertainty and standard deviation, concepts that will be described in more detail in a later section. Reliability of results can be improved by carefully controlling all variables apart from the independent variable. We will discuss other factors affecting reliability later in this section.

reliability the extent to which the results of assessments are consistent, replicable and free from error

Validity

The quality of the data affects the **validity** of an experiment. We describe data as being valid if the result is due to the independent variable only and can answer the research question. In our example with measuring the effect of temperature on enzyme function, if variabilities in pH in the environment are not properly controlled, they can also affect enzyme function. As a result, we cannot confidently conclude that the results measured from the experiment are due to the changes in temperature only. This type of variable is known as an **extraneous variable** and can affect the relationship between the dependent and independent variables. In this example,

validity the extent to which the experiment measures what it is intended to measure

extraneous variable any variable that is not directly related to the experiment but could affect the results of the experiment

confounding variable a variable that is related to the independent and dependent variables

error the difference between a measured value and true value

random error a variation that affects a measurement in a random way so that successive measured values may reflect small changes from each other

mean the average value of a set of values

pH is a specific type of extraneous variable known as a **confounding variable** because it relates to both the independent and dependent variables. This is why it is important to ensure that all variables other than the independent variable are controlled.

Apart from extraneous variables, **errors** can also affect the results of an experiment. The two main types of errors are random and systematic errors.

Errors

Random errors are unpredictable variations that can occur during measurement. When taking multiple readings of the same thing, random measurement errors cause small variations so that you end up recording a spread of readings. These errors affect the precision of a measurement and can be caused by limitations of measuring instruments. The effect of random errors can be reduced by making more or repeated measurements and calculating the **mean** (or average). To calculate the mean:

$$\text{Mean} = \frac{\text{sum of measured values}}{\text{total number of measurements taken}}$$

KEY FORMULA

Mean

$$\text{Mean} = \frac{\text{sum of measured values}}{\text{total number of measurements taken}}$$

The mean value is then regarded as the most likely or best estimate of the true value; however, we cannot be certain that it is the true value.

While random errors affect the precision of data, **systematic errors** affect the accuracy of a measurement. These errors cause the readings to differ from the true value by a consistent amount in the same direction. This can occur when measuring instruments are not properly calibrated, so readings differ from the true value

by the same amount. Systematic errors can also be caused by observational error if there is a consistent distortion in the way we view things that causes errors that are the same every time. For example, a tall person may read a thermometer from a higher viewpoint and record a lower measure than the true value every time. To minimise the impact of systematic errors, it is important to know how to use measuring tools properly and to calibrate them before use.

Figure DC.1.7 highlights the differences between random and systematic errors.

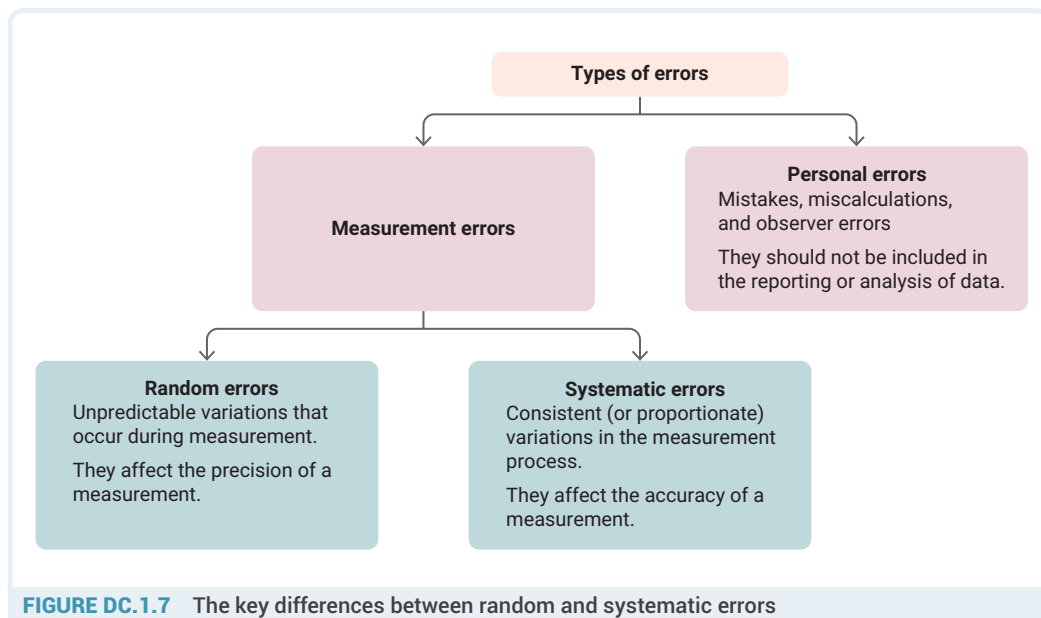


FIGURE DC.1.7 The key differences between random and systematic errors

Uncertainty

While systematic errors can be accounted for by subtracting or adding the value of the error, random errors contribute to the **uncertainty** of a measurement. This reflects the lack of exact knowledge of the true value of the measurement. All measurements are subject to uncertainty because there are many sources of variation. For example, **instrumental uncertainty** in measuring tools can result in variability and imprecision of results due to factors such as sensitivity, calibration and resolution. To minimise the effect of instrumental uncertainty, it is important to calibrate tools, ensure that the appropriate tools and techniques are used, and consider any limitations when designing the experiment. Uncertainty can also occur because of the way the person taking the reading interacts with the tool.

Errors and uncertainties can sometimes be quantified. We can estimate the uncertainty of a measurement, which is usually expressed as $\pm$ a certain value. This is known as **absolute uncertainty**.

Absolute uncertainty of repeated measurements

Most experiments require you to take multiple measurements. As mentioned above, doing so and averaging the results can help to reduce the effect of random errors. Imagine taking multiple measurements of your body temperature. The values are 35.6°C, 36.1°C, 35.9°C and 36.4°C. The difference between the maximum and minimum values is called the **range**. The absolute uncertainty is calculated as the halfway point between the maximum and minimum values, or half of the range:

$$\begin{aligned}\text{Absolute uncertainty} &= \pm \frac{\text{maximum} - \text{minimum}}{2} \\ &= \pm \frac{36.4 - 35.6}{2} \\ &= \pm 0.4^\circ\text{C}\end{aligned}$$

The measurement result would be the mean of the values:

$$\begin{aligned}\text{Mean} &= \frac{35.6 + 36.1 + 35.9 + 36.4}{4} \\ &= 36.0^\circ\text{C}\end{aligned}$$

The reported value includes both the mean and the absolute uncertainty. In this example, the reported value would be $36.0 \pm 0.4^\circ\text{C}$. In other words, the actual value could lie anywhere between 35.6°C and 36.4°C.

Absolute uncertainty of single measurements/device details

For analogue devices, the uncertainty is normally determined as half of the smallest division on the scale. For example, a glass thermometer with graduations of 1°C has an uncertainty of $\pm 0.5^\circ\text{C}$.

With digital devices, the uncertainty is normally defined as the smallest division because we cannot see in between divisions as we can in an analogue device. For example, a digital thermometer that measures in 1°C has an uncertainty of $\pm 1^\circ\text{C}$. One limitation of this calculation is that it does not indicate the direction of the error; we do not know if we overestimated or underestimated.

Percentage uncertainty

Absolute uncertainty can be used to calculate **percentage uncertainty**. Percentage uncertainty is calculated relative to the measured quantity, and is calculated by:

$$\text{Percentage uncertainty (\%)} = \frac{\text{absolute uncertainty}}{\text{measured value}} \times \frac{100}{1}$$

uncertainty the range of values for a measurement result, taking account of the likely values that could be attributed to the measurement result given the measurement equipment, procedure and environment

instrumental uncertainty the inherent limitations and potential errors associated with the measuring instruments or tools used in scientific experiments or observations

absolute uncertainty the magnitude of the difference between the observed/measured value and the true value

range the difference between the maximum and minimum values of a measured confidence interval

KEY FORMULA

Absolute uncertainty

$$\text{Absolute uncertainty} = \pm \frac{\text{maximum} - \text{minimum}}{2}$$

percentage uncertainty a measure of the uncertainty of a measurement compared with the size of the measurement, given as a percentage

KEY FORMULA**Percentage uncertainty**

$$\text{Percentage uncertainty (\%)} = \frac{\text{absolute uncertainty}}{\text{measured value}} \times \frac{100}{1}$$

A lower percentage uncertainty indicates a more precise measurement, whereas a high percentage uncertainty indicates that the measurement is less precise because of greater variability.

Once data has been processed in this way, a table can be presented that also includes these measurements of uncertainty (**Table DC.1.5**).

TABLE DC.1.5 An example of a summary table showing measurements of uncertainty

Voltaic cell	Cathode metal	Potential difference (± 0.05 V)			Mean potential difference (V)	Absolute uncertainty of mean ($\pm V$)
		Trial 1	Trial 2	Trial 3		
1	B(s)	2.25	2.40	2.20	2.28	0.10
2	C(s)	1.30	1.28	1.37	1.32	0.45
3	D(s)	3.11	3.15	3.04	3.10	0.55

QCAA Chemistry 2019 v1.3 IA2 sample assessment instrument August 2018

Significant figures

Once uncertainty has been calculated, results need to be quoted with the appropriate significant figures.

The following steps can be used to determine the significant figures of a single value.

1. Reading a value from left to right, start counting significant figures at the first non-zero number. For example, for a measurement of 0.024 307 g, the first significant figure is 2. In this case, the first two zeros are not considered significant.
2. Every number after the first significant is deemed as significant, including any zeros. In 0.024 307 g, there are five significant figures.

These rules apply even when numbers are expressed in scientific notation. If the above example was expressed as 2.4307×10^{-2} g, then the first significant figure is still 2 and every number thereafter is considered significant.

WORKED EXAMPLE DC.1.2

A student measured the mass of a sample to be 0.0103 g. Determine the number of significant figures in this value.

ANSWER**1 Identify the first non-zero digit.**

Reading from left to right, the first significant number is the first non-zero number. According to this rule, the first significant figure in this value is 1.

2 Count the number of figures in the value that is considered significant.

There are two significant figures 0 and 3 after the first significant figure. As such, there are three significant figures in this value.

When measurements are used to calculate a final value, the numbers and operations to arrive at the final answer contribute to the significant figures of the final answer.

1. For addition and subtraction, the final answer needs to be expressed to the least number of decimal places. For example, when adding the length of two pencils of 10.5 and 9.42 cm, the final answer should be expressed to 1 decimal place, 19.9 cm.

2. For multiplication and division, the final answer needs to be expressed to the least number of significant figures. For example, if we needed to calculate the percentage of sugar in a 50.0 g sample given that there is 2.4 g of sugar present:

$$\begin{aligned}\text{Percentage} &= \frac{2.4}{50.0} \times 100 \\ &= 4.8\% \text{ sugar}\end{aligned}$$

Because 2.4 has the least number of significant figures (two), the answer is expressed to two significant figures. For multistep questions, retain the appropriate number of significant figures at each step, where the final answer should be expressed based on the final step.

WORKED EXAMPLE DC.1.3

A student performed the following calculation using experimental data:

$$\frac{527.11 - 232.3}{5.4}$$

Determine the number of significant figures that the answer should be expressed in.

ANSWER

1 Identify the order of the steps involved in this calculation.

The calculation would be performed in the following order:

- i $527.11 - 232.3$
- ii Answer from step i $\div 5.4$

2 Identify the number of significant figures in the each step.

- i 4. Since it is a subtraction calculation, we need to express the answer to the least number of decimal places.
- ii 2. Since it is a division calculation, we need to express the answer based on the number with the least significant figures (5.4)

3 Determine the number of significant figures for the final answer.

Two

Graphs

Although tables can be an effective way to collect and record data, it is difficult to visualise any trends or relationships between the independent and dependent variables. Presenting data in graphical form makes it easier to identify if any trends exist between the variables.

Many different types of graphs can be used to represent data; for example, column graphs, pie charts, scatterplots and line graphs. Choosing the right graph depends on the nature of the data collected and what you are trying to show. For graphs that involve an x-axis and a y-axis, the independent variable is represented by the x-axis and the dependent variable is represented by the y-axis (**Figure DC.1.8**).

It is also important to choose an appropriate scale when drawing graphs because it helps to ensure that the data is represented in a way that can be easily interpreted. It also avoids misleading representations that imply inaccurate relationships between data. All graphs need to have a title that outlines the information being presented.

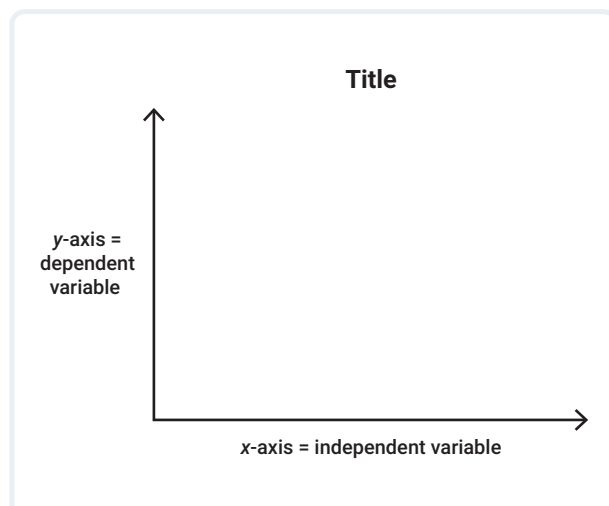


FIGURE DC.1.8 The positive quadrant of cartesian plane showing the variables represented on the x-axis and y-axis

Pie charts

Pie charts are best used to show parts of a whole and the percentage composition of each different category. For example, pie charts can be used to show the composition of a mixture of air – nitrogen, oxygen and other gases – and the percentage of each (Figure DC.1.9).

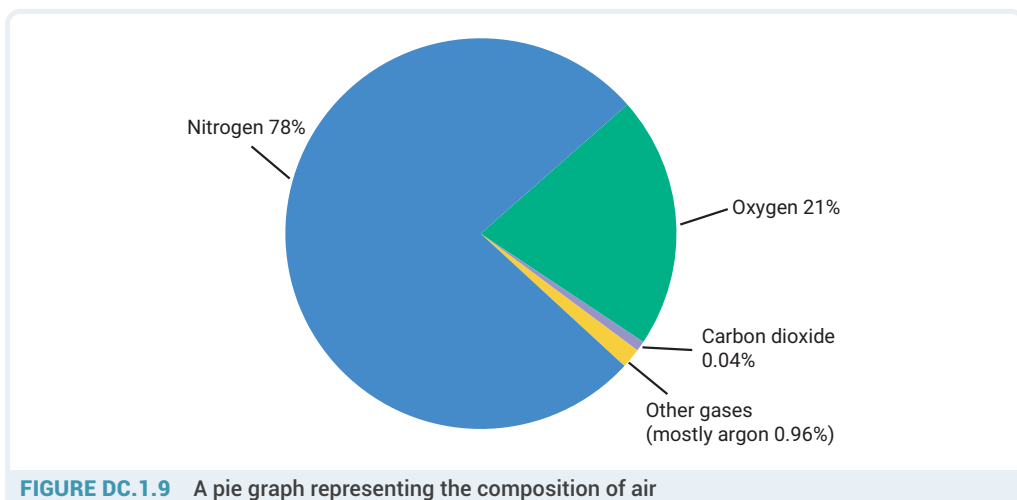


FIGURE DC.1.9 A pie graph representing the composition of air

A limitation with pie charts is that they become visually cluttered when there are many different categories.

Column graphs

Column graphs are useful when comparing quantities or different categories or groups (Figure DC.1.10).

These types of graphs are preferred for comparing categories or to show changes over time, or when comparing the differences between groups.

Line graphs

Line graphs are ideal when showing trends over time for continuous data, particularly when comparing multiple series over the same period. In line graphs, each data point is connected to the next and the relationship between the two variables can be represented as the equation:

$$y = mx + c$$

where: m = the gradient

c = the y intercept.

For example, the calibration curve measuring the absorbance of light based on concentration of a solution can be represented by a linear graph as shown in Figure DC.1.11. If using Excel to draw these graphs, it is sometimes useful to use the x,y scatter plot.

The gradient is a useful piece of information that helps to describe the relationship between the independent and dependent variables. For linear relationships, the gradient of the slope helps to identify the nature of the relationship

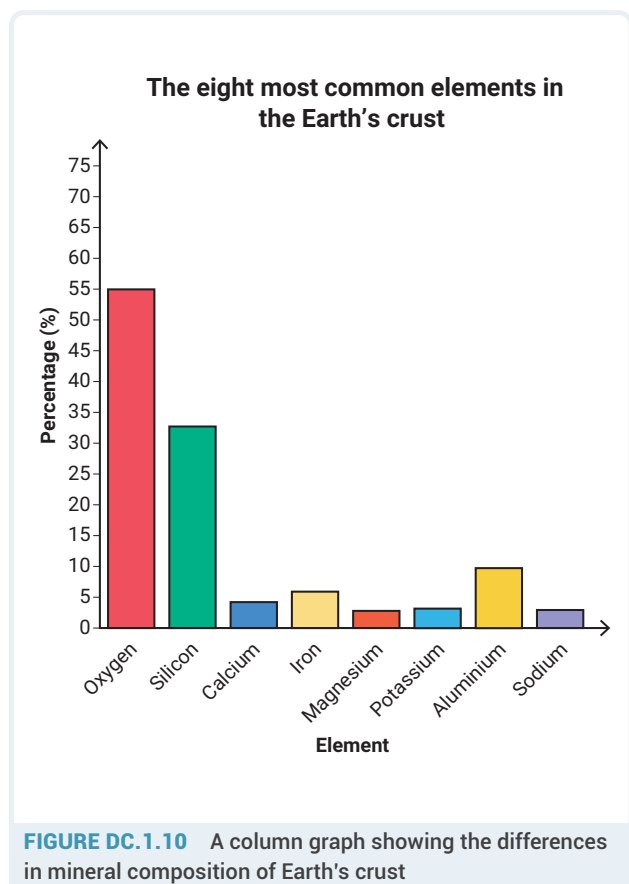


FIGURE DC.1.10 A column graph showing the differences in mineral composition of Earth's crust

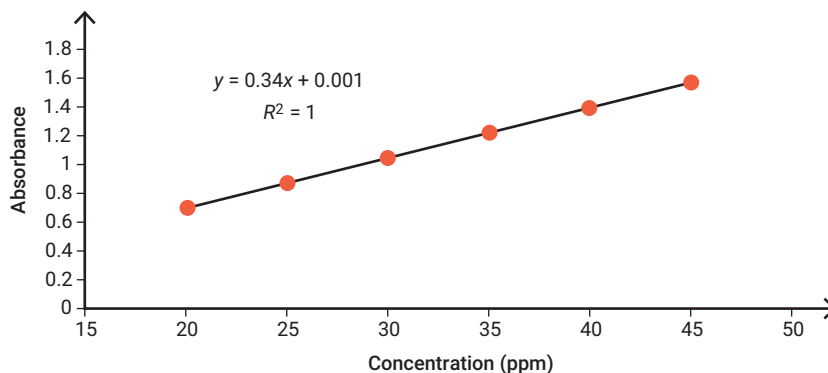


FIGURE DC.1.11 A line graph showing absorbance at different concentrations

between the independent and dependent variables. To calculate the gradient of a linear graph, m , where the equation is $y = mx + c$:

$$\text{Gradient } (m) = \frac{\Delta y}{\Delta x}$$

Determining the gradient in this way only requires two data points, where the difference in the y values is divided by the difference in the x values of the same two points. Depending on the value of the gradient, a:

- positive gradient (**Figure DC.1.12a**) indicates that as the x value (independent variable) increases, so does the y value (dependent variable)
- negative gradient (**Figure DC.1.12b**) indicates that as the x value (independent variable) increases, the y value (dependent variable) decreases
- gradient of zero (**Figure DC.1.12c**) indicates that as the x value (independent variable) increases, there is no change in the y value (dependent variable). As such, there is a constant relationship between the two variables.

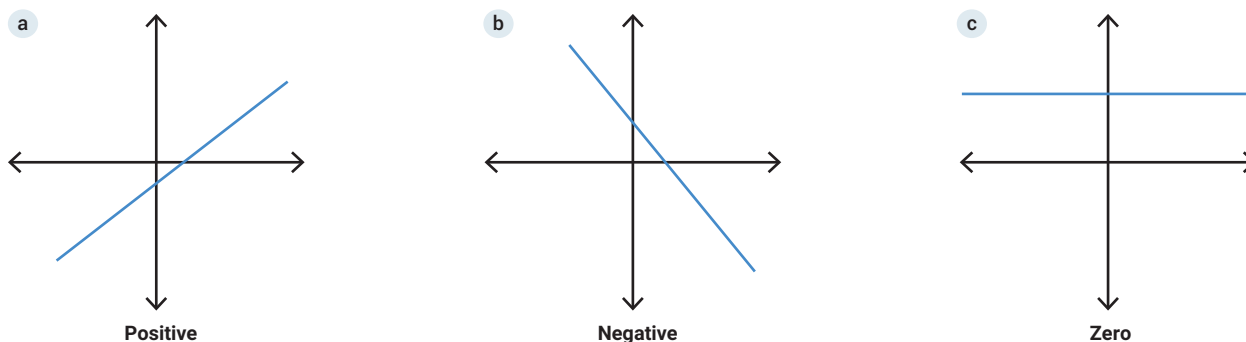


FIGURE DC.1.12 Linear graphs with (a) a positive gradient, (b) a negative gradient and (c) zero gradient

Analysing the gradient for non-linear relationships is a bit more complicated and requires us to calculate the gradient of different tangents at specific points along the graph and compare the changes.

Scatterplots

Scatterplots are similar to linear graphs in that they show individual data points, highlighting the relationship between the independent and dependent variables. However, unlike line graphs, the data points in scatterplots are not connected (**Figure DC.1.13**).

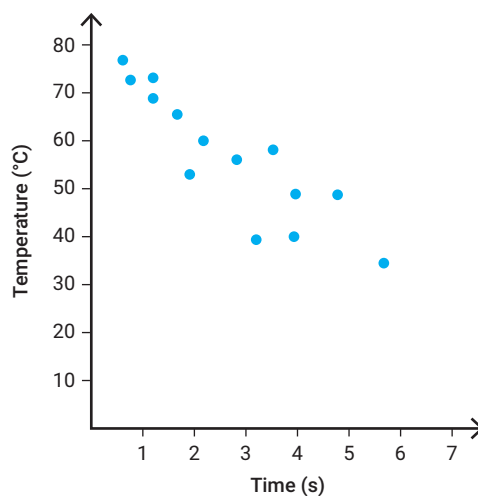


FIGURE DC.1.13 A scatterplot of changes in temperature over time

trendline a line that represents the general direction or pattern of data points

Although the points in a scatterplot are not connected, the way the data points are organised relative to each other can identify a relationship between the variables. **Trendlines** can be drawn through or near the datapoints to help make the relationship between the independent and dependent variable more visible (**Figure DC.1.14**) while also showing the strength of this relationship.

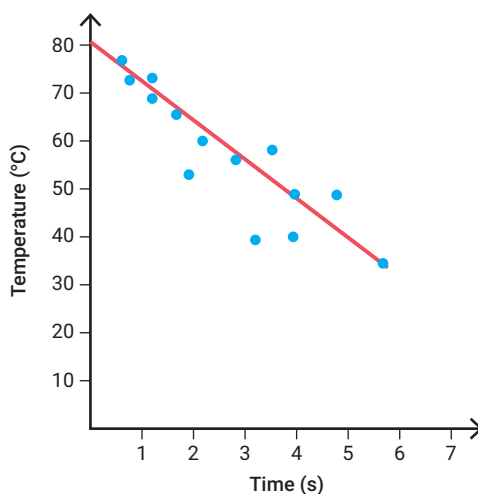
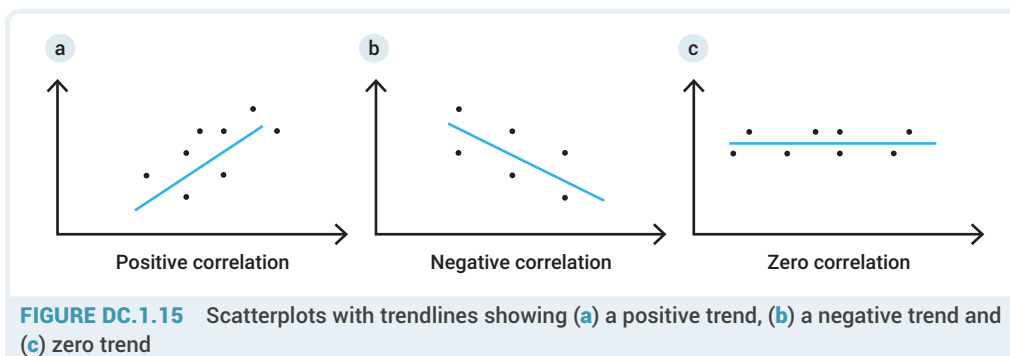


FIGURE DC.1.14 A scatterplot of changes in temperature over time including a trendline (red)

Although it's possible to draw trendlines manually, it is more accurate to use software to draw trendlines. When manually adding trendlines, the line should be drawn so that it minimises the distance between the line and the data points.

If a trend does exist, we can often easily identify whether it is positive (or positive correlation) or negative (negative correlation). In a positive trend, the dependent variable increases as the

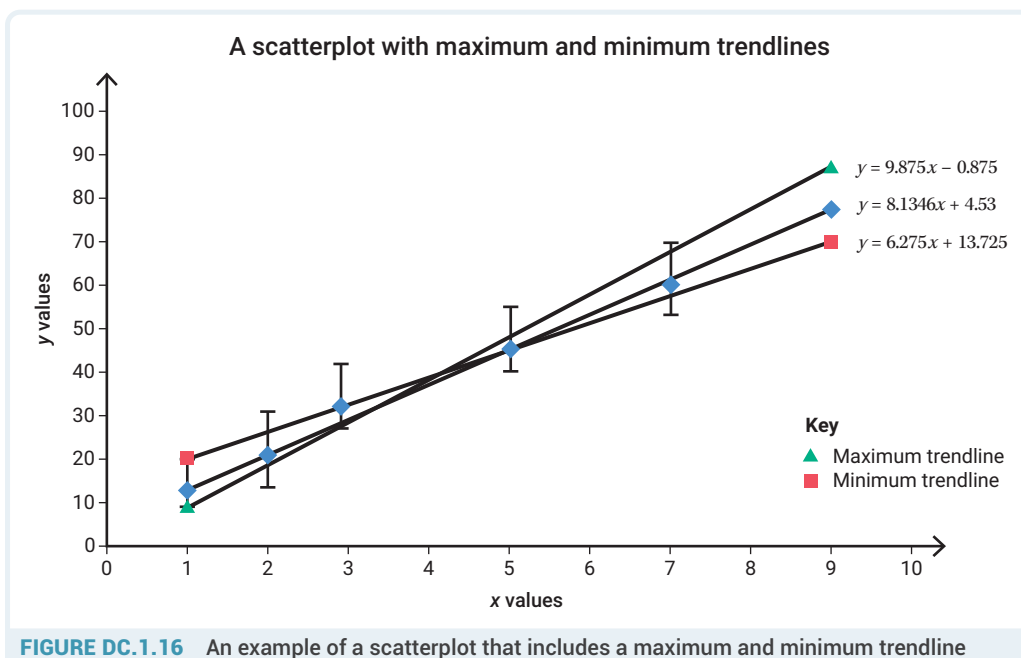
independent variable increases (Figure DC.1.15a), whereas in a negative trend, the dependent variable decreases as the independent variable increases (Figure DC.1.15b).



Maximum and minimum trendlines are visual representations of the strength of the relationship between the variables (Figure DC.1.16). A wider range between the two suggests a greater variability of uncertainty in the data, whereas a narrow range suggests a lower variability in the measured values. Analysing **maximum trendlines** and **minimum trendlines** together can help us predict the potential range of outcomes. For example, using trendlines to forecast temperature changes as a result of emissions can help us predict and prepare for worst-case scenarios. Maximum and minimum trendlines can also help identify potential errors in the experiment. Values that fall significantly outside the area between the maximum and minimum trendlines suggest an outlier that may be due to a random error.

maximum trendline
a trendline with the greatest gradient that fits within the data within the uncertainty values

minimum trendline
a trendline with the smallest gradient that fits within the data within the uncertainty values



Greater variability in certain areas of the graph may also suggest error. For example, when measuring the rate of a reaction at different temperatures, it may be evident that there is a large variability in the rate of the reaction at higher temperatures. This could imply an error in temperature control at higher temperatures.

A common strategy used to draw a maximum trendline involves drawing a line from the bottom of the error bar of the starting data point to the top of the error bar of the last data

line of best fit a straight line through data points in a graph that best expresses the relationship shown in a scatterplot

point. To draw a minimum trendline involves drawing a line from the top of the error bar of the starting data point to the bottom of the error bar of the last data point.

Although trendlines are more general and can be used for different types of graphs with linear and non-linear data, the **line of best fit** is better suited for linear relationships. Since the line of best fit is used for linear relationships, the data points can be used to establish the relationship expressed as $y = mx + c$.

Although this can be done manually, the calculations can become complex and therefore it is often easier (and more accurate) to use software such as Excel, which can both draw the graph and establish the corresponding equation for the line of best fit. Lines of best fit can be used to predict values not measured in the experiment (extrapolation) or estimate values within the range of data collected (interpolation) that was not directly measured (**Figure DC.1.17**).

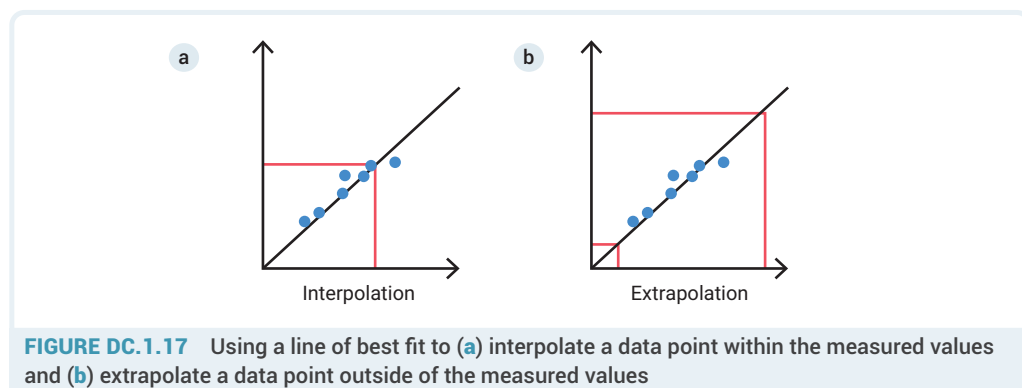


FIGURE DC.1.17 Using a line of best fit to (a) interpolate a data point within the measured values and (b) extrapolate a data point outside of the measured values

Drawing the line of best fit involves specific statistical models such as linear regression and is often accompanied by a quantifiable level of certainty, known as the R -squared value (R^2) (also referred to as the coefficient of determination). Regression analysis provides an equation for a graph so that predictions can be made about the data.

Linear regression is a basic and commonly used type of predictive analysis. Regression is used to examine two key questions:

1. Does a set of predictor variables do a good job at predicting an outcome (dependent) variable?
2. Which variables in particular are significant predictors of the outcome variable?

These regression estimates are used to explain the relationship between one dependent variable and one or more independent variables.

This can be calculated in Excel.

Before discussing R^2 values, we must first understand the significance of R values.

A method that can be used to quantitatively describe the direction and strength of a linear relationship between the independent and dependent variables is the **Pearson correlation coefficient (R)**, which measures the correlation between two sets of data. R values can be between -1 and 1 , where:

- $R = 0$ suggests no correlation
- $R = 1$ suggests a strong positive correlation
- $R = -1$ suggests a strong negative correlation.

The formula to calculate the R value is complicated and therefore it is much easier to use software to help with this calculation. Programs such as Excel have options for calculating R when plotting a graph.

Squaring the R gives us the R^2 value. This value indicates the strength of the correlation of the linear relationship between two sets of data. In simple terms, it answers the question, 'Can I draw a line graph to represent the data?' Calculating this coefficient does not allow you to fit a line to your data (use a regression analysis for this). However, the value is not able to tell the difference between the independent and dependent variables; for example, investigating a



Weblink

Linear regression and Excel

Pearson correlation coefficient (R) a statistical measure that quantifies the direction and strength of a relationship between two variables

high-calorie diet causing diabetes might give a correlation of 0.8. However, you could also get the same result with the variables switched around – diabetes causes a high-calorie diet. Therefore, as a researcher you must be aware of the data you are putting in and note the difference between correlation and causation:

- $R > 0.8 = \text{strong correlation}$.
- $R < 0.5 = \text{weak correlation}$.

Non-linear graphs

Not all trends show a linear pattern. For example, graphs showing the solubility of substances at different temperatures show a non-linear relationship (Figure DC.1.18).

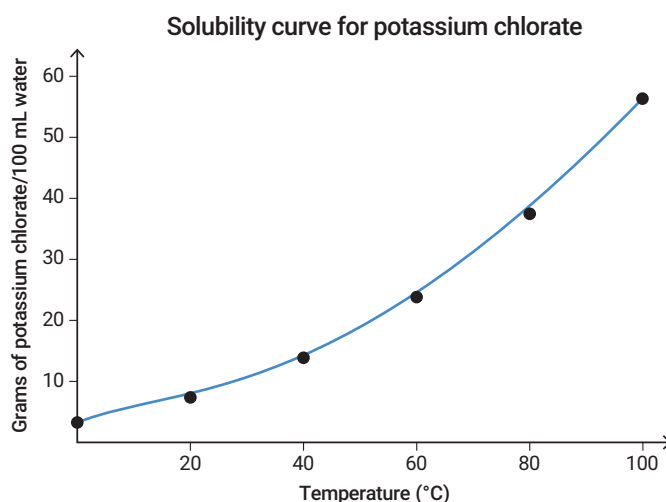


FIGURE DC.1.18 The solubility of potassium chlorate in 100 mL of water at different temperatures shows a non-linear relationship.

The simplest way to identify whether a relationship between two variables is linear or non-linear is to plot the data points on a graph to identify the overall trend. Gradient analysis can be conducted on non-linear graphs by calculating the instantaneous gradient of the tangent line at each data point and comparing the extent of the changes between each point. The tangent line is a straight line that 'touches' the data point and shares the same gradient as the curve at the given data point (Figure DC.1.19).

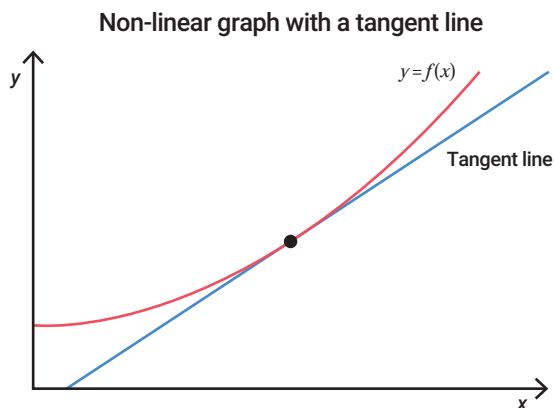


FIGURE DC.1.19 An example of a tangent touching a data point on a non-linear graph

Error bars

To illustrate uncertainty, your graphs should incorporate error bars. These extend from data points to demonstrate the uncertainty the measurement (**Figure DC.1.20**).



Weblinks
Error bars

Drawing graphs
with error bars

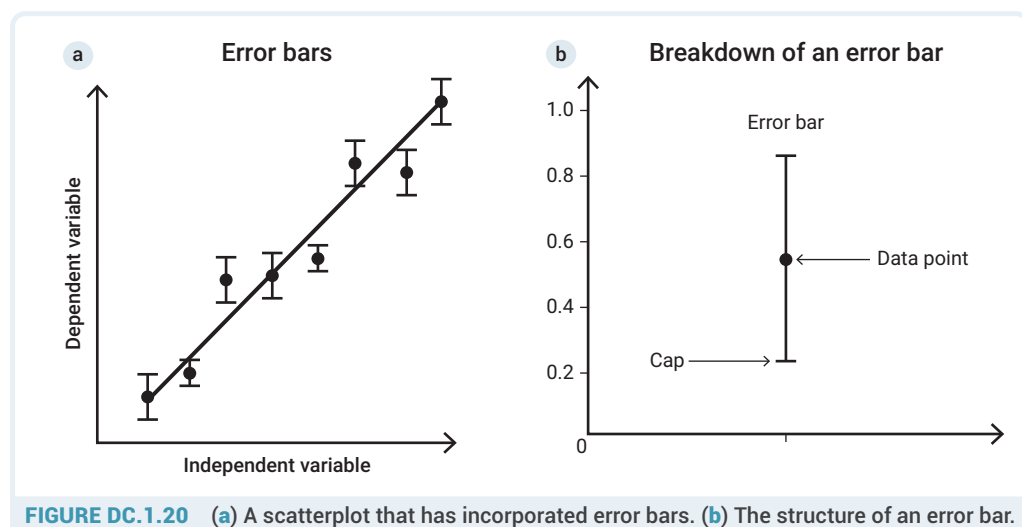


FIGURE DC.1.20 (a) A scatterplot that has incorporated error bars. (b) The structure of an error bar.

The upper and lower limits of the error bars can be determined by using descriptive statistics such as standard deviation or absolute uncertainty. (Note: There are different types error bars, e.g. standard deviation, confidence intervals, absolute uncertainty, percentage uncertainty.) To draw error bars on graphs:

1. Identify the data point.
2. Calculate the uncertainty of the mean for the data point. This will determine the upper and lower limits of the error bar.
3. Use the values from step 2 to identify the maximum and minimum value for the data point. Use this to draw the error bar.

Graphing applications such as Excel have an option to include error bars in graphs. This is a faster and often more accurate method of generating graphs with error bars.

A larger error bar indicates that the values are spread out and suggests greater uncertainty than a smaller error bar, which signals that the measurements are clustered around the data point. Error bars can be drawn for different types of graphs (**Figure DC.1.21**).

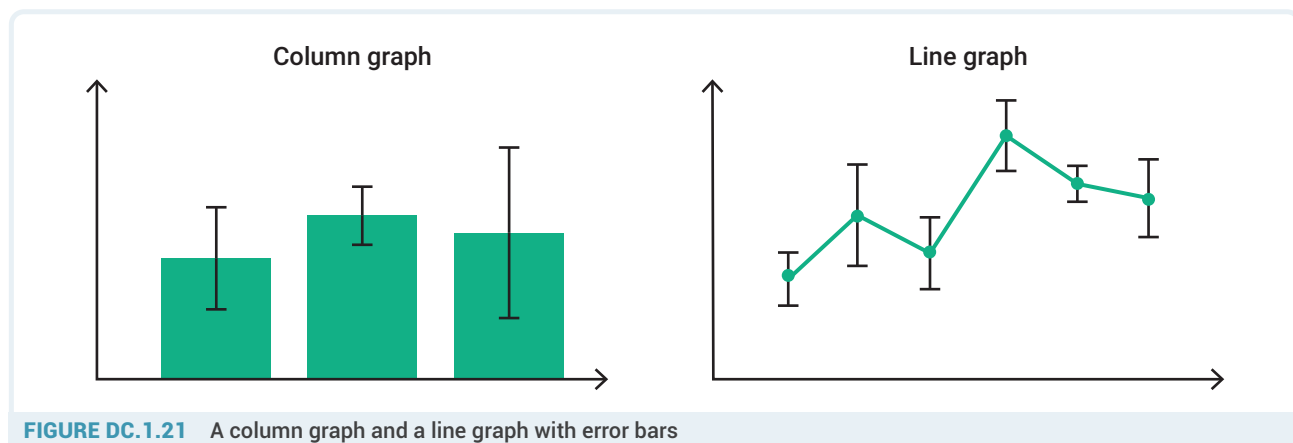


FIGURE DC.1.21 A column graph and a line graph with error bars

Other representations of data

Scientific drawings

Textbooks are full of scientific drawings that represent structures, organisms and processes. These drawings are highly detailed, accurate and clear. For example, consider the representation of covalent bonding between two hydrogen atoms in a hydrogen molecule (Figure DC.1.22).

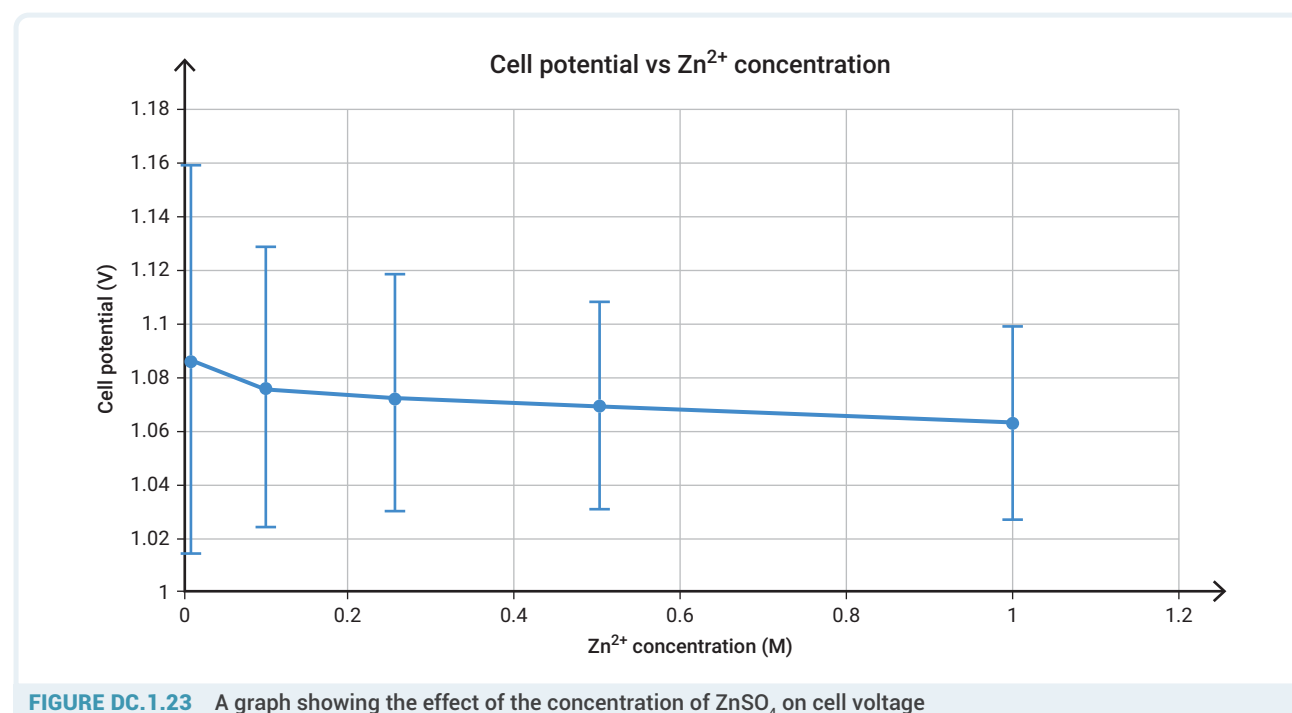
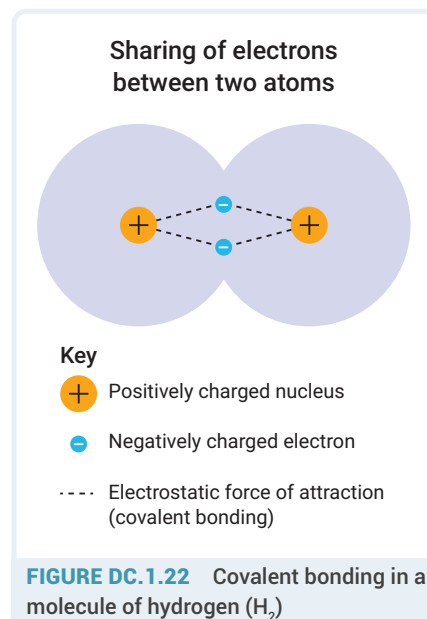
Scientific drawings include labels and annotations and are drawn to scale to show the relative proportions of the elements involved.

Identifying trends and relationships

The purpose of an experiment is to collect relevant data that can be analysed and used to understand the relationship between the independent and dependent variables.

Analysing graphs

When analysing graphs, it is important to consider all aspects presented in the graph. Consider the graph shown in Figure DC.1.23. What we tend to notice first is the overall trend in the data. The graph shows a negative trend where the cell potential decreases as the concentration of Zn^{2+} increases. We determine this visually based on the shape of the line; however, it is possible to also measure the negative gradient for the line. The large R^2 value of 0.9723 (very close to 1) suggests high correlation between both variables.



The graph in Figure DC.1.23 also contains error bars showing the overlap of all data points. When this happens, it suggests that the differences between the overlapping data points are statistically insignificant and are due to random errors. As such, the trend shown may not reflect the true relationship between the concentration and cell potential.

Although not as easy to visually identify as graphs, raw data tables can also be interpreted to identify the relationship between two variables. Consider the raw data table for the graph in Figure DC.1.23 (Table DC.1.6).

TABLE DC.1.6 A data table for the experiment testing the effect of the concentration of ZnSO_4 on cell voltage

ZnSO_4 conc. (M)	Mean cell potential (V)	Absolute uncertainty (V)	Percentage uncertainty (%)	Theoretical E_{cell} (V)	Absolute error (E)	Percentage error (%)
0.01	1.087	± 0.005	± 0.460	1.159	± 0.072	± 6.249
0.10	1.077	± 0.005	± 0.464	1.129	± 0.053	± 4.681
0.25	1.073	± 0.005	± 0.465	1.117	± 0.044	± 3.977
0.50	1.070	± 0	± 0	1.108	± 0.039	± 3.506
1.00	1.063	± 0.005	± 0.470	1.100	± 0.037	± 3.333

All results have an overall low percentage error, suggesting that the experiment has high validity.

Interpreting and evaluating evidence

To demonstrate that you have a robust understanding of the results of the experiment, it is important to identify the pattern/relationship between the variables and comment on the reliability and validity of the relationships using your calculations of errors and uncertainty. Although it sounds counterintuitive, highlighting sources of error in your experiment and describing its effect on your results strengthens your argument. It also allows you to identify any limitations and offer suggestions for improvements and/or extensions to your experiment. By doing so, you are demonstrating an ability to critically analyse data, which helps to develop well-informed arguments.

You need to scientifically justify the argument. This is done by referencing theory and previous studies to explain the phenomena being shown through the data. For example, in an experiment measuring the effect of temperature on enzyme function, we would want to refer to the theory relating to the current understanding of enzyme function and use that to justify the arguments made from the trends identified in the data.

The culmination of this allows us to draw well-informed conclusions that help to answer the research question.



Resource
An example student
experiment

LEARNING CHECK DC.1

DESCRIBING

- Describe** the difference between:
 - accuracy and precision
 - reliability and validity.
- Identify** two strategies to improve the accuracy of data.
- Identify** the type of data most suited to:
 - pie charts
 - line graphs
 - column graphs.

4 **Describe** the importance of:

- a a logbook
- b MSDSs
- c ethics in experiments.

5 **Sequence** the following sections of the scientific method in the correct order: methods, materials, discussion, research question.

APPLYING

6 **Consider** the following research question:

'How does increasing the temperature (K) affect the rate of a reaction (mL s^{-1})?'

Identify the:

- a dependent variable
- b independent variable.

7 A student wanted to conduct an experiment to see whether eating food before running had any effect on how far she could run. Write a research question for this experiment.

8 A student is conducting an experiment involving the use of a glass measuring cylinder to measure and pour a sample of acid into a 100 mL glass beaker. **Identify** one safety concern associated with the experiment and how the risk can be minimised.

9 In a medical experiment, a participant was asked to undergo a series of additional tests that could reveal sensitive information about their health situation. The participant refused to give consent to the tests. However, the experimenter ignored this and requested for the tests to be conducted anyway. Which ethical concept has the experimenter breached? **Explain** your answer.

10 **Identify** the number of significant figures in the following.

- a 0.0023
- b 2.43007
- c 8.1005
- d 0.7

11 A student measured a value of 20 cm in their experiment with an absolute uncertainty of 1 cm.

- a **Calculate** the percentage uncertainty.
- b What does this value suggest about the precision of the measurement?

12 A group of students designed an experiment to measure the melting points of different substances. The students obtained five different substances: A–E. They set up a melting point apparatus and heated each substance until it melted, recording the temperature at which melting occurred.

- a **Identify** an extraneous variable for this experiment.
- b Each student took turns measuring the melting point of each sample each day. Students brought in their own thermometers, and it was noticed that some used digital thermometers while others used analogue thermometers. **Identify** the type of error that occurred as a result of this.

13 In a particular set of measurements, a student recorded the following measurements: 14.2 cm, 14.1 cm and 14.3 cm. **Calculate** the absolute uncertainty.

14 The results from a student's experiment are shown below.

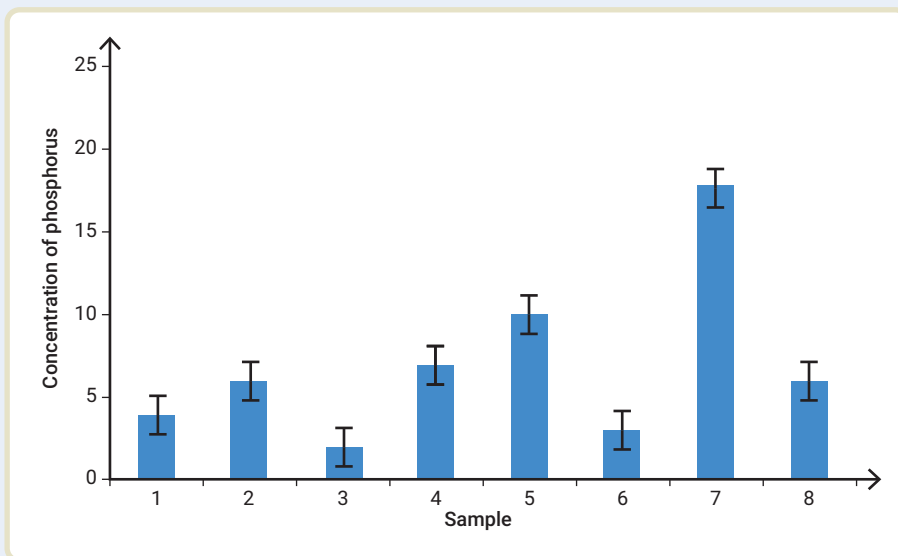
Concentration of reactant (mol L^{-1})	Trial 1 Rate of reaction ($\text{mol L}^{-1} \text{s}^{-1}$)	Trial 2 Rate of reaction ($\text{mol L}^{-1} \text{s}^{-1}$)	Trial 3 Rate of reaction ($\text{mol L}^{-1} \text{s}^{-1}$)
0.1	0.0050	0.0060	0.0050
0.2	0.010	0.011	0.0090
0.3	0.015	0.017	0.016
0.4	0.020	0.022	0.021

- a **Identify** the dependent and independent variables for this experiment.
- b **Calculate** the mean value for each concentration.

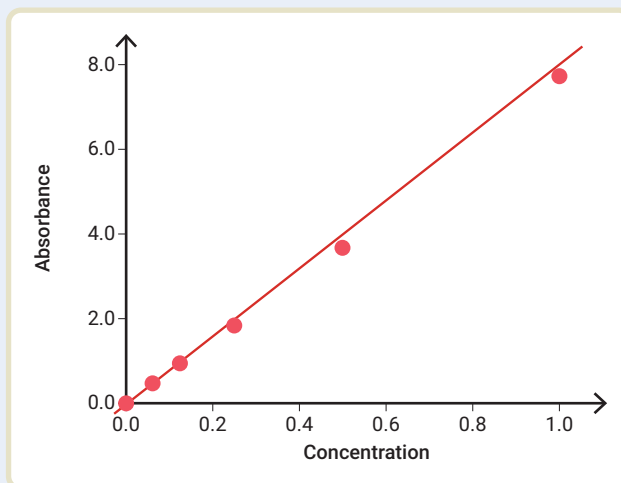
- c Use the values to draw a graph to represent this data.
- d **Determine** the correlation (if any) between the independent and dependent variables.

ANALYSING

15 Consider the following graph.



- a **Determine** the dependent variable.
 - b Which sample shows the greatest variability?
- 16 This graph was drawn for an experiment measuring the change of absorbance as a result of a change of concentration.
- a What is the name of the type of graph drawn?
 - b What is the name given to the line drawn in the graph?
 - c **Calculate** the gradient (m) of the line.
 - d **Determine** the approximate absorbance at a concentration of 0.2.
 - e Based on the graph, would you expect the R value to be closer to 1 or to 0? **Explain** your response.



DC.2 Research investigation

To help prompt your Research Investigation (IA3) assessment, your teacher will provide a list of claims that you can investigate. These claims will be related to particular topics outlined in the syllabus. After selecting a claim, you will be required to choose a research question to investigate. Unlike the student experiment, the research investigation requires you to collect and analyse **secondary data** about your topic and particular research question.

secondary data data that is collected by someone else

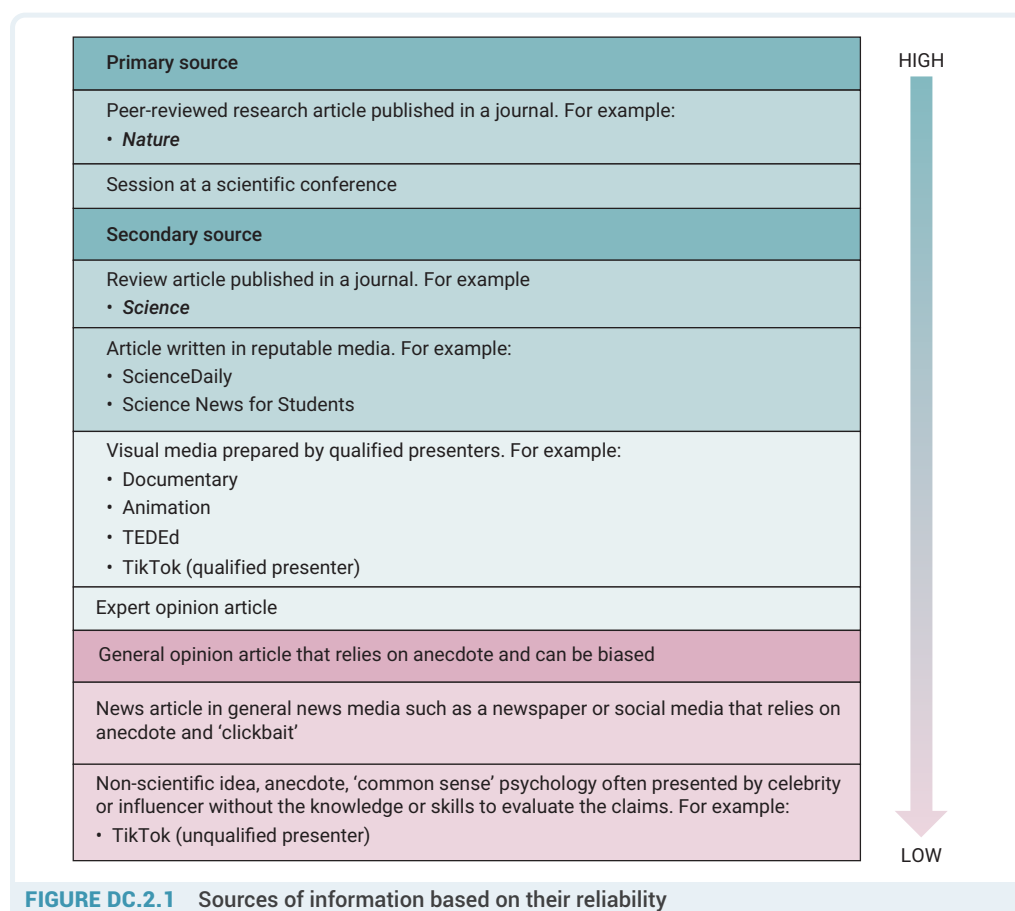
Forming and finding

Researching

As with the Student experiment, you will need to conduct research before developing a research question. This involves reading scientific articles and books and investigating other resources to develop a solid understanding of the topic. From blogs to scientific journals, there are many resources available to help develop your understanding. These sources may be available through open access (e.g. Google Scholar) or through organisations such as government websites and local or national libraries. Open Science, an initiative by Creative Commons, has a list of open-source scientific articles that are freely available. For scientific research, it is important to use a variety of credible resources. Therefore, you will need to be able to assess the reliability of the sources you are using. For example, blogs that can be written by anyone are not as reliable as a scientific article from a peer-reviewed journal (Figure DC.2.1).



Weblink
Open Science



This is why the peer-review process in scientific research is so important (Figure DC.2.2). For an article to be published in a journal, it must be reviewed by multiple experts, who evaluate it and make suggestions for further improvement. Before it can be resubmitted, the author must review and respond to the suggestions. This process can take months. Only once this process has been completed can the article be accepted for publication by the journal.

When using sources that are not from peer-reviewed journals, it is important to assess the reliability and validity of the information. It is helpful to ask yourself question such as:

- Is the author(s) an expert in this domain?
- Does the author use evidence to support their claims?
- Is the methodology valid?

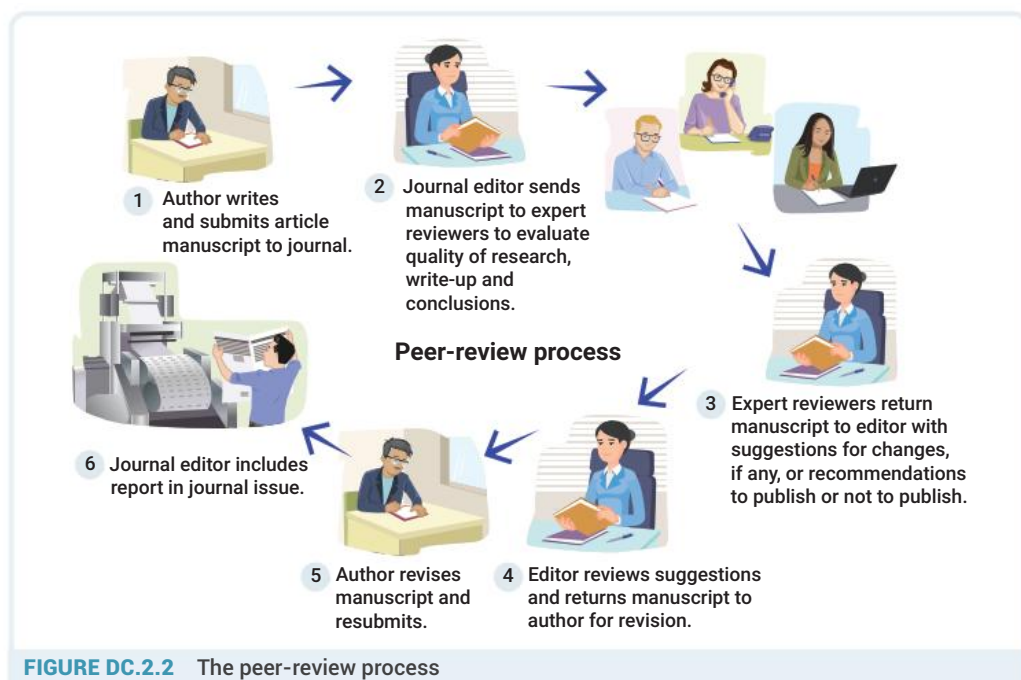


FIGURE DC.2.2 The peer-review process

- If evidence is used, where does the data come from?
- Is this publication trustworthy?
- Is there any bias; for example, is there a conflict of interest among the researchers?

It is also a good idea to cross-reference the information presented by these resources with other sources such as primary sources and textbooks. This initial research helps you to develop a rationale for your investigation, and as a result helps to craft a research question that is relevant to the claim. As with the student experiment, the research question needs to be able to be tested.

To help the reader have the necessary context for the research investigation, you need to provide a level of background. The background needs to include enough of a foundation that the reader can understand the theoretical underpinnings of the research while also showing that you have used scientific evidence to develop a research question that aligns with the claim.

Scientific conventions

The ability to communicate scientific understanding to an audience is often an overlooked skill. How we present the information depends on what we are sharing and the audience we are sharing with. For example, when communicating to a younger audience who are unfamiliar with many scientific concepts, it is important to use language that is accessible, and visuals to help foster a foundational understanding of the topic. When communicating findings to those in the scientific community, we need to use scientific language, including correct nomenclature, units and symbols specific to the scientific theory.

Since experiments and scientific research draw on the knowledge, thoughts and ideas developed by others, we need to appropriately acknowledge the source of the information. The referencing format that is required depends on the discipline; however, in most cases, science uses the APA (American Psychological Association) referencing style.

Analysing evidence

As part of your assessment, you will need to find scientific evidence from previous research related to your research question. The data derived from these studies is used in the same way as the data collected from your student experiment – to identify trends and relationships between variables to answer the research question.



Weblinks

Evaluating the
information you find

Referencing style guides

Referencing sources

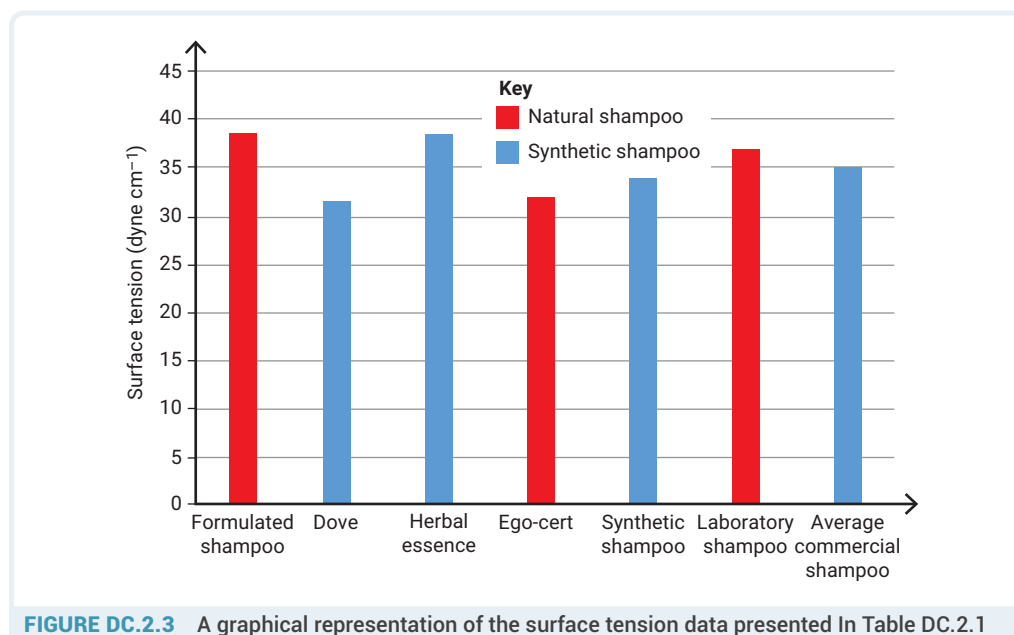
Because there are multiple data points, it is important to re-organise the data and present it in a way that can be analysed. For example, **Table DC.2.1** presents a comparison of data collected from two distinct scientific experiments investigating the properties of different shampoos. The data from all three studies has been organised to make it easier to compare.

TABLE DC.2.1 Comparison of results between different studies

	(Badi 2014)			(Isaiah 2015)		(Mainkar 2001)	
	Formulated shampoo	Dove	Herbal Essence	Ego-cert	Undisclosed synthetic shampoo	Laboratory shampoo	Average of five commercial shampoos
Surface tension (dyne cm ⁻¹)	38.72	31.68	38.36	32	34	37.1	32.8–37.7
Wetting time (s)	187	141	157			130	159–227
Solid content (%)	22.75	25	25	25.3	26.5		
Foam description	Small dense	Small dense	Small airy			Loose open	Dense creamy
Foam height (mL)	115	92	113	165	169	158	153–168
	2% solution			1% solution			
Detergency (% sebum removal)				93	95	61.14	61.1–80.12

QCAA Chemistry 2019 v1.3 IA3 sample assessment instrument August 2018

Additionally, the data can be used to construct graphs (**Figure DC.2.3**).



QCAA Chemistry 2019 v1.3 IA3 sample assessment instrument August 2018

Organising data in this way makes it easier to identify any trends, patterns or relationships that exist between the variables being tested. As in your student experiment, you would also need to identify any sources of error and levels of uncertainty that can impact the results.

Interpreting evidence

As you now know, presenting data and describing trends and relationships on their own is not sufficient. We need to be able to use the evidence and scientific theory to justify the argument being made and to draw a conclusion.

When developing your conclusion, ensure that it directly answers the research question. Sometimes when assessing studies at an individual level, the data may point to a particular conclusion. However, when studies in the same area are evaluated together, an overall analysis may suggest a different conclusion. If your investigation shows a different conclusion from the studies used, that in itself is an important conclusion. It highlights that further investigation is required to develop a deeper understanding of the area.

Evaluating evidence

The quality of evidence can affect the reproducibility of the research and strength of the conclusions drawn. We can assess the quality of the evidence by identifying any limitations caused by errors and/or uncertainty (Figure DC.2.4). This may include assessing:

- appropriateness of the method (where possible)
- sample size
- sources of error
- degree of uncertainty of the data.

It is also important to consider any bias (e.g. is the funding from a particular company?), recency of data and qualifications of the author(s).

The three studies used to determine whether a natural or synthetic shampoo is better at cleaning hair have not led to any conclusive results. The characteristics used to determine whether a shampoo cleans better were limited to the surface tension of the water when the shampoo was added, wetting time, foaming ability, solid content and detergency. In all three studies, the natural shampoo performed better in some characteristics, while the synthetic shampoo in other characteristics. Therefore, the review of the studies done into cleaning ability does not support or refute the claim that natural shampoos are better cleaners than synthetic shampoos.

FIGURE DC.2.4 An excerpt evaluating the evidence provided by previous scientific studies

QCAA Chemistry 2019 v1.3 IA3 sample assessment instrument August 2018

This would also help to make any interpolation or extrapolations of the findings to further analyse the research claim. It is important to suggest any further improvements for future studies related to this area (Figure DC.2.5). For example, identifying any changes that you would make to the methodology to improve the validity or reliability of the data from the experiment.

Further investigations required

To truly investigate whether natural shampoos are better than synthetic shampoos in terms of surface tension, wetting time, solid content, foaming ability and detergency, further testing would need to be carried out. This testing would need to involve a larger sample of both natural and synthetic shampoos. The amount of shampoo would need to remain constant in all testing and be similar to the quantity of shampoo used when washing hair, and the hardness of the water would also

FIGURE DC.2.5 Suggestions for further investigation for an experiment involving shampoos that showed inconclusive evidence

QCAA Chemistry 2019 v1.3 IA3 sample assessment instrument August 2018

Suggestions for improvement should also address any limitations present in the experiment including:

- experimental limitations; for example, time available to conduct the experiment, errors
- methodological limitations; for example, accuracy and reliability of measurement techniques, ethical constraints
- external limitations; for example, environmental factors that can introduce variability, access to proper equipment.

As you can appreciate, being able to interpret and evaluate data is crucial for reaching informed conclusions about your research. Not only do you need to speak about the data, you need to be able to use the evidence to justify any arguments made from the research.

LEARNING CHECK DC.2

DESCRIBING

- 1 **Identify** the difference between primary and secondary data.
- 2 **Describe** the role of the peer-review process in scientific research.

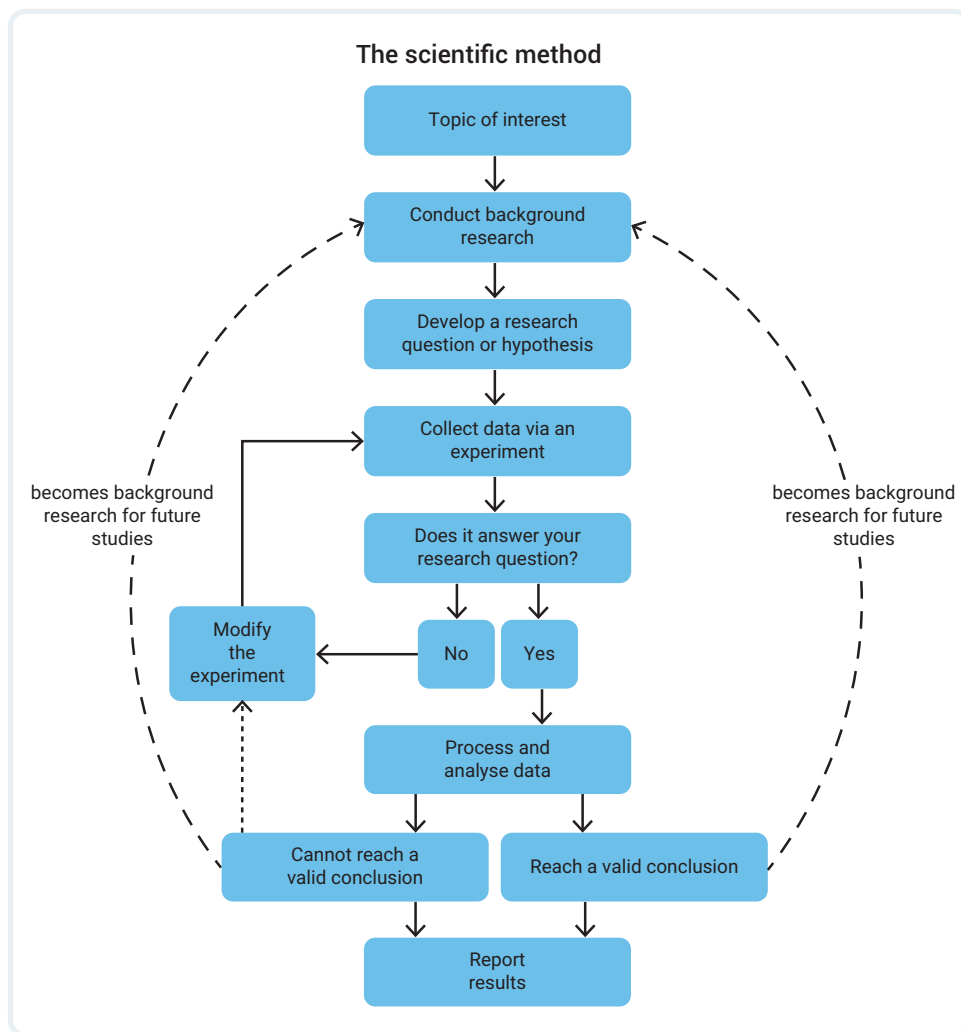
APPLYING

- 3 A student conducted an experiment to investigate the effect of different concentrations of hydrochloric acid on the pH of a solution. The student tested four different concentrations: 0.5 M, 1 M, 2 M and 4 M. The pH was measured using a pH meter. **Identify** one possible experimental and one methodological limitation for this experiment.
- 4 **Consider** the following passage:
The rate of reaction is a fundamental concept in the field of chemistry. Understanding the factors that influence reaction rates is crucial in predicting and controlling chemical reactions.
Rewrite the passage so that it can be read and understood by a primary school student who is studying science.
- 5 A student used the following source for their research investigation.
Cruzan, J. (2012). 'The most important solvent'. Retrieved from <http://www.drcruzan.com/Water.html>.
Use information from the 'Referencing sources' weblink to show how this resource would be referenced in the text.

CHAPTER SUMMARY

Conducting research

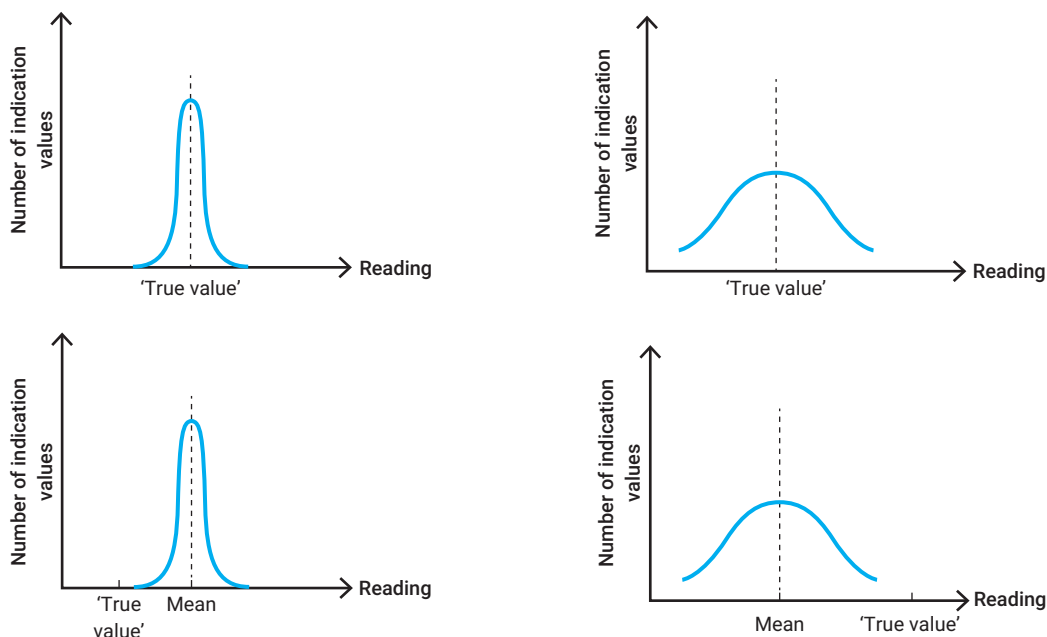
- The scientific method follows a particular process aimed to maximise accuracy, reliability and objectivity while minimising uncertainty and error.



Analysing data

- Precision describes the closeness of data, whereas accuracy describes how close the measured value is to the true value.
- The quality of data affects the validity of the experiment.
- Errors affect the accuracy and precision of data. These can be categorised into:
 - random errors
 - systematic errors.

Accuracy and precision



- Percentage error helps to indicate the accuracy of a measurement.

$$\text{Percentage error (\%)} = \left| \frac{\text{measured value} - \text{true value}}{\text{true value}} \right| \times \frac{100}{1}$$

- Uncertainty describes the variability in the measured results:

$$\text{Absolute uncertainty} = \pm \frac{\text{maximum} - \text{minimum}}{2}$$

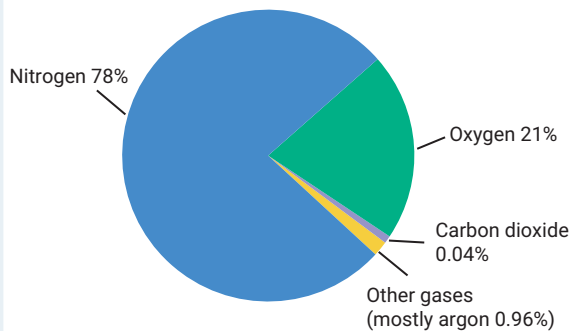
$$\text{Percentage uncertainty (\%)} = \frac{\text{absolute uncertainty}}{\text{measurement}} \times \frac{100}{1}$$

- Pearson's correlation (R) coefficient helps quantify the direction and strength of the relationships between the measured variables.
 - $R = 0$ suggests no correlation.
 - $R = 1$ suggests a strong positive correlation.
 - $R = -1$ suggests a strong negative correlation.

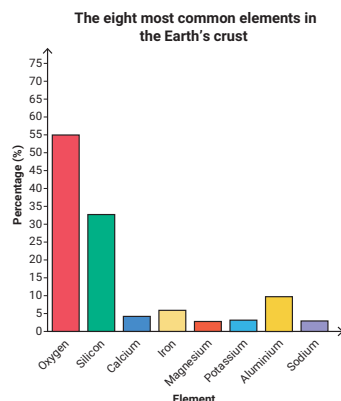
Graphs

- There are many different graphical representations of data, including:
 - linear graphs
 - column graphs
 - pie charts
 - scatterplots.

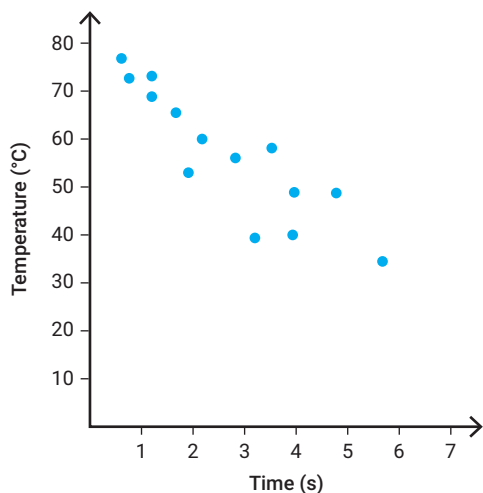
Pie charts



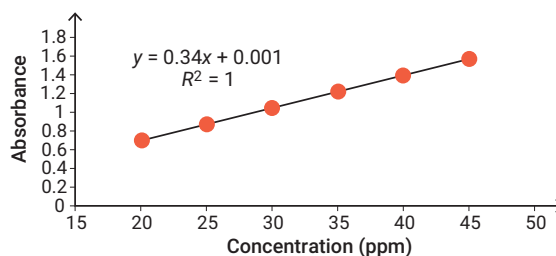
Column graphs



Scatterplots

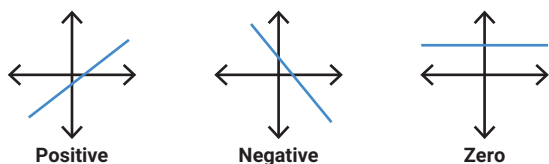


Line graphs

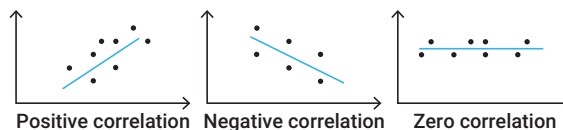


- Graphs help to show the relationship between variables. Although they can show correlation, this does not mean causation.

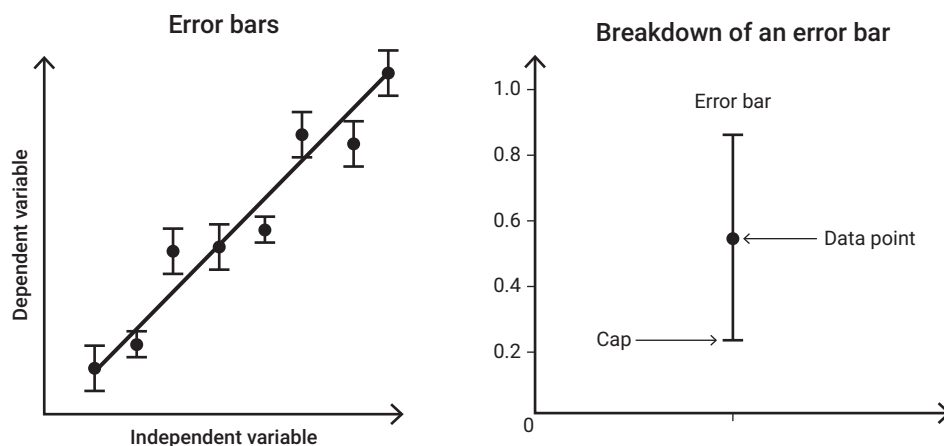
Correlation of line graphs



Correlation trends of scatterplots

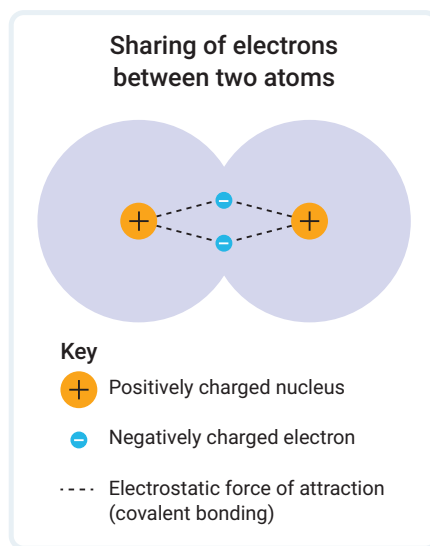


- Error bars on graphs help to visualise variability of measurements around the mean.
 - The central point shows the data point.
 - The upper and lower limits show the variability of the measured values.



Scientific drawings

- Scientific drawings are representations of scientific concepts.



Interpreting, analysing and evaluating evidence

- Apart from speaking to the trends shown in the data, when analysing data it is also important to assess:
 - appropriateness of the method
 - sample size
 - sources of error
 - degree of uncertainty of the data.
- When evaluating evidence, make sure to also address any limitations present in the experiment, including:
 - experimental limitations; for example, time available to conduct the experiment, errors
 - methodological limitations; for example, accuracy and reliability of measurement techniques, ethical constraints

- external limitations; for example, environmental factors that can introduce variability, access to proper equipment.

Communicating findings

- When communicating findings, make sure to:
 - use appropriate conventions and nomenclature
 - use language appropriate to the audience
 - reference appropriately using the relevant referencing system.

MULTIPLE CHOICE

1. Which measure provides information about the spread or variability of data points in a data set?
 - A Mean
 - B Mode
 - C Outlier
 - D Standard error
2. Which of the following data sets would be considered precise but not accurate?
 - A Volume of solution added to a flask using a graduated cylinder with a chipped spout (Readings: 22 mL, 21 mL, 23 mL)
 - B Measurements of the pH of a solution using a faulty indicator that consistently reads 0.5 units higher than the actual pH (Readings: 3.5, 3.5, 3.6)
 - C Mass of a metal bar measured on a balance that is not zeroed properly (Readings: 11.4 g, 15.8 g, 13.2 g)
 - D Temperature readings of a Bunsen burner flame measured with a thermometer that is positioned slightly farther from the flame each time (Readings: 1500°C, 1500°C, 1600°C)

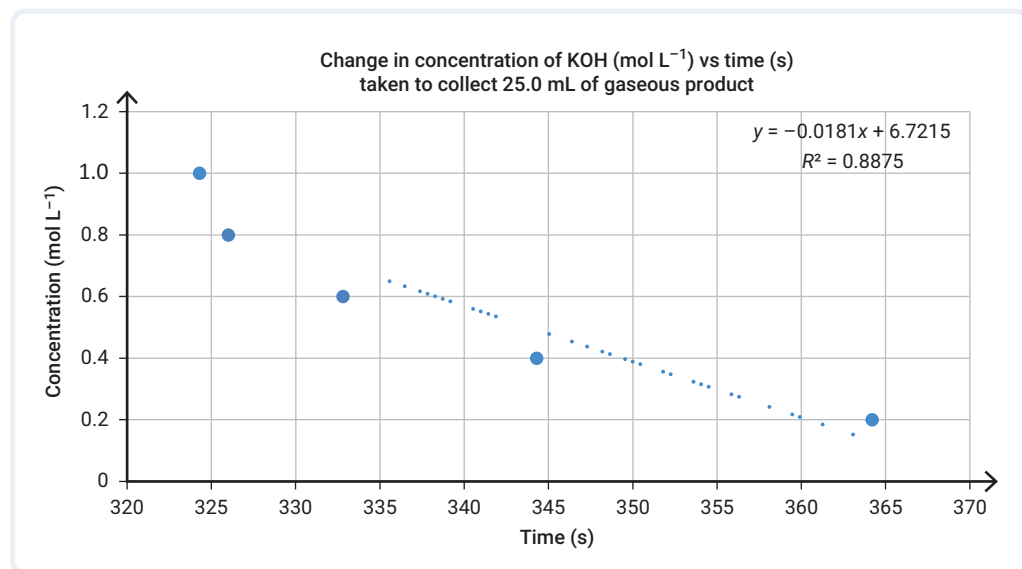
Questions 3 and 4 relate to the following information.

A Chemistry student measures the volume of a liquid using a graduated cylinder. The student reads the volume as 12.5 mL. However, a more precise instrument shows the actual volume to be 12.9 mL.

3. What is the percentage error in the student's measurement?
 - A 1.5%
 - B 2.4%
 - C 2.6%
 - D 3.1%
4. Which of the following actions would not improve the reliability of the student's results in subsequent measurements?
 - A Using a graduated cylinder with a smaller volume division (e.g. 0.1 mL instead of 1 mL)
 - B Repeating the measurement with the same graduated cylinder and averaging the results
 - C Calibrating the graduated cylinder by measuring a known volume of a different liquid
 - D Estimating the volume between the markings on the graduated cylinder
5. In a Chemistry experiment, students investigated the effect of the concentration of solutions on light absorbance. Which of the following can be considered an extraneous variable?
 - A Type of light source used
 - B Concentration of the solution
 - C Length of time the solution was exposed to light
 - D Temperature of the room during the experiment

Questions 6–8 relate to the following information.

A student conducted an experiment to measure the rate of a reaction. The rate was measured through the collection of the gaseous product.



6. The independent variable is:
- A concentration (mol L⁻¹).
 - B gas (mL).
 - C mass (grams).
 - D time (seconds).
7. Which of the following correctly describes the trend shown by the graph?
- A As the concentration increases, the time taken to collect 25.0 mL of gas increases.
 - B As the concentration decreases, the time taken to collect 25.0 mL of gas increases.
 - C As the concentration decreases, the time taken to collect 25.0 mL of gas decreases.
 - D As the concentration increases, the time taken to collect 25.0 mL of gas remains the same.
8. The R^2 shows:
- A a weak correlation between the variables.
 - B a strong correlation between the variables.
 - C a negative correlation between the variables.
 - D no correlation between the variables.
9. Which of the following options can be considered a random error in a chemistry experiment?
- A Using a calibrated balance
 - B Following the experimental procedure
 - C Reading a measuring cylinder at different angles from the meniscus
 - D Performing the experiment at a temperature higher than that outlined in the methodology

10. Which of the following cannot improve the reliability of results in an experiment measuring soil acidity levels in different areas?
- A Increasing the replicates in each area
 - B Using various measurement techniques
 - C Calibrating the pH meters before taking measurements
 - D Reducing the number of data points collected in each area

SHORT RESPONSE

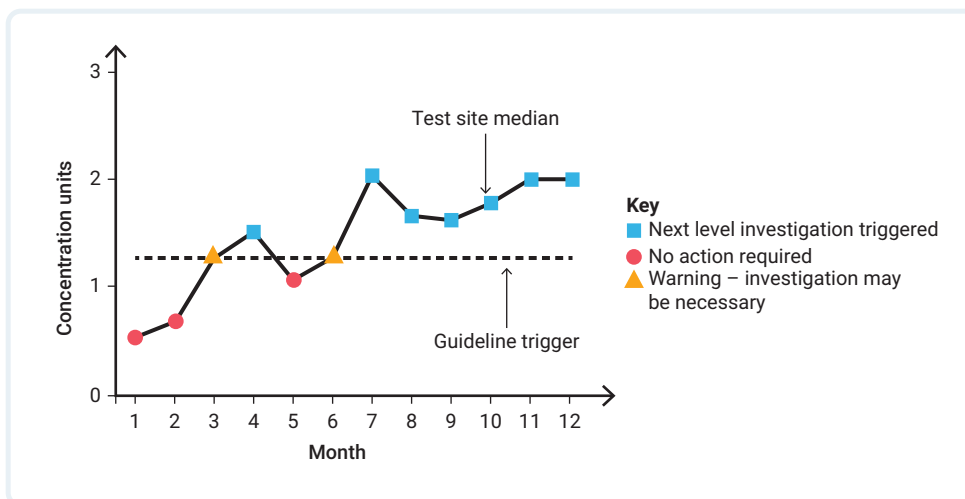
11. The following is a description of an experiment.

Students are investigating the effect of different concentrations of a chemical solution on the rate of a chemical reaction. In the experiment, the reaction is conducted with three different concentrations: 0.1 M, 0.5 M and 1.0 M.

All reactions are conducted at the same temperature and pressure, and the rate of the reaction is measured at regular time intervals.

Write an appropriate research question for this experiment.

12. Governments monitor water quality to ensure that the water is being provided to the population is safe and within guidelines. As part of ongoing monitoring, water samples are measured for the presence of different chemicals.



- a **Identify** the dependent variable.
- b At what point would the guideline be triggered?
- c **Describe** the trend shown by the graph.

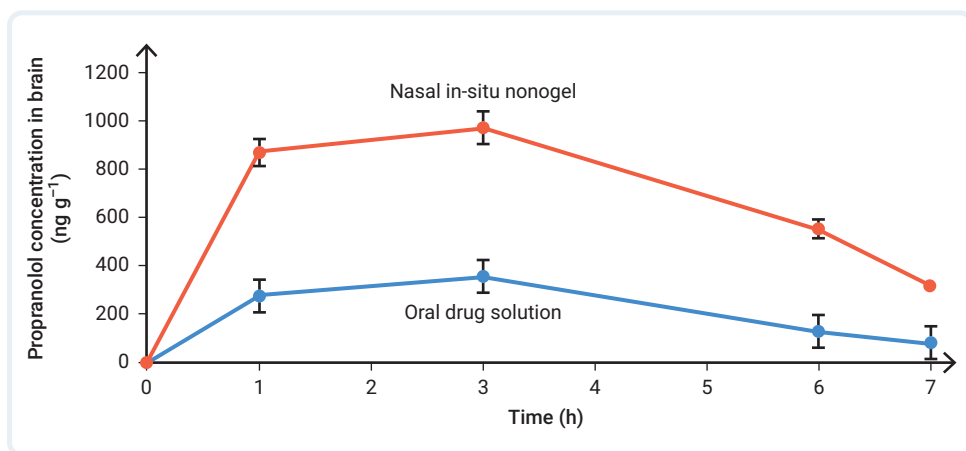
13. The data in the table was collected from an experiment testing the effect of different concentrations of a chemical solution on the rate of a chemical reaction.

Concentration of solution (%)	Time taken for a set mass of product to be produced (s)		
	Trial 1	Trial 2	Trial 3
2.0	$15 \pm 5\%$	$16.8 \pm 5\%$	$16.2 \pm 5\%$
1.5	$24 \pm 5\%$	$25.8 \pm 5\%$	$25.2 \pm 5\%$
1.0	$34.8 \pm 5\%$	$36 \pm 5\%$	$35.4 \pm 5\%$
0.5	$45 \pm 5\%$	$46.8 \pm 5\%$	$46.2 \pm 5\%$

Sketch an appropriate graph to represent this data. Include error bars in your graph.

14. According to the Global Burden of Disease, neurological diseases are one of the main causes of death globally. Migraine is a neurological disease affecting the central nervous system that causes severe pain to one side of the head. Propranolol is a drug developed to help manage migraine attacks; however, multiple factors affect how much of the drug can be absorbed into and used by the body.

The following graph was taken from a study exploring the best mechanisms to deliver the drug to improve the availability of propranolol.



- Describe** the trend seen in the graph.
- For someone suffering from migraines, which delivery method would be the preferred method? **Explain** your choice.

ANSWERS

CHAPTER DC SCIENCE RESEARCH

LEARNING CHECK DC.1

DESCRIBING

- Accuracy refers to how close a value is to the true value, whereas precision describes how close a set of measurements are to each other.
 - Reliability refers to the consistency of measurements upon repeat experiments, whereas validity describes the extent to which the experiment measures what it is intended to measure.
- Calibrate measuring instruments.
Record multiple measurements and take an average.
- Data that is in categories
 - Continuous data or measurements taken over time
 - Data that is in categories or discrete data
- A logbook contains detailed notes relating to the experiment, including observations, results and methodology. This information is important during the analysis of the results of the experiment.
 - MSDSs contain information about the safe handling of any chemicals or substances used in an experiment. This ensures the safety of the researchers.
 - Ethics help to ensure the safety of participants in the experiment, the integrity of the experiment and the proper use of the experimental results.
- Research question, materials, methods, discussion

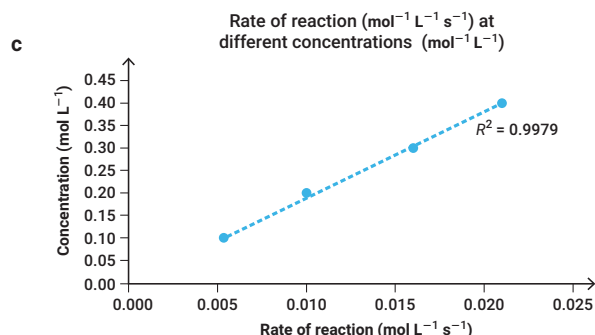
APPLYING

- Dependent variable: rate of reaction (mL s^{-1})
 - Independent variable: temperature
- Potential research question: Does the consumption of food (grams) before running affect the total distance a person can run (km)?
- Potential risk: The student could drop the measuring cylinder, causing it to break and shatter. The pieces of glass could cause potential damage to the skin.
Risk minimisation – Wear gloves and closed-toe shoes to protect against pieces of glass if broken.
- The experimenter breached informed consent. The experiment ignored the participant's refusal to be tested and performed the procedure anyway. Participants need to give formal consent to participate in experiments and the experimenter must respect their wishes, not go against them.

- 10 a 2 b 6 c 5 d 1

- Percentage uncertainty = $\frac{1}{20} \times 100 = 5\%$
 - This suggests that the measurements are relatively imprecise.
- The rate of cooling of each substance – this could affect the temperature recorded when the substance melts
 - Random error
- Absolute uncertainty = $\frac{14.2 - 14.1}{2} = \pm 0.1 \text{ cm}$
- Dependent variable : Rate of reaction
Independent variable: Concentration of reactant

Concentration (mol L^{-1})	Mean rate ($\text{mol L}^{-1} \text{s}^{-1}$)
0.1	0.0053
0.2	0.0073
0.3	0.016
0.4	0.021



- There is a positive correlation between the rate of reaction and concentration of reactant.

ANALYSING

- Concentration of phosphorus
 - Sample 3
- Scatterplot
 - Line of best fit
 - Gradient = $\frac{\Delta y}{\Delta x} = \frac{4 - 2}{0.480 - 0.250} = 8.70$ (approximately)
 - Approximately 1.500
 - Since there is a strong positive correlation between the two variables, the R value is likely to be closer to 1 than 0.

LEARNING CHECK DC.2

DESCRIBING

- 1 Primary data is firsthand data collected by the researcher, whereas secondary data refers to information collected from an experiment conducted by someone else.
- 2 The peer-review process helps to ensure the quality of the research and the results. It also helps to add credibility to the research investigation.

APPLYING

- 3 Experimental limitation: the accuracy of the pH meter
Methodological limitation: only testing three concentration levels
- 4 How fast a reaction occurs is the rate of reaction. Knowing how fast or slow reactions occur can help scientists with studying chemical reactions.
- 5 (Cruzan, 2012)

CHAPTER EXAM

MULTIPLE CHOICE

1 D	3 D	5 A	7 B	9 C
2 B	4 D	6 D	8 B	10 D

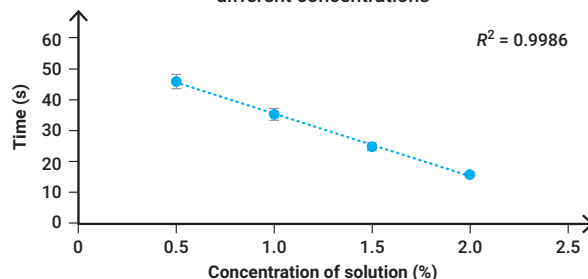
SHORT RESPONSE

- 11 How does the concentration(M) of a chemical solution affect the rate of a chemical reaction(mL s^{-1})?
- 12 **a** Concentration
b At a concentration above 1.2–1.3

- c** In the first 4 months, the concentration increased over time. From months 3–4, the concentration increased high enough for the guideline to be triggered. From months 4–5, the concentration decreased but began to increase again from months 5–7. From months 7–9, the concentration decreased over time but began to increase again from months 9–11 after which it remained stable. For the period of months 7–12, the concentration was greater than the guideline trigger point.

13

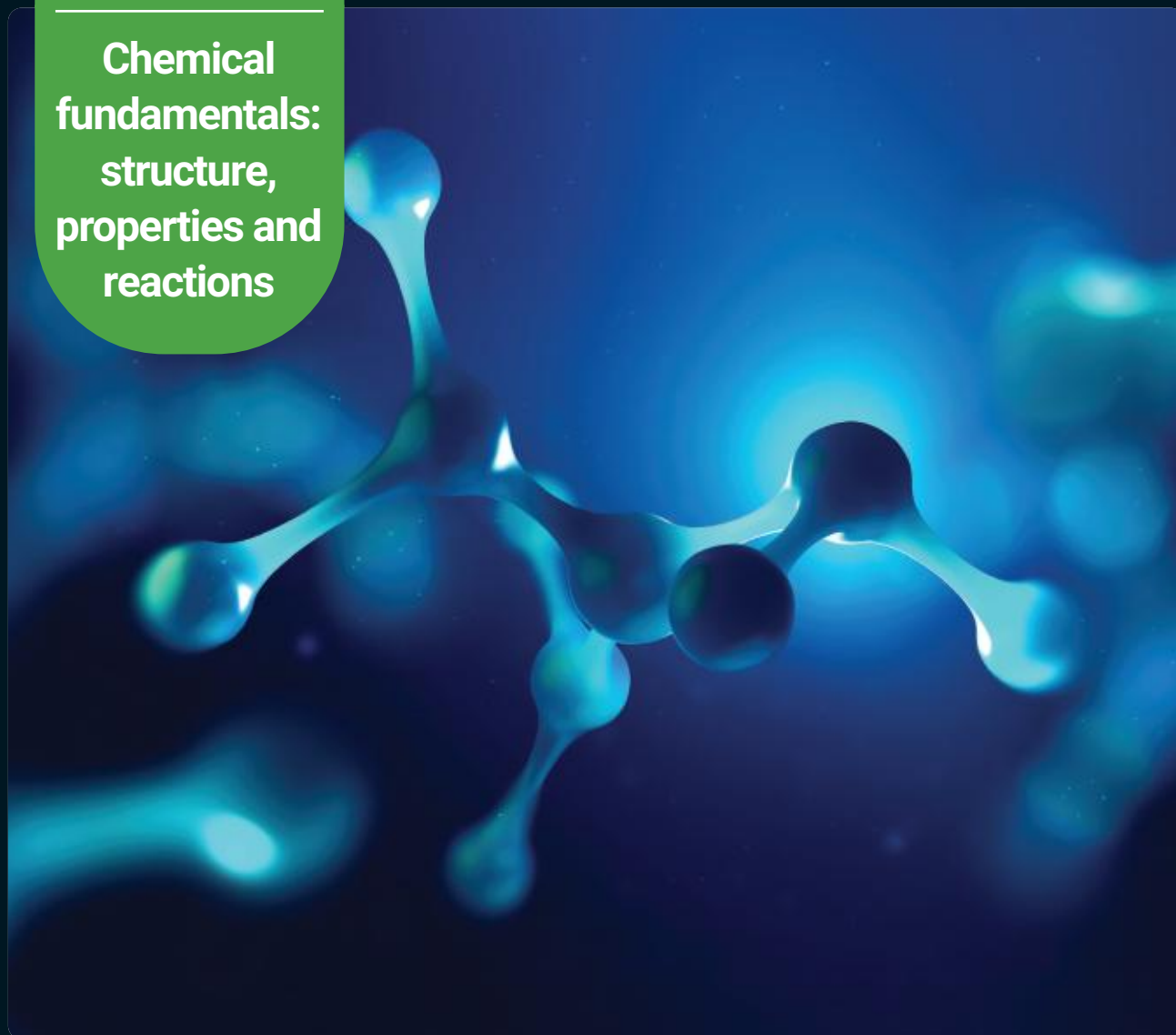
Time taken for product to be produced at different concentrations



- a** The graph shows that nasal in-situ nonogel allows for a higher concentration of propranolol in the brain over time compared to oral drug solution. For both graphs, the concentration of propranolol in the brain increases during the first 3 hours, after which the concentration decreases over time.
- b** For someone suffering from migraines who would want more of the migraine drug to be available quicker, nasal in-situ nonogel would be a preferred option because it allows for both a quicker and a higher concentration of medication to reach the brain than the oral drug solution. It also allows for more of the drug to be available in the brain over time.

UNIT 1

Chemical fundamentals: structure, properties and reactions



Yurchanka Siarhei/Shutterstock

Topic 1: Properties and structure of atoms

CHAPTERS RELATED TO THIS TOPIC AREA: 1–5

Topic 2: Properties and structure of materials

CHAPTERS RELATED TO THIS TOPIC AREA: 6 and 7

Topic 3: Chemical reactions – reactants, products and energy change

CHAPTERS RELATED TO THIS TOPIC AREA: 8–10

Chemical reactions occur when bonds are broken between atoms and molecules within the reactants and new bonds are formed to produce the products. How this occurs depends on the structure of the reactants and products and the bonding that exists between the atoms within them. Bonding depends on the arrangement of subatomic particles within atoms. The periodic table is a powerful tool that helps us to understand the patterns and trends observed in the properties of atoms of different elements and to explain them in terms of atomic structure. Concepts such as 'the mole' enable chemists to measure the quantities of particles involved in a chemical reaction accurately, and to predict the amounts of products that will form.

UNIT OBJECTIVES

By the end of this unit, students should be able to:

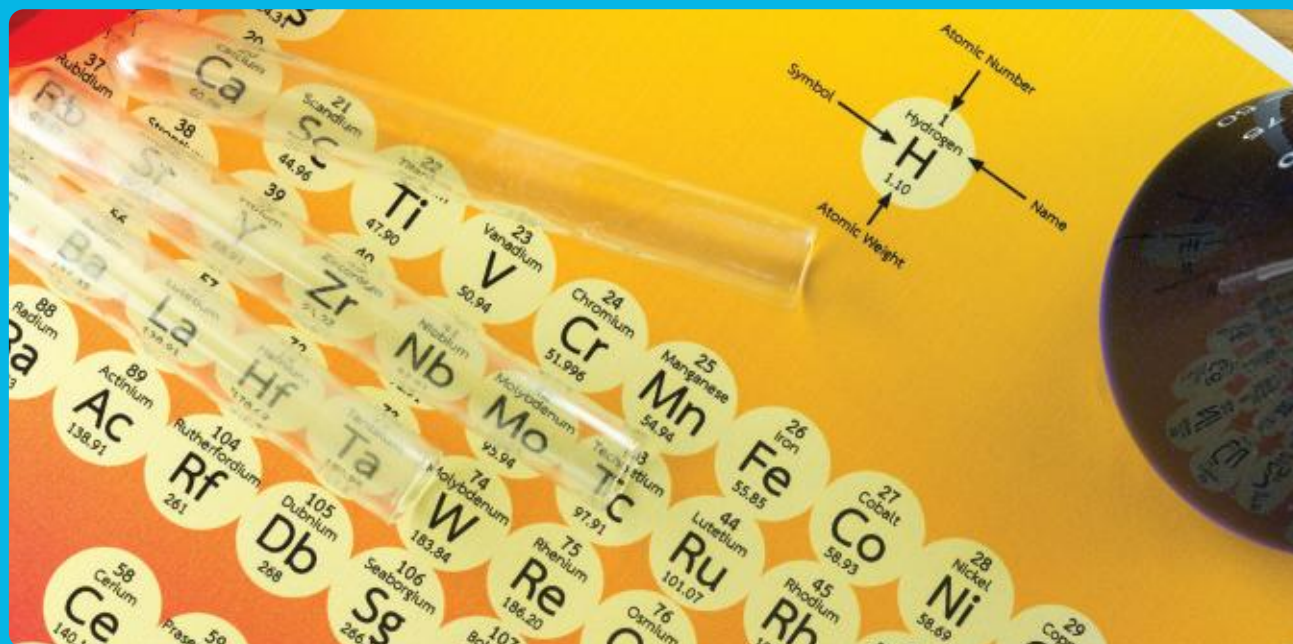
1. describe ideas and findings about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change 2. apply understanding of properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change 3. analyse data about properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy 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 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

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

CHAPTER

1

Atomic structure



Shutterstock.com/Dabarti CGI

SYLLABUS DOT POINTS

SCIENCE UNDERSTANDING

- State that elements are represented by symbols.
- 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*.
- Describe the relationship between the structure of the periodic table and the electronic configuration of atoms.
- Apply the nuclear symbol notation ${}^A_ZM$ 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 in energy levels.
- Identify that the periodic table is arranged into four blocks associated with the four sub-levels: s, p, d and f.
- Determine full and condensed electron configurations for atoms and ions up to $Z = 36$ (e.g. $1s^2 2s^2 2p^6 3s^2 3p^5$ and $[\text{Ne}]3s^2 3p^5$).



- Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to $Z = 36$.
- Identify the electron configuration of Cr and Cu as exceptions.

SCIENCE AS A HUMAN ENDEAVOUR

- **Appreciate** that experiments provided evidence that enabled scientists to develop models of the atom.
- **Consider** the role the Geiger–Marsden gold foil experiments played in the development of atomic theory.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

The current understanding of the atom and its structure has been developed over a long time as a result of a number of important experiments. Knowledge of atomic structure allows scientists to make predictions about the properties and reactions of different elements and compounds.

Practical

- Gold foil experiment

Worksheets

- The atom
- Electron configurations



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ All matter is made up of atoms.
- ✓ An atom is made up of smaller particles.
- ✓ Elements are composed of only one type of atom.
- ✓ Each element is represented by a unique symbol.
- ✓ The periodic table lists all the known elements.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ describe how protons, neutrons and electrons are arranged in an atom
- ✓ distinguish between atomic number and mass number
- ✓ apply appropriate notation to determine number of protons, neutrons and electrons in an atom
- ✓ distinguish between atoms and ions in terms of numbers of protons and electrons
- ✓ state the definition of an element and how elements are represented
- ✓ identify how the organisation of the periodic table is related to atomic number and atomic subshells
- ✓ relate the structure of the periodic table to the electronic configuration of atoms
- ✓ compare the relative energies and number of electrons found in s, p and d orbitals and subshells
- ✓ recall the Aufbau principle, Hund's rule and the Pauli exclusion principle
- ✓ apply appropriate rules and principles to write full and condensed electron configurations
- ✓ identify exceptions to general electron configuration rules.

atom the smallest stable unit of matter that can exist by itself

electron a subatomic particle with a negative charge that is present in all atoms

alpha particle a nucleus of a helium atom that is emitted by some radioactive substances as alpha radiation

1.1 The atom

In 1808, after examining information accumulated by many chemists and along with data from his own experiments, English chemist John Dalton, concluded that the early Greek philosophers had been right. All matter was made of separate tiny particles, **atoms**.

In 1897, J.J. Thomson, while experimenting with cathode ray tubes, discovered the electron. He showed that the stream of glowing material produced when an electric current was passed through a vacuum tube could be bent by electric and magnetic fields. He concluded the rays consisted of charged particles. He realised that, within the atom, there were negatively charged subatomic particles, which are now recognised as **electrons**. In 1904 he proposed that the atom was made of a positively charged material, throughout which electrons were distributed, like currants in a fruit cake or a plum pudding (**Figure 1.1.1**). Thomson's model of the atom is often referred to as the 'plum pudding model'.

Thomson's model of the atom was not universally accepted by scientists because Thomson could not develop a stable model of his concept. Between 1908 and 1913, a series of experiments, performed by Hans Geiger and Ernest Marsden under the direction of Ernest Rutherford, led to a different model of atomic structure. Geiger and Marsden fired a beam of **alpha particles** at thin foils of various metals, including gold, and analysed the resultant scattering patterns of the alpha particles. If Thomson's model was correct, the alpha particles should have passed straight through the metal atoms with little or no deflection. However, Geiger, Marsden and Rutherford found that the particles were scattered in all directions, with some particles having significant deflection.

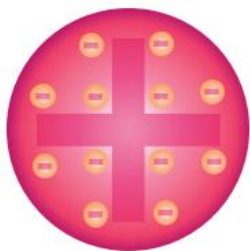
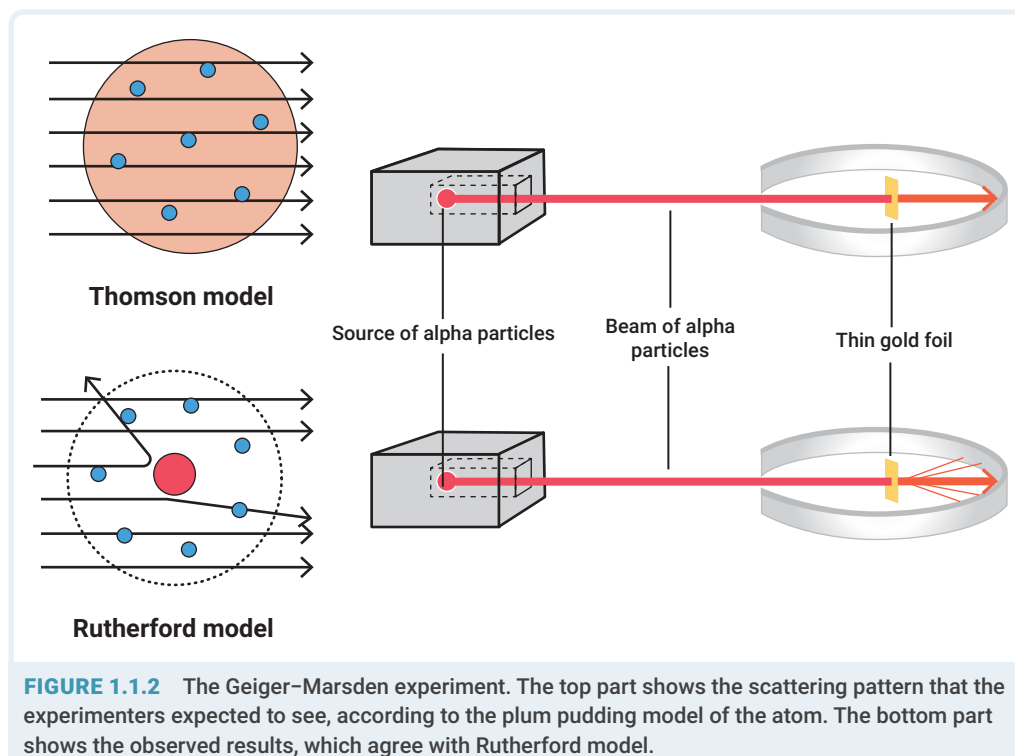


FIGURE 1.1.1 J.J. Thomson's plum pudding model of the atom

These unexpected results led Rutherford, in 1911, to publish a landmark paper in which he proposed that most of the atom is empty space but there is a small, dense, charged structure at its centre. Data from a series of further experiments by Geiger and Marsden, in 1913, led Rutherford to conclude the centre or **nucleus** (as he now called it) was positively charged. A comparison of Thomson and Rutherford's models is shown in **Figure 1.1.2**.

nucleus the positively charged centre of an atom, containing positively charged protons and electrically neutral neutrons, to which the orbiting electrons are attracted



Weblink
Rutherford's gold foil experiment

PRACTICAL ACTIVITY 1.1.1

GOLD FOIL EXPERIMENT

Introduction

The Geiger-Marsden gold foil experiments helped develop our understanding of atomic theory. Thanks to this, we have been able to make advancements in many other fields of science.

Research question

How does the pattern of movement of particles help identify the structure of an atom?

Aim

To simulate the Geiger-Marsden gold foil experiment

Materials

- 1 hula hoop
- 1 golf ball
- 1 ping-pong ball
- string

Procedure

- 1 Work in pairs.
- 2 Tie the golf ball onto the hula hoop so it hangs approximately in the middle of the hula hoop when it is held vertically.
- 3 With their eyes closed, one student lightly tosses a ping-pong ball 30 times randomly through the hula hoop.

- 4 The other student records how many times the ping-pong ball hits the golf ball.
- 5 Swap roles and repeat steps 3 and 4.

Results

Collate a class set of results.

Analysis of results

- 1 Explain what the hula hoop, the golf ball and the ping-pong ball each represent, in relation to the experiment.
- 2 Using the class set of results, compare the number of times the golf ball was hit by the ping-pong ball with the number of times it was missed. Explain whether this aligns with the results of the gold foil experiment.

Interpretation

- 3 Explain whether the results of this simulation would lead to the same conclusion as Rutherford's.

Evaluation

- 4 Suggest any changes that could be made to improve the simulation.

proton a positively charged particle that exists in the nuclei of all atoms

neutron an uncharged particle that exists in the nuclei of all atoms except some hydrogen atoms



Weblinks

Atoms: the space between

Protons, neutrons
and electrons

On the basis of Rutherford's work and the contributions of many other scientists, such as Henry Moseley and James Chadwick, scientists now consider that the nucleus contains subatomic particles called **protons** and **neutrons**. The nucleus contains most of the mass of the atom but occupies only a small volume of the atom. Electrons surround the nucleus. Electrons are very small, in constant and extremely rapid motion, and are spread out over a relatively large distance. This produces an electron cloud around the nucleus (**Figure 1.1.3**). Despite the large amount of space they cover, electrons contribute almost no mass to the atom. If the atom were the size of the Brisbane Cricket Ground, including the grandstands, then each electron would be the size of a pea somewhere in the whole ground, and the nucleus would be no bigger than a peach in the centre of the playing field.

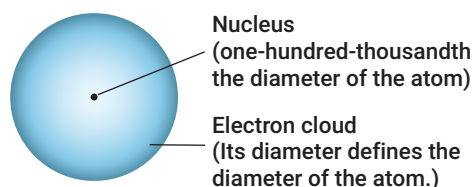


FIGURE 1.1.3 The structure of an atom, showing a nucleus and an electron cloud

Subatomic particles

An atom contains three types of subatomic particle – the proton, the neutron and the electron. The properties of each particle are shown in Table 1.1.1.

TABLE 1.1.1 Symbols, charges and comparative masses of subatomic particles

Particle	Symbol	Charge	Comparative mass
Proton	p	1+	1
Neutron	n	0	1
Electron	e	1–	$\frac{1}{1800}$ (approx.)

What holds an atom together?

In an atom, the nucleus has an overall positive charge, because it consists of positively charged protons and uncharged neutrons. The electrons around the nucleus are negatively charged. The attraction between the positive nucleus and the negative electrons keeps the atom together. This is called **electrostatic attraction**.

An atom can be modelled as a nucleus surrounded by electrons, held together by electrostatic forces of attraction between the nucleus and the electrons.

electrostatic attraction
a force that occurs between particles with opposite charges

LEARNING CHECK 1.1

DESCRIBING

- 1 **Name** the three particles in an atom and **state** the charge of each particle.
- 2 **Identify** where each of the three particles is located in the atom.

APPLYING

- 3 **Explain** the balance of forces that exists in an atom that makes it stable.
- 4 Rutherford deduced that the nucleus of an atom is very small, extremely dense and positively charged. **Explain** how the Geiger–Marsden experiments led him to draw these conclusions.
- 5 Japanese scientist Hantaro Nagaoka also rejected Thomson’s model. In 1904, he proposed instead that electrons orbit the positive charge like the rings around Saturn. **Compare** Nagaoka’s and Rutherford’s models.

1.2 Atoms, elements and the periodic table

Elements

In his 1665 book *The Sceptical Chymist*, Robert Boyle proposed that an **element** was a substance that could not be reduced to something simpler. It wasn’t until 1808 that a link was made between elements and atoms when John Dalton asserted that the specific properties of elements were based on the atoms of which they were composed. Dalton was also responsible for recommending that symbols be used to represent different atoms. The definition, of an element as a substance consisting of only a single type of atom, is central to chemistry.

By 1900, almost all naturally occurring elements had become known. Today, there are 118 known elements, from which all other substances and materials are derived. An element is defined by the number of protons in the nucleus of the atoms of which it is composed. The number of protons is called the **atomic number** and this number is unique to each element. For example, a hydrogen atom has one proton, so hydrogen has the atomic number 1. Oxygen has the atomic number 8, which means each atom of oxygen has eight protons.

Elements can be represented by symbols. Some elements are represented by the first letter of their names; for example, hydrogen is ‘H’ and carbon is ‘C’. Some other elements have two letters; for example, aluminium is ‘Al’ and magnesium is ‘Mg’. The first letter of an element’s symbol is always a capital letter. The second letter is always lower case. Some elements are represented by letters from their names originating in other languages such as Latin (e.g. gold (Au) from *aurum*, meaning ‘gold’; potassium (K) from *kalium*, meaning ‘potash’) and Greek (e.g. chlorine (Cl) from *chlorós*, meaning ‘pale green’).

element a substance consisting of only one type of atom

atomic number (Z) the number of protons in one atom of an element – each element has a unique atomic number

IUPAC the International Union of Pure and Applied Chemistry

periodic table of elements
a table listing the known elements in order of atomic number, based on their chemical and structural properties, so that patterns and trends are easily recognisable

period a horizontal row on the periodic table

group a vertical column on the periodic table

Element names are determined by the International Union of Pure and Applied Chemistry (**IUPAC**). Many of the elements have been known for a long time, but some elements have only recently been synthesised in laboratories. IUPAC approves names suggested by scientists, companies or universities that were instrumental in the discovery of the element. For example, element 116 was named livermorium (Lv) in 2012 after the Lawrence Livermore National Laboratory in California.

The periodic table of elements

The **periodic table of elements**, commonly known as the ‘periodic table’, is the chemist’s method of arranging all the elements in a systematic way so that trends and patterns relating to their individual properties can be seen. The periodic table was developed by the Russian chemist Dmitri Mendeleev and first published in 1869.

The periodic table is arranged in a series of horizontal rows called **periods** and vertical columns called **groups**. The elements are listed in the periodic table in order of atomic number, from element 1, hydrogen (H), to element 118, oganesson (Og). A simple periodic table showing groups and periods is shown in **Figure 1.2.1**. Notice that the periods and groups are numbered. This makes it easier to locate an element when referring to it on the periodic table.

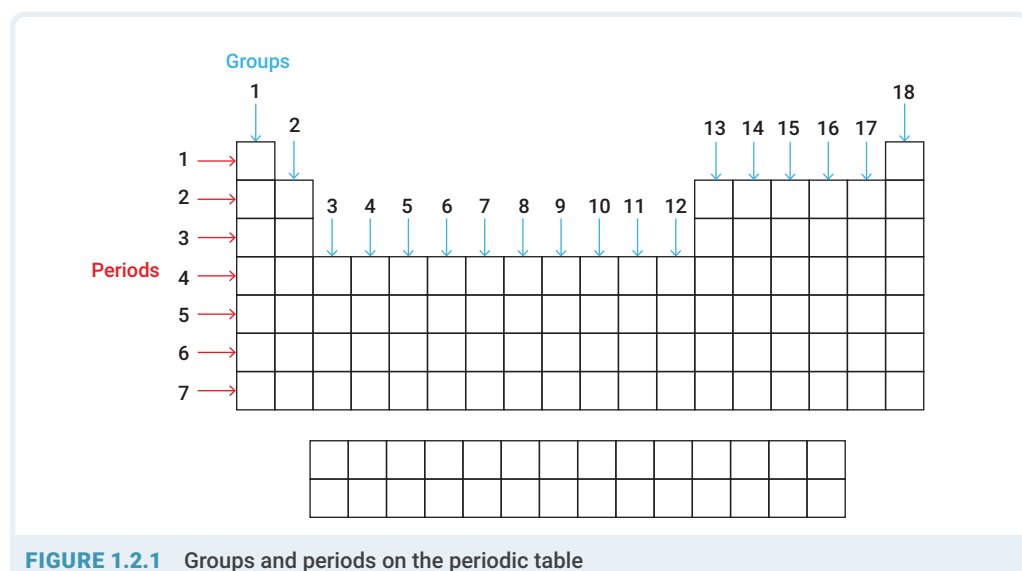


FIGURE 1.2.1 Groups and periods on the periodic table

Protons, neutrons and electrons

The composition of elements can be represented using the nuclear symbol notation shown in **Figure 1.2.2** to show the number of protons, neutrons and electrons in an atom.

The atomic number (Z) represents the number of protons in an atom and is the number that defines an atom. For example, an atom of carbon will always have six protons, even if the

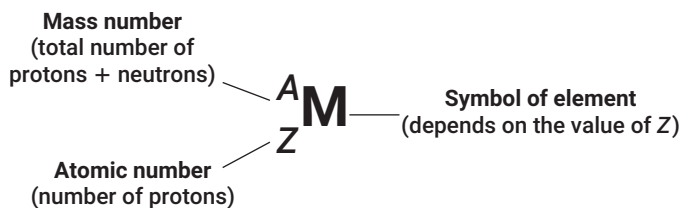
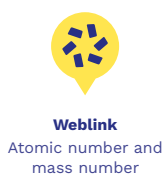


FIGURE 1.2.2 Atoms can be represented by a symbol that includes the atomic number and mass number.

number of neutrons or electrons changes. A neutral (uncharged) atom has equal numbers of electrons and protons, so the atomic number also gives the number of electrons in the atom.

The **mass number (A)** is the total number of protons and neutrons in the atom. To find the number of neutrons in an atom, subtract the atomic number, Z , from the mass number, A .

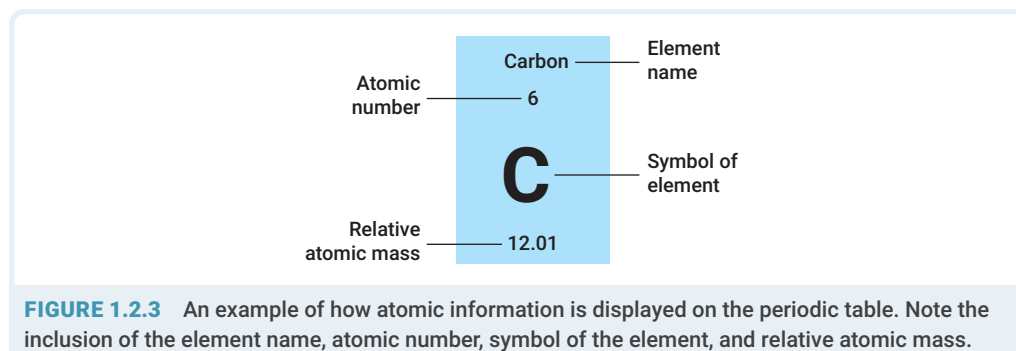
$$\text{Number of neutrons} = \text{mass number } (A) - \text{atomic number } (Z)$$

Often, a periodic table displays the **relative atomic mass (A_r)** of an element, rather than the mass number. You will learn more about this in Chapter 2.

The information about mass and atomic numbers found on the periodic table is shown in **Figure 1.2.3**.

mass number (A) the number of protons and neutrons in the nucleus of an atom

relative atomic mass (A_r) the average mass of one atom of an element, relative to the mass of a carbon-12 atom



Worksheet
The atom

WORKED EXAMPLE 1.2.1

Sodium is represented by $^{23}_{11}\text{Na}$. Determine the number of protons, neutrons and electrons in sodium.

ANSWER

1 Deconstruct the meaning of the representation.

Remember that in the representation ^A_ZM , A represents the mass number, Z represents the atomic number, M represents the chemical symbol of the element.

2 Recall the definition of each key term.

- Mass number: the number of protons and neutrons
- Atomic number: the number of protons

3 Determine the number of protons, neutrons and electrons.

- The atomic number of sodium is 11. This means that it has 11 protons.
- The mass number of sodium is 23, so it has $23 - 11 = 12$ neutrons.
- The atom is neutral, so there is an equal number of electrons and protons.

Therefore, sodium has 11 protons, 12 neutrons and 11 electrons.

Sometimes, when atoms undergo chemical reactions, an atom loses or gains electrons. A loss of one or more electrons means the atom has more protons than electrons, resulting in a positively charged particle. A gain of one or more electrons means the atom has more electrons than protons, resulting in a negatively charged particle. These charged particles are called **ions**. A positive or negative ion can be identified by comparing the number of electrons with the atomic number.

Ions are identified by a superscript $+$ or $-$ to the right of the element symbol. If more than one electron has been lost or gained, a number is added to show how many. For example, H^+ represents a positive hydrogen ion formed when a hydrogen atom lost one electron, whereas P^{3-} represents a negative phosphorus ion formed when a phosphorus atom gained three electrons. In nuclear symbol notation, a beryllium ion, formed by the loss of two electrons is represented by $^8_4\text{Be}^{2+}$.

ion an atom that has become charged through the gain or loss of one or more electrons



Syllabus link
Chapter 5 will cover ions in more detail.

WORKED EXAMPLE 1.2.2

An atom of oxygen is identified as having 10 electrons.

- a Identify whether this atom is an ion.
- b If it is an ion, state the overall charge on the ion.

ANSWERS

- a **1 Identify the atomic number of oxygen.**

The atomic number of oxygen is 8. This means that it has eight protons.

- 2 Use the atomic number to determine the number of electrons present in a neutral (uncharged) atom.**

In a neutral atom, the number of electrons is equal to the number of protons.

- 3 Identify whether the atom is an ion.**

The atom is stated to have 10 electrons so there are two more electrons than protons. Because the atom has an overall charge, it must be an ion.

- b **Determine the discrepancy between the number of protons and electrons.**

The atom has 10 negatively charged electrons and eight positively charged protons.

Therefore, the atom must be negatively charged overall. Because there are two more electrons than protons, the overall charge is 2⁻.

LEARNING CHECK 1.2

DESCRIBING

- 1 **Define** 'element'.
- 2 **State** why an element is represented by a unique symbol.
- 3 **Distinguish** between a group and a period on the periodic table.
- 4 **Identify** the difference between a neutral atom and an ion.

APPLYING

- 5 **Discriminate** between atomic number and mass number.
- 6 **Outline** how to determine the numbers of protons, neutrons and electrons in an atom.
- 7 **Identify** each element below, and calculate the number of protons, neutrons and electrons in each.



- 8 Represent each of the following as ${}^A_Z\text{M}$, using the periodic table as a reference.

- a An atom of aluminium with a mass number of 27
- b An atom with four protons and a mass number of 9

- 9 Copy and complete the following table. You will need to extract information from the periodic table for some elements.

Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
Hydrogen	1	1			
Magnesium			12	12	
Boron	5			6	
Chlorine		35			17
Nickel		59		31	

- 10 An atom of potassium is identified as having 18 electrons. **Determine** the overall charge on this atom.

1.3 The electron configuration of atoms

Identifying the locations of electrons in the electron cloud surrounding the nucleus of an atom was more challenging for scientists than identifying the locations of protons and neutrons in the nucleus. Understanding the arrangement of electrons in atoms is of great chemical significance because it allows chemists to make predictions about chemical reactions and the formation of compounds.

In 1913, Niels Bohr was the first person to develop a theory to explain the arrangement of electrons in atoms. Like all other aspects of the atom model, his theory was modified and further developed by other scientists, particularly Erwin Schrödinger.

Today, our most developed model for the behaviour of electrons is called the ‘quantum theory’. Within this theory, there is an ‘energy level model’ for the electrons within an atom and it establishes the following principles:

- Electrons are found in specific **electron energy levels** (sometimes referred to as ‘energy shells’) around the nucleus of an atom. Electrons in the same energy level have the same amount of energy.
- The level with lowest energy is closest to the nucleus. The further away from the nucleus the energy level, the greater the energy it has (**Figure 1.3.1**).
- Electrons in an atom cannot exist between energy levels, although they can move from one energy level to another as energy is absorbed or released.
- The difference in energy between successive energy levels decreases the further away the levels are from the nucleus.
- Each energy level within an atom can hold only a specific maximum number of electrons. The general formula $2n^2$ can be used to determine the maximum number of electrons an energy level can hold, where n is the energy level number. This rule applies to the first four energy levels. For example, energy level $n = 3$ can hold $2 \times 3^2 = 18$ electrons. **Table 1.3.1** shows the maximum number of electrons that can be found in each energy level.

electron energy levels a region of an atom where electrons have the same energy

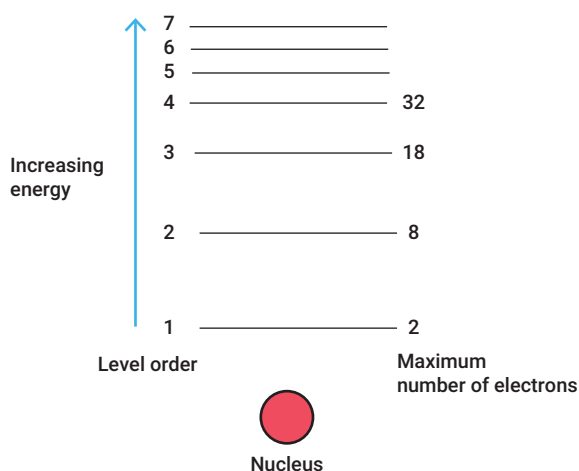


FIGURE 1.3.1 Electron energy levels in an atom

electron configuration
the specific arrangement
of electrons within an
individual atom

Aufbau principle
electrons around an
atom will fill the lowest
available energy level
first, before filling higher
energy levels



Weblink

Electron configuration

TABLE 1.3.1 The maximum number of electrons in each energy level

Energy level (n)	Calculation of electron number ($2n^2$)	Number of electrons
1	2×1^2	2
2	2×2^2	8
3	2×3^2	18
4	2×4^2	32

Determining electron configurations

The **electron configuration** of each atom is unique. To determine the electron configuration of an atom, a series of rules should be followed. The **Aufbau principle** states that electrons around a nucleus fill up energy levels in order from the lowest to the highest. This means that the single electron in hydrogen goes into the lowest energy level ($n = 1$) and the configuration is written as 1. The two electrons in helium also go into the lowest energy level, because it can hold two electrons. The electron configuration is written as 2.

Lithium has three electrons. The first two of these exist in the lowest energy level ($n = 1$), which is now full. The third electron moves to the next highest energy level ($n = 2$). The configuration of the lithium atom is written as 2,1.

Configurations are written to show how many electrons are in each level, and the order of the levels. The configurations for the first 10 elements are shown in **Table 1.3.2**. Notice that for neon, the second energy level is filled with the maximum of eight electrons it can hold.

TABLE 1.3.2 The electron configurations of the first 10 elements

Element	Number of electrons	Electron configuration
Hydrogen	1	1
Helium	2	2
Lithium	3	2,1
Beryllium	4	2,2
Boron	5	2,3
Carbon	6	2,4
Nitrogen	7	2,5
Oxygen	8	2,6
Fluorine	9	2,7
Neon	10	2,8

With the first two energy levels filled, the electrons of the next eight elements (sodium to argon) fill up the third energy level. These configurations are shown in **Table 1.3.3**.

TABLE 1.3.3 The electron configurations of the elements of period 3

Element	Number of electrons	Electron configuration
Sodium	11	2,8,1
Magnesium	12	2,8,2
Aluminium	13	2,8,3
Silicon	14	2,8,4
Phosphorus	15	2,8,5
Sulfur	16	2,8,6
Chlorine	17	2,8,7
Argon	18	2,8,8

Atomic orbitals and subshells

Look back at Figure 1.2.1, groups and periods of the periodic table (page 10). Do you notice a pattern emerging? The electrons of elements in period 1 are filling energy level 1, period 2 are filling energy level 2 and period 3 are filling energy level 3.

At group 18 for periods 1 and 2, the level is also full. This is not consistent for period 3 (level 3) because the maximum number of electrons in level 3 is 18. Where are the other 10 electrons going?

The reason the pattern becomes more complicated is that each major energy level contains one or more **subshells**, which are designated with the letters 's', 'p', 'd' and 'f'. Electrons within a subshell have the same amount of energy and move within regions of physical space, called **atomic orbitals**. The **Pauli exclusion principle** states that each atomic orbital can contain a maximum of two electrons; if an orbital contains two electrons, they spin in opposite directions. Each subshell type consists of a specific number of orbitals, as shown in **Table 1.3.4**.

subshell an energy sublevel made up of multiple orbitals that contain electrons with the same amount of energy

atomic orbital a region of three-dimensional space around an atom's nucleus containing a maximum of two electrons

Pauli exclusion principle no two electrons can occupy the same quantum state; hence, two electrons in the same orbital will have opposing spins

TABLE 1.3.4 Subshells, orbitals and electrons

Subshell type	Number of orbitals in the subshell	Number of electrons in the subshell (2 × number of orbitals)
s	1	2
p	3	6
d	5	10
f	7	14

The first energy level has only one subshell, the s subshell, and it is given the designation 1s. The second energy level has two subshells, s and p, given the designation 2s and 2p. In the third energy level there are 3s, 3p and 3d subshells. The fourth energy level has 4s, 4p, 4d and 4f. As you can see, the number in front represents the main energy level, and the letter represents the particular subshell in that energy level.

Table 1.3.5 shows the numbers of subshells, orbitals and electrons in each level, confirming the maximum number of electrons that each level can hold.

TABLE 1.3.5 The orbitals and electrons within the first four energy levels

Energy level number (n)	Subshell	Number of orbitals	Number of electrons	Maximum number of electrons in energy level ($2n^2$)
1	1s	1	2	2
2	2s	1	2	8
	2p	3	6	
3	3s	1	2	18
	3p	3	6	
	3d	5	10	
4	4s	1	2	32
	4p	3	6	
	4d	5	10	
	4f	7	14	

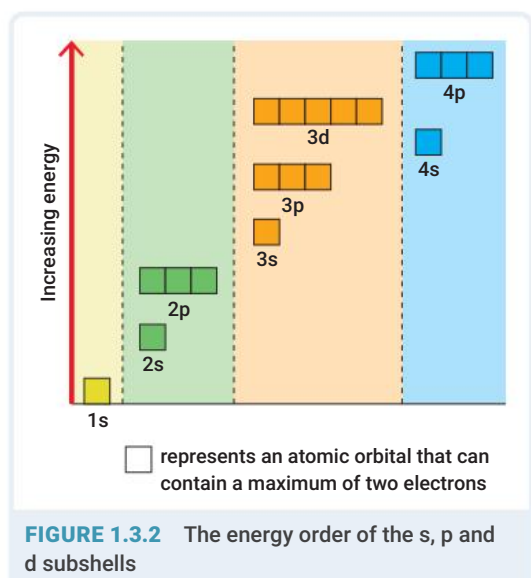


FIGURE 1.3.2 The energy order of the s, p and d subshells

Each subshell has a different level of energy. In a particular energy level, the s subshell always has the lowest energy, followed by the p subshell, then the d and then the f subshells. **Figure 1.3.2** shows all the subshells in order of increasing energy. Each box in the diagram represents an orbital of that subshell.

As shown in **Figure 1.3.2**, in higher energy levels, there is some overlap of energy for subshells in different energy levels. The 4s energy level has a lower energy than the 3d subshell. According to the Aufbau principle, lower energy levels fill up first, so the 4s energy level is occupied before the 3d subshell. The order of the subshells in increasing energy is $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d$.

Electron configurations and the periodic table

Although the concept of electron configurations can seem complex and confusing, the periodic table provides a very useful way of quickly determining the electron configuration of an element. Because the design of the periodic table was based on properties of the different elements, and these properties are determined by electron configurations, the periodic table is actually organised according to the position of electrons in filling these different subshells. **Figure 1.3.3** shows how each block of the periodic table corresponds to each subshell. Groups 1 and 2 are the s block, groups 13–18 are the p block, (notice that helium is an exception), groups 3–12 (also called the **transition metals**) are the d block and elements 57–71 (the **lanthanoids**) and 58–103 (the **actinoids**) are the f block.

The organisation of the periodic table lets us use it to predict which subshells have higher or lower energy than others. As atomic number increases, the subshells are filled in order of their increasing energy. You can see that for potassium and calcium, the 4s subshell is filled before the 3d subshell begins filling for scandium. It is important to identify and remember that each of the transition metal rows are filling a d subshell that belongs to the energy level of the period above (e.g. the d subshell in the fifth row of the periodic table is the d subshell in the fourth energy level, subshell 4d) and that each row of lanthanides fills an f subshell that belongs to the energy

transition metal a metallic element from the d block of the periodic table

lanthanoid a metallic element of atomic number 57–71 that along with the actinoids make up the f block of the periodic table

actinoid a metallic element of atomic number 89–103 that along with the lanthanoids make up the f block of the periodic table

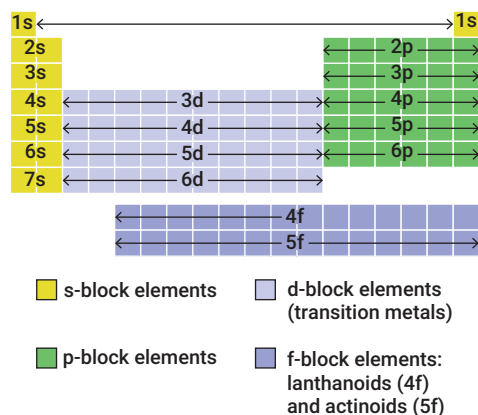


FIGURE 1.3.3 The subshells related to position in the periodic table

level two periods above (e.g. the row of lanthanides technically sits in period 6 and the outer shell for these elements is the 4f subshell). Identifying the order of the subshells is necessary for writing accurate and informative electron configurations for the elements.

Writing electron configurations

Tables 1.3.2 and 1.3.3 showed a simple way of writing electron configurations. However, now we can write electron configurations that show information about subshells and atomic orbitals.

For example, hydrogen has the electron configuration $1s^1$, where the superscript 1 represents the number of electrons in the subshell. Helium is $1s^2$. To write the configuration for lithium, the next electron needs to go into the next subshell, so lithium is $1s^2 2s^1$. When the 2s orbital is filled, the electrons need to go into 2p, then 3s and so on.

If we use the subshell filling order, $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p$ and so on, we can write an electron configuration that shows the energy levels, subshells and arrangement of electrons for any element. For example, if we look at a periodic table and Figure 1.3.3, we see aluminium is in period 3 and is the first element in the p block. This means all the level 1 and 2 subshells will be filled as will the 3s orbital, so its electron configuration will be $1s^2 2s^2 2p^6 3s^2 3p^1$.



Worksheet
Electron configurations

WORKED EXAMPLE 1.3.1

Write the electron configuration for neon (Ne).

ANSWER

1 Determine the number of electrons present in Ne.

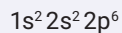
Neon has an atomic number of 10; therefore, it has 10 protons and 10 electrons.

2 Determine the electron configuration of Ne.

The order of filling of electron shells is:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p \dots$$

Therefore, we need to allocate electrons into subshells until all electrons have been assigned into a subshell:



noble gas an element from group 18 on the periodic table

As the atomic number increases, so does the length of the electron configuration. Chemists have a convention for writing a condensed version of the electron configuration using the configuration of the nearest previous group 18 element. Group 18 elements (called **noble gases**) are used because they are always the last element in a period. Consider the element gallium (Ga) in period 4. The last noble gas before gallium is argon (Ar) in period 3. Therefore, the condensed version of the electron configuration of gallium is written as $[\text{Ar}] 4s^2 3d^{10} 4p^1$, where $[\text{Ar}]$ represents the configuration of an argon atom.

WORKED EXAMPLE 1.3.2

Write the condensed electron configuration for magnesium (Mg).

ANSWER

1 Determine the last noble gas before magnesium.

Ne

2 Determine the condensed electron configuration.

The first 10 electrons of Mg can be represented by the configuration of neon $[\text{Ne}]$. The remaining two electrons need to be allocated into the next subshell with the lowest energy level after $2p^6$:

$[\text{Ne}] 3s^2$

Orbital notation

An alternative way of writing electron configurations for atoms is to use orbital notation. This method provides even more information about the location of electrons because it shows the number of electrons in each orbital of a subshell. This additional information helps explain more about the behaviour and properties of elements, as you will see in Chapter 3.

When using orbital notation there are two additional rules that must be applied.

- The first rule is the Pauli exclusion principle, which states that no two electrons can occupy the same quantum state. However, the maximum number of electrons that a single orbital can have is *two*, not one. This means that the two electrons in any orbital in a subshell must have different quantum states. The two states differ by their magnetic spin quantum numbers, which can be thought of as 'spin up' or 'spin down'. Scientists designate electrons in the same orbital by arrows, pointing up or down, to distinguish their spin states.
- The second rule is **Hund's rule of maximum multiplicity**, which applies when there is more than one orbital in a subshell. This states that for an atom in its lowest energy, or **ground state**, every orbital in a subshell must contain one electron before any orbital contains two electrons.

For example, nitrogen has the electron configuration of $1s^2 2s^2 2p^3$. The preferred ground state arrangement of the electrons in nitrogen is shown in **Figure 1.3.4**. Notice each of the three $2p$ orbitals has one electron, in accordance with Hund's rule.

Hund's rule of maximum multiplicity every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied

ground state the lowest energy electron configuration of an atom

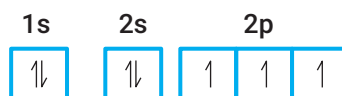


FIGURE 1.3.4 The orbital notation for the ground state electron configuration of nitrogen



TABLE 1.3.6 The electronic configurations of the first 18 elements using subshell and orbital notation

Element	Atomic number	Electron configuration	Orbital notation
Hydrogen	1	1s ¹	[↑]
Helium	2	1s ²	[↑↓]
Lithium	3	1s ² 2s ¹	[↑↓] [↑]
Beryllium	4	1s ² 2s ²	[↑↓] [↑↓]
Boron	5	1s ² 2s ² 2p ¹	[↑↓] [↑↓] [↑][][]
Carbon	6	1s ² 2s ² 2p ²	[↑↓] [↑↓] [↑][↑][]
Nitrogen	7	1s ² 2s ² 2p ³	[↑↓] [↑↓] [↑][↑][↑]
Oxygen	8	1s ² 2s ² 2p ⁴	[↑↓] [↑↓] [↑↓][↑][↑]
Fluorine	9	1s ² 2s ² 2p ⁵	[↑↓] [↑↓] [↑↓][↑↓][↑]
Neon	10	1s ² 2s ² 2p ⁶	[↑↓] [↑↓] [↑↓][↑↓][↑↓]
Sodium	11	1s ² 2s ² 2p ⁶ 3s ¹	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑]
Magnesium	12	1s ² 2s ² 2p ⁶ 3s ²	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓]
Aluminium	13	1s ² 2s ² 2p ⁶ 3s ² 3p ¹	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑][][]
Silicon	14	1s ² 2s ² 2p ⁶ 3s ² 3p ²	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑][↑][]
Phosphorus	15	1s ² 2s ² 2p ⁶ 3s ² 3p ³	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑][↑][↑]
Sulfur	16	1s ² 2s ² 2p ⁶ 3s ² 3p ⁴	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑↓][↑][↑]
Chlorine	17	1s ² 2s ² 2p ⁶ 3s ² 3p ⁵	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑↓][↑↓][↑]
Argon	18	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑↓][↑↓][↑↓]

Electron configurations for elements in period 4

As shown in Table 1.3.6, element number 18, argon, has the configuration $1s^2 2s^2 2p^6 3s^2 3p^6$. This follows the Aufbau principle where subshells are filled in order of increasing subshell energy, as follows: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$.

After the 3p subshell, the next orbitals filled are those in the 4s subshell, rather than the 3d. Consequently, potassium and calcium, elements 19 and 20, have electrons in the fourth level (period 4, s block on the periodic table), despite the third level not being completely filled. For scandium, element 21, the 4s subshell is filled, so the remaining electron is held in the 3d subshell, which is of lower energy than the 4p subshell. This pattern continues until element 31, gallium; because the 3d subshell is now complete, the remaining electron is held

in the next available orbital, the 4p subshell. The full electron configuration of gallium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$.

The electron configurations of elements 19–36 are shown in **Table 1.3.7**. In this table, the convention of the [Ar] symbol has been used to save space.

TABLE 1.3.7 The electron configurations for elements 19–36

Element	Atomic number	Electron configuration	Orbital diagram																		
Potassium	19	[Ar] 4s ¹	[Ar] <table><tr><td>1</td></tr></table>	1																	
1																					
Calcium	20	[Ar] 4s ²	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table>	1	1																
1																					
1																					
Scandium	21	[Ar] 4s ² 3d ¹	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table>	1	1	1															
1																					
1																					
1																					
Titanium	22	[Ar] 4s ² 3d ²	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table>	1	1	1	1														
1																					
1																					
1																					
1																					
Vanadium	23	[Ar] 4s ² 3d ³	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table>	1	1	1	1	1													
1																					
1																					
1																					
1																					
1																					
Chromium	24	[Ar] 4s ¹ 3d ⁵	[Ar] <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1												
1																					
1																					
1																					
1																					
1																					
1																					
Manganese	25	[Ar] 4s ² 3d ⁵	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1											
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Iron	26	[Ar] 4s ² 3d ⁶	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1										
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Cobalt	27	[Ar] 4s ² 3d ⁷	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1									
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Nickel	28	[Ar] 4s ² 3d ⁸	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1								
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Copper	29	[Ar] 4s ¹ 3d ¹⁰	[Ar] <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1								
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Zinc	30	[Ar] 4s ² 3d ¹⁰	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1						
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Gallium	31	[Ar] 4s ² 3d ¹⁰ 4p ¹	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td></td></tr></table> <table><tr><td></td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1	1					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Germanium	32	[Ar] 4s ² 3d ¹⁰ 4p ²	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td></td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1	1	1				
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Arsenic	33	[Ar] 4s ² 3d ¹⁰ 4p ³	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1			
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Selenium	34	[Ar] 4s ² 3d ¹⁰ 4p ⁴	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1		
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Bromine	35	[Ar] 4s ² 3d ¹⁰ 4p ⁵	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table> <table><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1		
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
Krypton	36	[Ar] 4s ² 3d ¹⁰ 4p ⁶	[Ar] <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table> <table><tr><td>1</td></tr><tr><td>1</td></tr></table>	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1	1
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					
1																					

Electron configuration of ions

The subshell notation can be used to represent the electron configuration of ions in the same way as for atoms.

According to the Aufbau principle, an atom loses electrons from the highest occupied energy orbital first. For example, potassium, with configuration [Ar] $4s^1$ loses the 4s electron and becomes a positive ion with configuration [Ar], the same configuration as argon. When an atom gains electrons, the electrons are added into the next available orbital. For example, bromine, with configuration [Ar] $4s^2 3d^{10} 4p^5$, has the electron added to the 4p subshell and becomes a negative ion with the same electron configuration as krypton: [Ar] $4s^2 3d^{10} 4p^6$.



Syllabus link
Chapter 5 discusses the formation of ions and ionic bonding in more detail.

WORKED EXAMPLE 1.3.3

Write the electron configuration for a sodium ion (Na^+).

ANSWER

1 Determine the number of electrons in a sodium atom.

Sodium has an atomic number of 11, so it has 11 electrons.

2 Identify the impact of sodium becoming an ion.

To form an ion with a charge of $1+$, the sodium atom must lose its highest energy electron.

3 Determine the electron configuration of sodium ion.

Therefore, using the 10 remaining electrons, the electron configuration for a sodium ion:



WORKED EXAMPLE 1.3.4

Write the electron configuration for a sulfide ion (S^{2-}).

ANSWER

1 Determine the number of electrons in a sulfur atom.

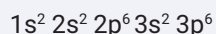
Sulfur has an atomic number of 16, so it has 16 electrons.

2 Identify the impact of sulfur becoming an ion.

To form an ion with a charge of $2-$, the sulfur atom must gain two electrons.

3 Determine the electron configuration of sulfur ion.

Therefore, the electron configuration for this ion is:



Chromium and copper – apparent exceptions in electron configuration

The electron configurations for chromium and copper do not follow the Aufbau principle as they do not fill lower energy subshells before filling higher ones (see Table 1.3.7). There are other exceptions to the Aufbau principle within the transition metals too, but this section will focus on understanding and recognising copper and chromium as exceptions.

The Aufbau principle says that lower energy subshells are more stable and this is where electrons will tend to be. So when filling the subshells, electrons are usually added to the 4s subshell before the 3d subshell. However, the stability of subshells is also affected by how full they are.

The stability for subshell occupancy in order from most to least stable is empty, full, half-full, partially full.

For copper, the arrangement where the 4s orbital is half-filled and the 3d orbital is full is more stable than a filled 4s and partially filled 3d orbital. For chromium, the more stable arrangement is a half-filled 4s and completely filled 3d rather than a full 4s and partially full 3d orbital (**Table 1.3.8**).

This type of rearrangement in defiance of the Aufbau principle only occurs when subshells are very close in energy level such as between s and d subshells.

TABLE 1.3.8 The electron configurations for copper and chromium

Element	Atomic number	Expected electron configuration	Actual electron configuration
Chromium	24	[Ar] 4s ² 3d ⁴	[Ar] 4s ¹ 3d ⁵
Copper	29	[Ar] 4s ² 3d ⁹	[Ar] 4s ¹ 3d ¹⁰

LEARNING CHECK 1.3

DESCRIBING

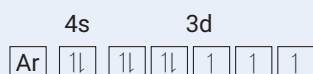
- State** the maximum number of electrons that can be found in the first three energy levels.
- Explain** what information the electron configuration of an atom provides.
- Explain** how the Aufbau principle relates to the electron configuration of an atom.
- State** the relative energies of the s, p and d orbitals.
- Identify** the relationship between the structure of the periodic table and the s, p, d and f subshells.

APPLYING

- Determine** the electron configurations for the following, using s, p, d, f notation.
 - Nitrogen atom
 - Phosphide P³⁻ ion
 - Scandium atom
 - Chromium atom
 - Germanium Ge⁴⁺ ion
 - Represent** the electron configurations for i and iii using condensed notation.
 - Represent** the configuration of iv using orbital notation.
- Identify** the elements that have the following electron configurations.
 - 2,8,2
 - 1s²
 - 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁶
- The following electron configurations are incorrect. **Identify** each element, **explain** why the configuration is incorrect and **state** the correct configuration for that element in the s, p, d, f form.
 - 2,9
 - 2,8,10
- Explain** why the electron configuration of a copper atom is unusual compared with the other atoms in the first row of the transition metals.

ANALYSING

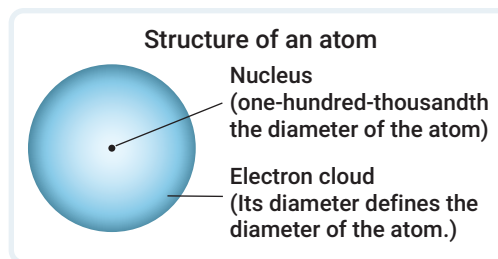
- Examine** the orbital notation configuration of cobalt.



Identify where the Aufbau principle, Pauli exclusion principle and Hund's rule have been applied.

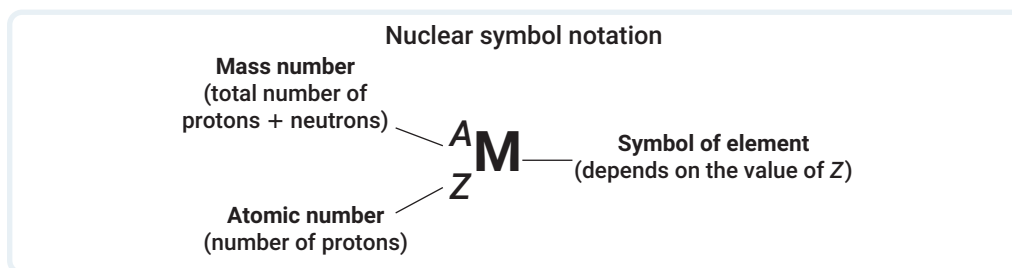
The atom

- An atom is composed of a central nucleus, which contain the protons and neutrons, surrounded by an electron cloud with the electrons located in distinct energy levels.



Representing elements

- Elements are represented by a unique name and symbol and each element is identified by the number of protons in the nucleus of its atoms.
- The information provided in nuclear symbol notation allows the determination of the number of protons, neutrons and electrons in atoms and ions.

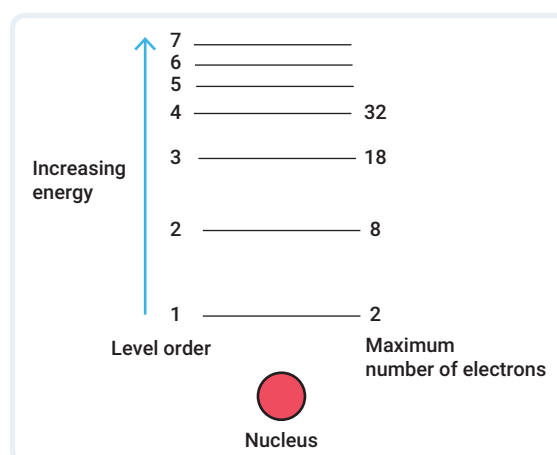


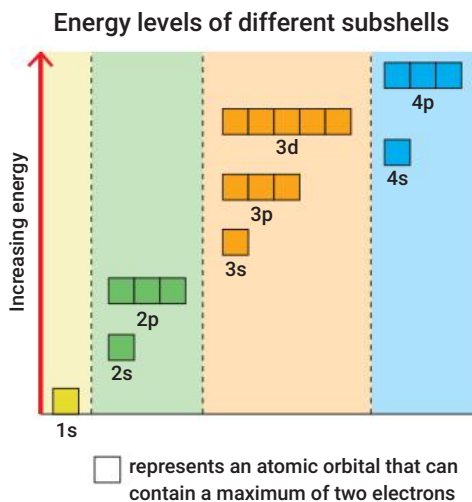
- Information about atomic and mass numbers is provided in the periodic table, which lists all known elements.

	Carbon	Element name
Atomic number	6	
	C	Symbol of element
Relative atomic mass	12.01	

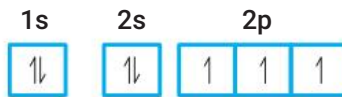
Electron configuration

- Electrons can only be located in discrete energy levels in the electron cloud surrounding the nucleus.
- Each energy level, n , can hold only a specific maximum number of electrons equal to the value of $2n^2$.

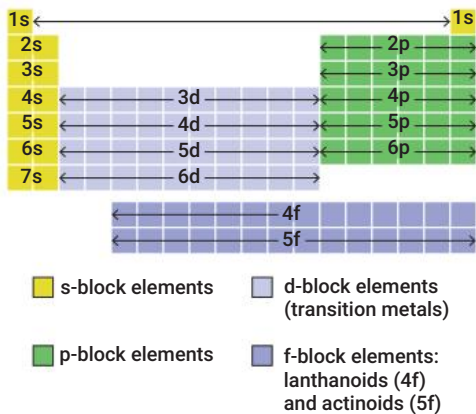




- ### An example of orbital notation



- ## Electronic structure and the periodic table



Element	Atomic number	Expected electron configuration	Actual electron configuration
Chromium	24	[Ar] 4s ² 3d ⁴	[Ar] 4s ¹ 3d ⁵
Copper	29	[Ar] 4s ² 3d ⁹	[Ar] 4s ¹ 3d ¹⁰

MULTIPLE CHOICE

- Which of these statements is *false*?
A Protons have a positive charge.
B Electrons are negatively charged.
C The nucleus of an atom is positively charged.
D Electrons are found in the nucleus of an atom.
- The atomic number of an element is equal to the:
A number of protons plus neutrons in a neutral atom of the element.
B number of neutrons in a neutral atom of the element.
C number of electrons in a neutral atom of the element.
D mass number of the element.
- Which of the following is not the electron configuration of a transition metal atom?
A $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
B $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
C $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$
D $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$
- An ion that contains seven protons, eight neutrons and nine electrons will have a mass number and charge corresponding to:

	Mass number	Charge
A	15	2+
B	8	2-
C	15	2-
D	8	2+

- What is the ground state electron configuration for S^{2-} ?
A $1s^2 2s^2 2p^6 3s^2 3p^4 4s^2$
B $1s^2 2s^2 2p^6 3s^2 3p^2$
C $1s^2 2s^2 2p^6 3s^2 3p^4$
D $1s^2 2s^2 2p^6 3s^2 3p^6$
- The 3d subshell has:
A three orbitals and can hold up to six electrons.
B five orbitals and can hold up to 10 electrons.
C three orbitals and can hold up to nine electrons.
D five orbitals and can hold up to 15 electrons.
- Which of the following subshells does *not* exist in any atom?
A 4s
B 3p
C 2p
D 2d

8. The ion X^{2+} has the following electronic configuration.
 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
In the periodic table, element X would be located in:
A group 2.
B group 4.
C group 6.
D group 18.
9. Which set of the following contains particles that have the same electronic configuration?
A Ca^{2+} , K^+ , Ar, Cl^-
B Li, Be, B, C
C O^{2-} , F, Na^+ , Ne
D Li^+ , Na^+ , K^+ , Cs^+
10. Which one of the following statements *best* expresses the idea of the Pauli exclusion principle?
A An orbital can never contain three or more electrons.
B An orbital must contain two electrons.
C An orbital may contain one electron.
D An orbital must always contain one or two electrons.

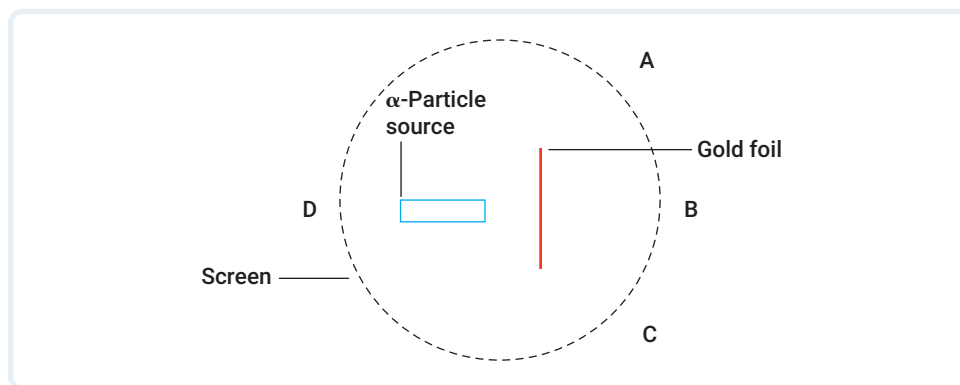
SHORT RESPONSE

11. **Describe** the information about the structure of an atom that can be determined by the:
a group an element is in on the periodic table
b period an element is in on the periodic table.
12. Consider the elements sulfur, neon, chlorine and calcium. **Deduce** which element has:
a its highest energy level completely filled
b the highest occupied energy level
c four electrons in the 3p orbitals
d a completely filled energy level when it gains one electron.
13. The atomic number of potassium is 19.
a **Identify** the electron configuration, in terms of energy levels and subshells, for the potassium atom in its ground state.
b **Identify** the electron configuration using orbital notation, in terms of energy levels and subshells, for the potassium ion (K^+).

DATA ANALYSIS

14. Apply understanding

In the Geiger–Marsden gold foil experiment, a thin piece of gold foil was bombarded with α -particles as shown in the simplified diagram.



Observations: Approximately 1 in 8000 alpha particles (<0.001%) bounced off the gold foil at large angles.

- a Referring to the diagram, use the letters to identify where:
 - i the particles were expected to be detected
 - ii detection was the most unexpected.
 - b **State** three pieces of information about the nucleus that Rutherford deduced from the results from this experiment.
 - c The development of quantum theory added to Rutherford's atomic model. Briefly **describe** these additions.
15. A neutral atom has eight neutrons and seven electrons. Use this information to represent the atom using ${}^A_Z\text{M}$ notation.



European southern Observatory/Science Photo Library (SPL)

SYLLABUS
DOT POINTS

SCIENCE UNDERSTANDING

- Discriminate between the terms atomic number (Z), mass number (A) and isotopes of an element.
- Apply the nuclear symbol notation ${}^A_ZM$ to determine the number of protons, neutrons and electrons in atoms, ions and isotopes.
- Describe that isotopes are atoms of the same element that have different numbers of neutrons.
- State that isotopes can be represented in the form ${}^AX$ (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 $\frac{1}{12}$ the mass of an atom of carbon-12.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Some remarkable techniques in several different fields, such as the detection of medical conditions and the dating of ancient objects, can be attributed to our understanding of isotopes.

Worksheets

- Isotopes
- Isotopes and relative abundance



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ Atoms are composed of protons, neutrons and electrons.
- ✓ Protons, neutrons and electrons each have a unique combination of mass and charge.
- ✓ Atomic number (Z) is the number of protons in an atom and mass number (A) is the number of protons and neutrons in the nucleus of an atom.
- ✓ The nuclear symbol notation ${}_Z^A\text{M}$ can be used to represent an atom of an element and to determine the number of protons, neutrons and electrons in an atom.
- ✓ Ions are formed when an atom loses or gains electrons.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ define 'isotope'
- ✓ apply the nuclear symbol notation ${}_Z^A\text{M}$ to identify isotopes
- ✓ state representations for isotopes, including ${}^A\text{X}$ (IUPAC) and X-A
- ✓ identify that isotopes of the same element have the same electron configuration and similar chemical properties, but different physical properties
- ✓ define 'relative atomic mass'
- ✓ distinguish between relative atomic mass and isotopic mass
- ✓ explain how relative atomic mass is calculated
- ✓ solve problems related to the calculation of relative atomic mass
- ✓ consider why isotopes may undergo radioactive decay.

isotopes two or more atoms of the same element with the same number of protons but with different masses, resulting from their different number of neutrons



Weblink
Isotopes

2.1 Different forms of the same element

Isotopes are atoms of the same element that have the same number of protons but different numbers of neutrons in their nuclei. This means that two isotopes of the same element will have the same atomic number, but different mass numbers. Isotopes of the same element show some similarities and some differences in their properties, and their relative abundance on Earth is different. Isotopes of different elements can be naturally occurring or synthesised for a specific purpose.

Isotopes are named by their mass number. This can be represented by writing the name (or symbol) of the element, followed by a hyphen and the mass number of the isotope. For example, chlorine has two isotopes that can be written as chlorine-35 and chlorine-37 or Cl-35 and Cl-37.

In Chapter 1, the nuclear symbol notation ${}_Z^A\text{M}$ was used to give information about the number of protons, neutrons and electrons in an atom. This notation is also used for isotopes. When referring to isotopes, the nuclear symbol notation may be abbreviated by leaving off Z because atoms of the same element have the same number of protons. IUPAC shows this as ${}^A\text{X}$, where A is the mass number and X is the element symbol. For example, carbon has several isotopes, all containing six protons but different numbers of neutrons. Carbon-12 or ${}^{12}_6\text{C}$ or ${}^{12}\text{C}$ has six neutrons, whereas carbon-14 or ${}^{14}_6\text{C}$ or ${}^{14}\text{C}$ has eight neutrons. Different elements have different numbers of naturally occurring isotopes and some, such as fluorine, have only one.

Table 2.1.1 gives two examples of the isotopes of different elements and shows the number of protons, neutrons and electrons in each isotope.

TABLE 2.1.1 The numbers of protons, neutrons and electrons in isotopes of carbon and uranium

Name	Symbol	Number of protons	Number of neutrons	Number of electrons (neutral atom)
Carbon-12	$^{12}_6\text{C}$	6	6	6
Carbon-13	$^{13}_6\text{C}$	6	7	6
Carbon-14	$^{14}_6\text{C}$	6	8	6
Uranium-234	$^{234}_{92}\text{U}$	92	142	92
Uranium-235	$^{235}_{92}\text{U}$	92	143	92
Uranium-238	$^{238}_{92}\text{U}$	92	146	92



Worksheet
Isotopes

Chemical properties

Chemical properties of atoms or elements are related to how an atom bonds and participates in chemical reactions. Examples are reactivity, flammability and acidity. The chemical properties of an atom or element are determined by the electron configuration. Since isotopes have the same number and arrangement of electrons, their chemical properties are very similar. For example, water (H_2O) is formed in exactly the same reaction, no matter whether the isotope of hydrogen involved is ^1H or ^2H .

chemical property a property of a substance that relates to the ways in which it bonds or reacts with other substances, e.g. acidity and flammability

Physical properties

Physical properties of elements are the properties it exhibits without chemically changing into something else. Examples are melting point, density and mass. Physical properties are heavily dependent on the mass of the atoms in the elements and since isotopes have different masses, they can have different physical properties. For example, H_2O made from hydrogen-1 isotopes has a density of 0.9982 g mL^{-1} but ‘heavy water’, which has H_2O molecules containing hydrogen-2 isotopes instead, has a density of 1.1056 g mL^{-1} . In addition, because of the different numbers of neutrons, some isotopes of certain elements are radioactive whereas others are not.

physical property a property of a substance that does not involve a chemical reaction, e.g. melting point and electrical conductivity

Stable and unstable isotopes

Some isotopes of elements are stable because the forces in the nucleus are balanced. Other isotopes have unstable nuclei in which the forces are not balanced. In some nuclei, the repulsive forces are stronger than the attractive forces. In other nuclei, the attractive forces are stronger than the repulsive forces. In both cases, an unstable nucleus forms and the nucleus undergoes **radioactive decay** to become stable. During radioactive decay, high-energy particles or radiation is emitted. If uncontrolled, these high-energy particles can be very dangerous, but if harnessed appropriately, they can be used for medical diagnosis and treatments, and in other industries.

radioactive decay the process by which an unstable nucleus changes into a more stable nucleus by releasing small particles and/or electromagnetic radiation

LEARNING CHECK 2.1

DESCRIBING

- 1 **Describe** what 'isotope' means.
- 2 **State** the different representations that can be used for isotopes. Include an example.

APPLYING

- 3 **Discriminate** between isotopes nitrogen-14 and nitrogen-13 in terms of differences in the numbers of subatomic particles.
- 4 **Explain** why isotopes of the same element have the same chemical properties.
- 5 Cobalt-60, which is used for irradiation of some cancers, is made from cobalt-59.
 - a **Identify** what needs to be done to Co-59 to produce the Co-60 isotope.
 - b **Describe** two differences in the physical properties of these isotopes.

2.2 Relative atomic mass

An atom is extremely small and so there are different ways that atomic mass can be measured. We would rarely measure an atom's mass in kilograms because it is such a small number. Sometimes, the mass of an atom is measured using a very small mass unit known as the atomic mass unit (amu), where $1 \text{ amu} = 1.66 \times 10^{-27} \text{ kg}$. Instead, scientists use the term **relative atomic mass (A_r)**. The word 'relative' means that the mass of an atom is being compared with the mass of another atom that scientists use as a standard. It is important to note that relative atomic mass has no units because it is a ratio of masses.

Scientists arbitrarily picked carbon-12 as the standard for comparison to calculate relative atomic mass. Hydrogen and oxygen were considered and even used for a time, but carbon was selected for ease of use in experiments and because it is a very common element on Earth. The relative atomic mass of an atom is the mass of the atom relative to the mass of a carbon-12 atom, which has six protons and six neutrons. The mass of a carbon-12 atom is considered to be exactly 12 and all other atoms are compared to this.

Magnesium-24 has twice the mass of carbon-12 so has a mass of $\frac{24}{12} \times 12 = 24$. Hydrogen-1 is one-twelfth the mass of carbon-12, so it has a mass of $\frac{1}{12} \times 12 = 1$. A visual representation of atomic mass comparisons for He-4, C-12 and Ti-48 are shown in **Figure 2.2.1**.

relative atomic mass (A_r)
the average mass of
one atom of an element,
relative to the mass of
carbon-12



Weblink

Relative atomic
mass calculations

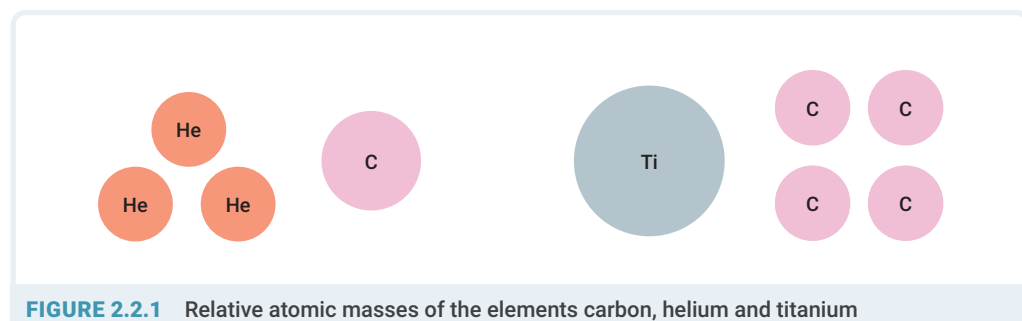


FIGURE 2.2.1 Relative atomic masses of the elements carbon, helium and titanium

Roland Smith

Isotopes of the same element have different atomic masses. The term ‘isotopic mass’ is used to refer to the mass of a particular isotope of an element. They also have different abundances on Earth, which means they are found in different amounts. For example, two of carbon’s isotopes are carbon-12 and carbon-13. On Earth, 98.9% of all carbon is carbon-12 and only 1.1% is carbon-13. Other isotopes of carbon, such as carbon-14, are present in such low levels they are not considered here.

The periodic table displays the relative atomic mass for each element, rather than the mass number. The mass and abundance of naturally occurring isotopes of an element are used to calculate these mass numbers, so it is not usually a whole number. For the two isotopes of carbon (^{12}C and ^{13}C), the generalised equation used to calculate relative atomic mass is:

$$\text{Relative atomic mass} = \frac{(\text{percentage abundance} \times \text{isotopic mass}) + (\text{percentage abundance} \times \text{isotopic mass})}{100}$$

using the information above, this becomes:

$$\begin{aligned} &= \frac{(98.9 \times 12) + (1.1 \times 13)}{100} \\ &= 12.01 \end{aligned}$$



Worksheet
Isotopes and relative
abundance

Worked example 2.2.1 shows how to calculate the relative atomic mass seen in the periodic table for any element. Worked example 2.2.2 shows how to calculate the percentage abundance for an isotope if you are provided with the relative isotopic mass.

WORKED EXAMPLE 2.2.1

Chlorine has two isotopes with relative isotopic masses of 35 and 37, respectively. The percentages of each isotope present on Earth are 75% for chlorine-35 and 25% for chlorine-37.

What is chlorine’s relative atomic mass?

ANSWER

1 Write the correct formula to use.

$$\text{Relative atomic mass of chlorine} = \frac{(\text{percentage abundance} \times \text{isotopic mass}) + (\text{percentage abundance} \times \text{isotopic mass})}{100}$$

2 Substitute into the formula and calculate the answer.

$$\begin{aligned} \text{Relative atomic mass of chlorine} &= \frac{(75 \times 35) + (25 \times 37)}{100} \\ &= 35.5 \end{aligned}$$

Chlorine’s relative atomic mass is 35.5.

WORKED EXAMPLE 2.2.2

The element lithium has two isotopic forms: lithium-6 and lithium-7. If lithium's relative atomic mass is 6.94, what is the percentage abundance of each isotope?

ANSWER

1 Write the correct formula to use.

$$\text{Relative atomic mass of lithium} = \frac{(\text{percentage abundance} \times \text{isotopic mass}) + (\text{percentage abundance} \times \text{isotopic mass})}{100}$$

2 Identify the abundance of each isotope.

Because the abundance of each isotope is unknown, assign the percentage abundance of one isotope as 'x'. If the abundance of lithium-6 is x, then the abundance of lithium-7 is 100 - x, because the total of the abundances must be 100. So:

$$6.94 = \frac{(x \times 6) + ((100 - x) \times 7)}{100}$$

Simplify the expression:

$$694 = (6x + 700 - 7x)$$

$$x = 700 - 694$$

$$x = 6$$

3 Determine the percentage abundance of each.

Percentage abundance of lithium-6 is 6%.

Percentage abundance of lithium-7 is 100 - 6 = 94%.

LEARNING CHECK 2.2

DESCRIBING

- 1 Define** 'relative atomic mass'.
- 2 Explain** how the relative atomic mass shown in the periodic table is calculated.

APPLYING

- 3** A given element has a mass three times that of carbon. **Calculate** its relative atomic mass.
- 4** Copper has two isotopes: copper-63 with an abundance of 69.17% and copper-65 with an abundance of 30.83%. **Calculate** the relative atomic mass of copper.
- 5** Vanadium (V) has two naturally occurring isotopes: vanadium-50 and vanadium-51. Given that the relative atomic mass for vanadium in the periodic table is 50.94, **predict** which isotope is more abundant and explain your reasoning.
- 6 Calculate** the exact abundances of the vanadium isotopes in Question 5.

Isotopes

- Isotopes of an element have the same atomic number but different mass numbers because they have different numbers of neutrons.
- Isotopes can be represented using different formats including ${}^A_Z\text{M}$, ${}^A\text{X}$ and X-A .
- Some isotopes of an element may be unstable and undergo radioactive decay.
- Isotopes of an element have the same electron configuration and similar chemical properties, but different physical properties.

Numbers of protons, neutrons and electrons in isotopes of carbon and uranium

Name	Symbol	Number of protons	Number of neutrons	Number of electrons (neutral atom)
Carbon-12	${}^{12}_6\text{C}$	6	6	6
Carbon-13	${}^{13}_6\text{C}$	6	7	6
Carbon-14	${}^{14}_6\text{C}$	6	8	6
Uranium-234	${}^{234}_{92}\text{U}$	92	142	92
Uranium-235	${}^{235}_{92}\text{U}$	92	143	92
Uranium-238	${}^{238}_{92}\text{U}$	92	146	92

Isotopes and relative atomic mass

- The relative atomic mass of each element can be calculated using the percentage abundance of each isotope.

$$\text{Relative atomic mass} = \frac{(\text{percentage abundance} \times \text{isotopic mass}) + (\text{percentage abundance} \times \text{isotopic mass})}{100}$$

CHAPTER EXAM

MULTIPLE CHOICE

1. ^{12}C and ^{14}C are regarded as isotopes of the same element because their:
- A atomic numbers differ.
 - B atoms have the same mass number.
 - C atoms have six protons in their nucleus.
 - D atoms have similar chemical properties.

Questions 2 and 3 refer to the following information.

The numbers of protons, neutrons and electrons present in each of five atoms are shown in the table.

Atom	Number of protons	Number of neutrons	Number of electrons
I	6	8	6
II	7	7	7
III	7	7	8
IV	8	8	8
V	8	10	8

2. Which of the two atoms are isotopes of the same element?
- A II and III
 - B IV and V
 - C III and IV
 - D I and IV
3. If X were the symbol for the element of atom III, how would it be represented?
- A $^{14}_7\text{X}$
 - B $^{15}_7\text{X}$
 - C $^{15}_8\text{X}$
 - D $^{14}_8\text{X}$
4. The element hydrogen consists of three isotopes: protium (^1H), deuterium (^2H) and tritium (^3H). If the relative atomic mass of hydrogen is 1.01, which isotope is the most abundant?
- A Protium
 - B Deuterium
 - C Tritium
 - D It is not possible to determine which isotope is most abundant without examining the mass spectrum of hydrogen.
5. Which of the following is a correct notation for isotope of boron that has a mass of 11?
- A Bo-11
 - B $^{11}_{11}\text{Bo}$
 - C ^{11}B
 - D $^{11}_5\text{B}$

6. Which of the following is a definition of the relative atomic mass of an element?
- A the weighted average of the masses of all possible isotopes of the element on a scale that naturally occurring carbon has a mass of 12 exactly
 - B the weighted average of the masses of the most abundant isotopes of the element on a scale that naturally occurring carbon has a mass of 12 g exactly
 - C the weighted average of the isotopic masses of the element on a scale in which an atom of carbon-12 is taken to be 12 exactly
 - D the weighted average of the isotopic masses on a scale such that an atom of naturally occurring carbon has a mass of 12 g exactly

7. Data for an element is given in the table below.

Relative isotopic mass	63.9	65.9	66.9	67.9
Percentage abundance	48.9	27.8	4.10	19.2

What is the relative atomic mass of the element closest to, based on the data provided?

- A 65.2
 - B 65.3
 - C 65.4
 - D 65.5
8. The isotopic masses of three of the four isotopes of lead are shown in the following table.

Isotope	Mass number	Abundance (%)
Lead-204	204	4
Lead-206	206	24
Lead-207	207	21

If the relative atomic mass of lead is 207.2, which of the following is the correct mass number and abundance of the remaining isotope of lead?

	Mass number	Abundance (%)
A	205	8
B	205	51
C	208	8
D	208	51

9. A carbon-12 atom has a mass of approximately 20.4×10^{-27} kg. The mass of a phosphorus-31 atom would therefore be closest to:
- A 2.6×10^{-27} kg.
 - B 31×10^{-27} kg.
 - C 52.7×10^{-27} kg.
 - D 632×10^{-27} kg.
10. ${}^{230}_{90}\text{Th}$ undergoes radioactive decay to form ${}^{226}_{88}\text{Ra}$. In order for this to happen, the thorium-230 atom must change so that overall it has:
- A lost two protons and four electrons.
 - B lost two neutrons, two protons and two electrons.
 - C lost two protons, two electrons and four neutrons.
 - D lost two neutrons and gained four protons and electrons.

SHORT RESPONSE

11. a **Define** 'isotope'.
b **Explain** which properties of isotopes are similar and which are different.
12. **Explain** why there are only 118 known elements but many more different types of atoms.

CROSS-CHAPTER QUESTION

13. Consider the isotopes $^{31}_{15}\text{M}$, $^{51}_{23}\text{M}$, $^{33}_{15}\text{M}$, $^{127}_{52}\text{M}^-$, $^{52}_{23}\text{M}^+$, $^{127}_{53}\text{M}$, $^{58}_{28}\text{M}$, where M may be the same element or a different element.
- a **Determine** how many different elements there are.
b **Identify** isotopes of the same element.
c **Identify** the isotopes with the same number of electrons.
d **Identify** the isotopes with the same number of neutrons.

DATA ANALYSIS

14. **Analyse data**

Silicon has three isotopes, as shown in the following table.

Isotopic mass	Abundance (%)
28	92.23
29	4.68
30	3.09

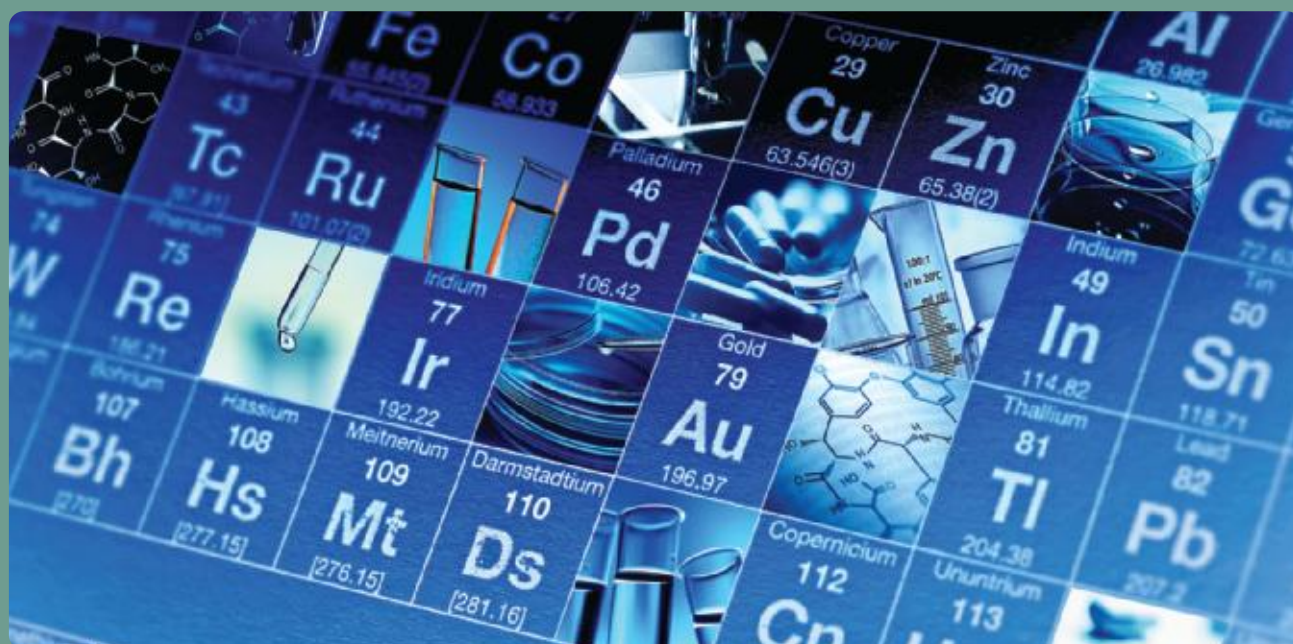
Calculate the relative atomic mass of silicon.

15. **Apply understanding**

Bromine has two stable isotopes: bromine-79 and bromine-81. Given the relative atomic mass of bromine, **calculate** the percentage abundance of each isotope.

35
Br
79.90

The periodic table and trends



isak55/Shutterstock.com

SYLLABUS
DOT POINTS

SCIENCE UNDERSTANDING

- 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, 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, 1st ionisation energy and electronegativities to determine periodic trends, patterns and relationships.
- Explain how successive ionisation energy data is related to the electron configuration of an atom.
- Identify that the valency is a measure of the number of bonds that an atom can form.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

The periodic table is the chemist's method of arranging all the elements in a systematic way so that regularly repeating or 'periodic' trends and their properties can be seen. This means that the properties of an individual element can be easily predicted from its position in the table.

Practical

- Reactivity of group 2 metals – the alkaline earth metals

Worksheets

- Atomic radius and ionisation energy
- Ionisation energy and electronegativity



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Atoms are composed of a nucleus containing protons and neutrons, and electrons in distinct energy levels.
- ✓ An element is composed of only one type of atom, which is identified by its atomic number and represented by a unique name and symbol.
- ✓ The relative atomic mass of an element is the average mass of one atom of the element relative to the mass of carbon-12.
- ✓ The average mass of an atom of an element is calculated from the mass and abundance of isotopes of that element.
- ✓ An ion is formed when an atom loses or gains electrons.
- ✓ The Aufbau principle, Hund's rule and the Pauli exclusion principle are applied to write the electron configuration of a neutral atom of an element.
- ✓ The structure of the periodic table is based on increasing atomic number and is organised into rows called periods and columns called groups.
- ✓ The periodic table is arranged into four blocks associated with the four energy subshells – s, p, d and f – and this organisation can be used to determine the electron configuration of atoms.

LEARNING OBJECTIVES

By the end of this chapter, you should be able to:

- ✓ identify the structure of the periodic table
- ✓ recognise and explain trends in chemical and physical properties down groups and across periods
- ✓ identify the trends in and compare the reactivity of alkali metals and halogens
- ✓ identify trends in metallic and non-metallic characteristics of elements
- ✓ identify that oxides are formed when an element reacts with oxygen
- ✓ identify the chemical nature of the oxides formed by period 3 elements
- ✓ analyse data for atomic radii, valencies, ionic radii, first ionisation energy and electronegativities to determine periodic trends, patterns and relationships
- ✓ relate changes in successive ionisation energies to the electron configuration of an element
- ✓ explain periodic trends and relationships in terms of electron configuration and atom structure.

3.1 The periodic table of elements

Arrangement of the periodic table of elements

As you learnt in Chapter 1, an element is a substance made up of only atoms with the same number of protons (atomic number), although the atoms of an element may have different numbers of neutrons (isotopes) or different numbers of electrons (ions).

On the periodic table, the elements are listed in order of atomic number, from element 1, hydrogen (H), to element 118, oganesson (Og).

The periodic table was developed by the Russian chemist Dmitri Mendeleev and published in 1869. Mendeleev organised the names of the elements with their atomic masses and other properties in rows (periods) and columns (groups) by first ordering the elements by mass, and then grouping them so that elements with similar properties were in the same vertical column.



Syllabus link

The periodic table was introduced in Chapter 1.

Mendeleev was so confident in his patterns and groupings that he left gaps for elements not yet discovered, predicting their mass and properties. Later, these elements were discovered and named. They included gallium, germanium and scandium.

Groups and periods in the periodic table

Elements with similar chemical and physical properties are found in vertical groups numbered 1–18. For example, group 2 includes the elements beryllium, magnesium, calcium and strontium, while group 15 includes nitrogen, phosphorus and arsenic. The groups are numbered, but some of them have special names to classify the elements within them into specific element types that have similar properties (**Figure 3.1.1**):

- Group 1 elements are also known as the alkali metals
- Group 2 elements are also known as the alkali earth metals.
- Group 17 elements are also known as the halogens.
- Group 18 elements are known as the noble gases.

Key
Atomic number
Symbol
Element name

Alkali metals																		Noble gases																	
1																		2																	
H																		He																	
hydrogen																		helium																	
3																		4																	
Li																		Be																	
lithium																		beryllium																	
11																		12																	
Na																		Mg																	
sodium																		magnesium																	
19																		20																	
K																		Ca																	
potassium																		calcium																	
37																		38																	
Rb																		Sr																	
rubidium																		strontium																	
55																		56																	
Cs																		Ba																	
caesium																		barium																	
87																		88																	
Fr																		Ra																	
francium																		radium																	
57-71																		72-73																	
Lanthanoids																		Hf																	
La																		Ta																	
lanthanum																		tantalum																	
58																		59																	
Ce																		Pr																	
cerium																		praseodymium																	
60																		61																	
Nd																		Pm																	
neodymium																		promethium																	
62																		63																	
Sm																		Eu																	
samarium																		europium																	
64																		65																	
Gd																		Tb																	
gadolinium																		terbium																	
66																		67																	
Dy																		Ho																	
dysprosium																		holmium																	
68																		69																	
Er																		Tm																	
erbium																		thulium																	
70																		71																	
Yb																		Lu																	
ytterbium																		lutetium																	
89																		90																	
Ac																		Th																	
actinium																		thorium																	
91																		92																	
Pa																		U																	
protactinium																		uranium																	
93																		94																	
Np																		Pu																	
neptunium																		plutonium																	
95																		96																	
Am																		Cm																	
americium																		curium																	
97																		98																	
Bk																		Cf																	
berkelium																		californium																	
99																		100																	
Es																		Fm																	
einsteinium																		fermium																	
101																		102																	
Md																		No																	
mendelevium																		nobelium																	
103																		104																	
Lr																																			
lawrencium																																			

Rare earth metals
Lanthanoids
Actinoids

FIGURE 3.1.1 The modern periodic table

Each element in the same group has similar chemical properties. For example, each of the group 1 elements reacts with water to form a similar type of compound and hydrogen gas. The reaction with water becomes more violent as you go down the group. Lithium reacts vigorously, sodium reacts a little more violently, and rubidium and caesium explode violently. This is an example of a trend you can see in the groups. The **reactivity** of the elements in group 1 increases as you go down the group. The same trend applies to group 2 elements. The elements in both of these groups are classified as metals, so they all have similar physical properties. Metals and non-metals will be considered in more detail in section 3.3.

reactivity the likelihood of an element or substance undergoing a chemical reaction

Although all the elements in group 17 have similar chemical and physical properties, their reactivity decreases as you go down the group. So for group 17, the most reactive element is fluorine. This reversal of reactivity trend is due to the different way in which non-metals and metals react. The noble gases (elements in group 18) are named for their lack of chemical reactivity. They are so chemically stable that they do not naturally undergo chemical reactions.

In groups 13–16, properties change quite significantly going down the group. For example, in group 14, carbon is a non-metal whereas tin and lead at the bottom of this group are metals; so the physical and chemical properties vary.

The periods (rows) are numbered 1–7, with period 1 containing only two elements: hydrogen and helium. The elements in a period display very different chemical reactivity. There is often a gradual change in properties across any one period. This gradual change allows us to identify trends; however, periods are less useful for working out the actual properties of individual elements. For example, one physical property that varies periodically is melting point, as shown for the period 3 elements in **Table 3.1.1**.

TABLE 3.1.1 The melting points of the period 3 elements

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Melting point (°C)	98	650	660	1410	44	113	–101	–189

The melting point increases to a maximum at silicon, and then decreases to a minimum at argon. For all periods, the melting point follows a similar series of maximum and minimum. While the minimum value always corresponds to the noble gas at the end of the period, the maximum varies because of the presence of the transition metals in periods below period 3.

LEARNING CHECK 3.1

DESCRIBING

- 1 **Describe** how Mendeleev arranged the elements on the periodic table.
- 2 **Identify** the groups that have been given common names, such as the noble gases. Make a table showing the group numbers and the common names of these groups.
- 3 **Explain** why organising the elements into groups is useful.

APPLYING

- 4 **Identify** two elements you would expect to have similar properties to chlorine.
- 5 If beryllium reacts with fluorine gas to make beryllium fluoride, predict what would happen if you placed some magnesium metal in a gas jar containing chlorine gas. **Explain** your answer.
- 6 Argon does not react readily with any other element. (It is called an inert gas.) **Identify** which other elements could be called inert gases.
- 7 Silicon is a well-known semiconductor in electronics. **Identify** another element that is also likely to display semiconductor properties.

3.2 Trends in the periodic table

The periodic table is a powerful tool because it reveals specific patterns in the properties of the elements. This means that scientists can gain a lot of information about an individual element and how it behaves simply from its position in the periodic table.

In Chapter 1, you learnt about the relationship between the organisation of periodic table and the electron configuration of atoms of elements. This knowledge can be used to explain the patterns identified in the periodic table.

Stable electron configurations

As you learnt in Chapter 1, elements in the same group have the same number of electrons in the same outermost subshell.

As you know, noble gases have either completely filled energy levels, as seen for He and Ne, or levels filled with eight electrons (Ar, Kr, Xe and Rn). Since the noble gases generally do not undergo chemical reactions, we are able to deduce that filled outer energy levels or filled outer p subshells are the most stable electron configurations.

A generalisation that can be made is that elements in the same group of the periodic table tend to gain or lose the same number of electrons to achieve the stable electron configuration of the nearest noble gas.

Atomic radius

Atomic radius is a measure of the size of an atom, defined as the distance from the nucleus to the boundary of the cloud of electrons surrounding it (Figure 3.2.1). The boundary is hard to measure because it is not a well-defined physical entity.

Instead, an estimate of atomic radius can be made by taking half the experimentally measured distance between the nuclei of two neighbouring atoms of the same element; for example, metal atoms in a metallic solid.

The trends in atomic radii observed with respect to the periodic table remain consistent regardless of which measurement is used:

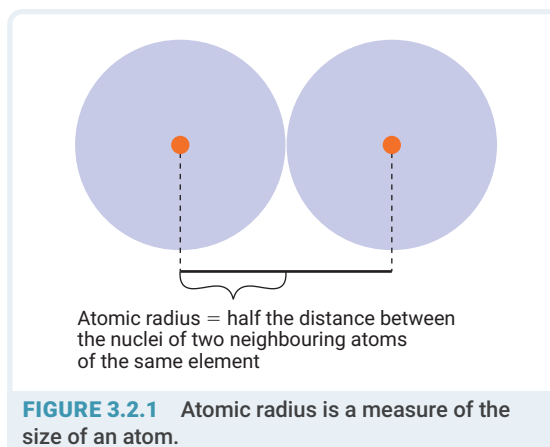
- Across the periodic table, the nucleus contains more protons and as such the positive charge of the nucleus increases. As the nucleus becomes more positive, the electrons in the outer energy level are more strongly attracted to the positive nucleus and they move closer together, making the atom smaller. As a result, atomic radius decreases from left to right across any period of the table.
- As you move down a group, the elements at the *bottom* of a group have more energy levels filled than those at the *top* of a group. As more energy levels are filled, down the group, outer electrons get further away from the nucleus and the atomic radius increases.

These trends are illustrated in Figure 3.2.2.

atomic radius a measure of the size of an atom – half the distance between the nuclei of two neighbouring atoms of the same element – the distances are very small and are measured in picometres (pm) ($1 \text{ pm} = 1 \times 10^{-12} \text{ m}$)



FORMULA AND
DATA BOOK



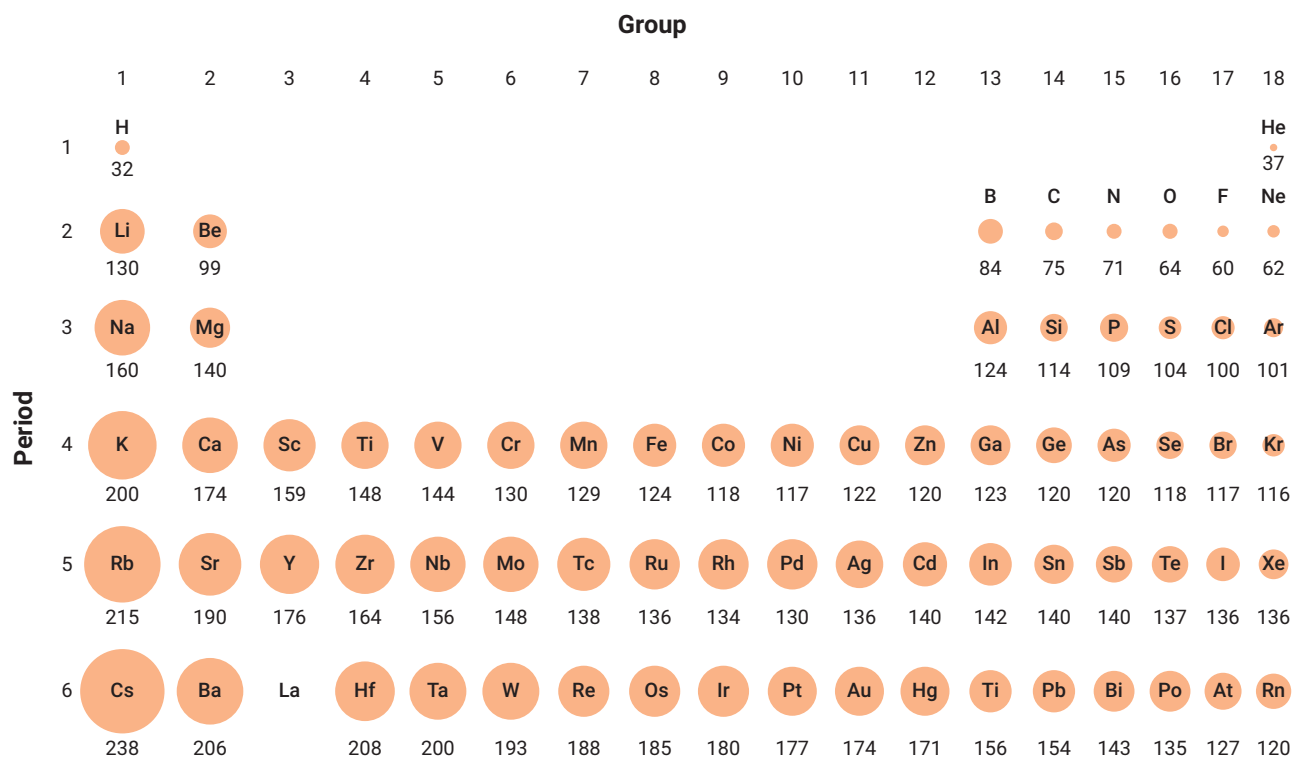


FIGURE 3.2.2 Trends in atomic radii (picometres) in the periodic table: atomic radius decreases across a period and increases down a group. Note that no data is available for the calculated atomic radius of lanthanum.

Valency

valence shell the outermost electron energy level of an atom

valency a measure of the number of chemical bonds an atom can form with other atoms

The number of electrons in the **valence shell** (valence electrons) of the elements in groups 1, 2 and 13–18 can be determined from the periodic table and from their electron configurations. The number of valence electrons also determines the **valency** of each element. Valency is an atom's ability to form chemical bonds with another atom. It is a number referring to the combining power of an element when compounds form. **Table 3.2.1** shows the relationship between the group number, number of electrons and valency. Once you know the valency, you can use the number of bonds an atom can form to determine a ratio for combining different elements.

TABLE 3.2.1 How valency is related to groups in the periodic table

Group number	1	2	3–12	13	14	15	16	17	18
Number of valence electrons	1	2	Variable	3	4	5	6	7	8
Valency	1	2	Variable	3	4	3	2	1	0

Calculating the valency of atoms

Some simple rules for calculating the valency of atoms are as follows:

1. Refer to the periodic table and determine the number of valence electrons for the atom.
2. If the number of valence electrons is four or fewer, the valence is equal to the number of electrons. For example, magnesium has two valence electrons, so magnesium has a valency of 2. Carbon has four valence electrons, so it has a valency of 4.

3. If the number of valence electrons is greater than four, then calculate the valency by subtracting the number of valence electrons from eight. For example, nitrogen has five valence electrons, so it has a valency of 3 ($8 - 5$). Chlorine has seven valence electrons so it has a valency of 1 ($8 - 7$).

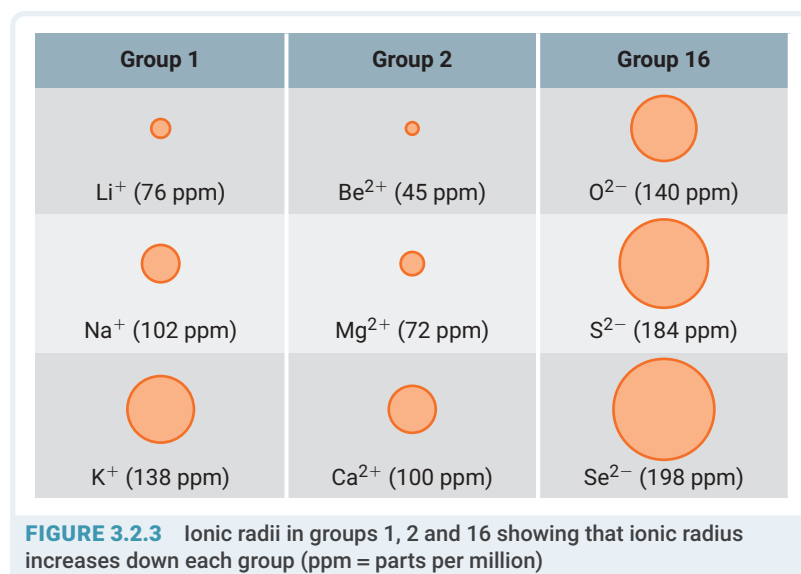
Valency can also be determined from the electron configuration of an element. As mentioned earlier, group 18 elements, (the noble gases) have stable electron configurations, so elements will lose or gain electrons to attain noble gas electron configuration. Therefore, the valency is equal to the number of electrons needed to be lost or gained for an atom to have a complete energy level, i.e. the electron configuration of the nearest noble gas.

For example, the alkali metals (group 1) have similar chemical properties and all have electron configurations with one more electron in the next energy level than the previous noble gas, i.e. s^1 . When alkali metals react, they tend to lose one electron to obtain the stable electron configuration of the nearest noble gas. They have a valency of 1. Similarly, the alkaline earth metals (group 2) have similar chemical properties as well as having two more electrons (s^2) than the nearest noble gas, so they tend to react to lose two electrons. They have a valency of 2.

The group 17 halogens all have similar chemical properties and one less electron (p^5) than the nearest noble gas, so they tend to gain one electron to obtain noble gas configuration. They have a valency of 1.

Ionic radius

To form ions, some atoms gain electrons, whereas other atoms lose electrons. When an atom gains or loses an electron, there is an imbalance of positively charged protons and negatively charged electrons. As such, the attractive forces between the nucleus and the electrons change, so the ionic radius is different from the atomic radius. Atoms in different groups form ions in different ways, so there is no clearly observable trend in the way that ionic radius changes across a period. There is a pattern, though, in that elements on the left side of the table (metals) have ionic radii smaller than their atomic radius after losing valence electrons, while elements on the right of the table (non-metals) generally have ionic radii bigger than their atomic radius after adding electrons to their valence shell. However, within every group, the **ionic radius** increases down the group, as the number of the energy level being filled increases (Figure 3.2.3).



ionic radius the distance between the centre of the nucleus and the outermost electron shell in an ion

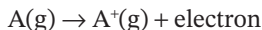
First ionisation energy

Since electrons are attracted to the positive charge in a nucleus, energy is required to remove them when an atom becomes a positive ion. The **first ionisation energy** is the amount of energy needed to completely remove an electron from a neutral atom when it is a gas. An atom that has a low first ionisation energy becomes an ion (a charged atom) very easily.

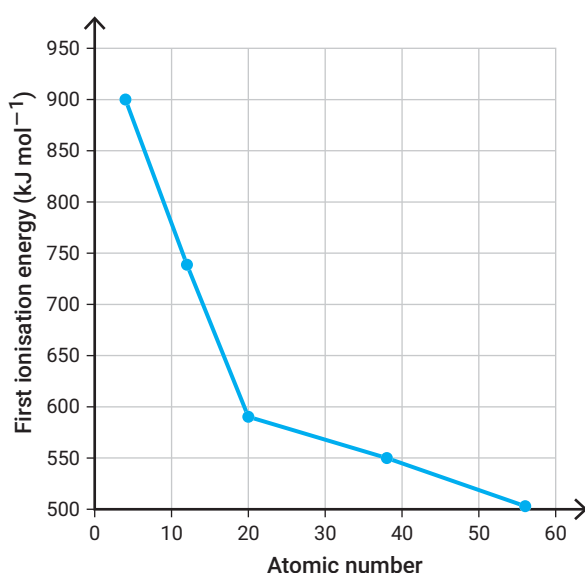
first ionisation energy the energy required to remove a single electron from an atom in the gaseous state



This process can be represented by the equation shown below, where the highest-energy electron is removed from a gaseous atom (A), to form a positive ion (A^+):



First ionisation energy decreases down a group. As we move down a group, the electrons in their successive energy levels are getting further away from the nucleus and are less strongly bound to the positive nucleus. This means it takes less energy to remove an electron from an outer energy level because there is a weaker attraction between the positive nucleus and the outermost negative electrons. The 'shielding effect', in which inner electrons 'shield' outer electrons from the attraction of the nucleus (lowering their ionisation energy in the process), also contributes to the weaker attraction of the outer electrons to the nucleus. This trend is observed among group 2 elements and is shown in **Figure 3.2.4**.



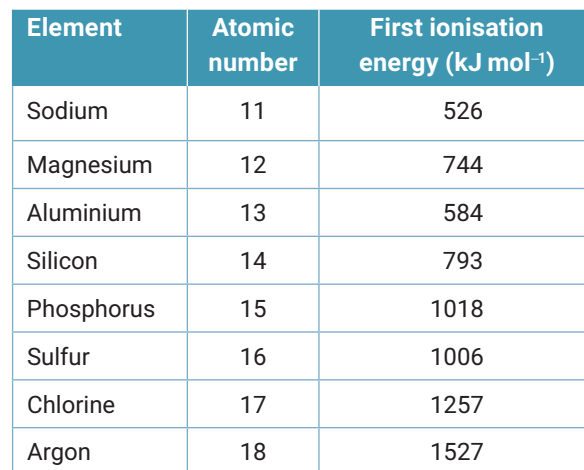
Element	Proton number	First ionisation energy (kJ mol^{-1})
Beryllium	4	906
Magnesium	12	744
Calcium	20	596
Strontium	38	556
Barium	56	509

FIGURE 3.2.4 The first ionisation energies of the group 2 elements

First ionisation energy increases from left to right across a period. Within a period, the valence electrons are in the same energy level so are all approximately the same distance from the positively charged nucleus. However, across a period, the number of protons (atomic number) increases, so the positive charge of the nucleus also increases. The attraction between the positive nucleus and the outermost negative electrons (in the same energy level) becomes stronger. This makes it harder to remove an outer electron and requires more energy, so the first ionisation energy increases across a period. **Figure 3.2.5** shows this trend for period 3. For example, the first ionisation energy is greater for magnesium than for sodium. Likewise, the first ionisation energy is greater for argon than for chlorine. The trend does display some minor exceptions. There are two relatively small decreases in ionisation energy from magnesium to aluminium and from phosphorus to sulfur. The reason for these anomalies can be explained by considering electron configurations and the stability differences of full, half-full and partially filled subshells.

Explaining anomalies in ionisation energy

The data in **Figure 3.2.5** shows that more energy is required to remove an outer electron from magnesium and phosphorus than from the elements immediately following them. We know that as the attraction between the positive nucleus (increasing atomic number) and the outermost



negative electrons becomes stronger, it is harder to remove an outer electron from the same energy level. So there is obviously another factor making the electron configurations of these elements more stable.

Mg $1s^2$ $2s^2$ $2p^6$ $3s^2$

P $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^3$

Table 3.2.2 shows the first ionisation energies for the s and p block elements of periods 1–4. The noble gases have the highest ionisation energy in each period, thus indicating that it is difficult to remove an electron from these elements. This confirms that noble gas electron configurations are extremely stable.



H 1318																	He 2379	
Li 526	Be 906												B 807	C 1093	N 1407	O 1320	Fl 1687	Ne 2087
Na 502	Mg 744												Al 584	Si 793	P 1018	S 1006	Cl 1257	Ar 1527
K 425	Ca 596											Ga 585	Ge 768	As 953	Se 947	Br 1146	Kr 1357	



Worksheet
Atomic radius and
ionisation energy

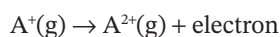
successive ionisation energies the energies required to remove the electrons from an atom, in sequence, starting with the outermost

For each period, the group 2 element, with its filled 2s subshell, has a higher ionisation energy than the adjacent p block element, and the group 15 element, with its half-filled p subshell, has a higher ionisation energy than the adjacent group 16 element.

Therefore, this data confirms that for the s and p block elements, filled and half-filled subshells are more stable than partially filled subshells. Note that d block elements have not been discussed because, as mentioned in Chapter 1, some of these elements have electron configurations that are exceptions in their electron configuration.

Successive ionisation energies

A second electron can be removed after the first, then a third can be removed and so on. The energy required to remove a second electron and form A^{2+} is known as the second ionisation energy:



If this trend were to continue, then successive electrons could be removed from the atom, giving rise to the term **successive ionisation energies**.

As previously discussed, the magnitude of the ionisation energy depends on the extent of the attraction between the nucleus and the electron being removed. With each additional electron that is removed, the ion becomes more positively charged and the attraction between the nucleus and the remaining electrons increases. This means that successive ionisation energies keep increasing.

Importantly, the size of the increase between successive ionisation energies provides crucial information about an atom's electron configuration and the position of the element on the periodic table. **Figure 3.2.6** is a graph of the successive ionisation energies for an atom of sodium.

Sodium has 11 electrons, and the configuration $1s^2 2s^2 2p^6 3s^1$. Figure 3.2.6 shows that there are particularly large increases in ionisation energy from the first to the second ionisation energy and from the ninth to the tenth ionisation energy. This corresponds to the change in electron energy levels. When an electron is removed from a new energy level closer to the nucleus, the distance between the remaining electrons and the nucleus is reduced and the force of attraction between the electrons and nucleus is significantly greater. Accordingly, the increase in ionisation energy is larger than it would be if an additional electron were removed from the same energy level.

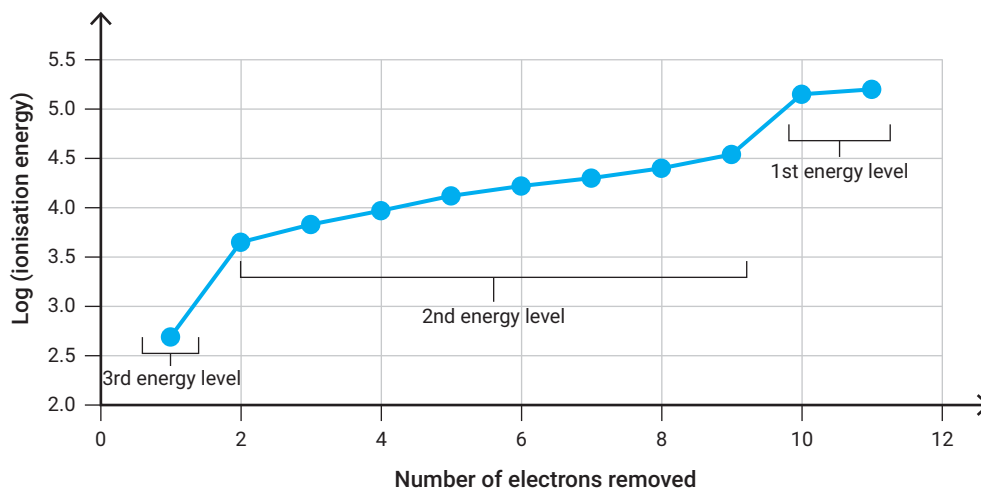


FIGURE 3.2.6 Successive ionisation energies of sodium

Table 3.2.3 shows the first four ionisation energies of some of the group 2 elements. The data indicates that it is easy to remove the first two electrons, then there is a large increase in the energy required to remove the third electron, followed by a smaller increase in energy required to remove

the fourth electron. All these elements have an electron configuration with a filled outer energy level s subshell (s^2). After the outer two s subshell electrons have been removed, the resultant 2^+ ion has the same electron configuration as that of the nearest noble gas. This electron configuration is extremely stable, making the removal of the third (and subsequent) electron more difficult.

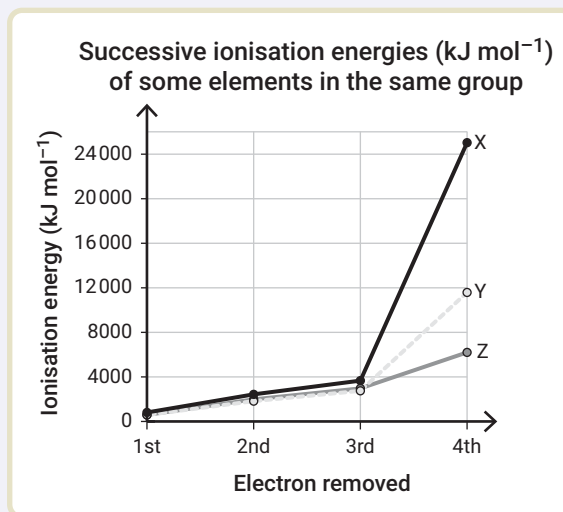
TABLE 3.2.3 Successive ionisation energies (kJ mol^{-1}) of some group 2 elements

Element	Ionisation energy (kJ mol^{-1})			
	First	Second	Third	Fourth
Be	906	1763	14 855	21 013
Mg	744	1457	7739	10 547
Ca	596	1152	4918	6480
Sr	556	1071	4210	5500

When studying a graph or table of an atom's successive ionisation energies, you can quickly identify the electron configuration by counting the number of electrons removed between each large increase in ionisation energy.

WORKED EXAMPLE 3.2.1

The graph below shows the first four ionisation energies for elements X, Y and Z, which are in the same group of the periodic table.



- Identify which group of the periodic table these elements belong to.
- List the elements in order of increasing period number.
- Justify your decisions.

ANSWERS

- 1 Cross reference the data from the graph with Table 3.2.2.
 X has an ionisation energy similar to Be.
 Y has an ionisation energy similar to Al.
 Z has an ionisation energy similar to Ga.

2 Determine the group in which these elements are found.

The elements are in group 13 of the periodic table.

b 1 Determine the location of each element within the group.

X occurs earliest, followed by Y and Z.

2 Identify the order of increasing period number.

X, Y, Z

c 1 Use the information from parts a and b to form a justification.

For each element, only a small increase in energy is required to successively remove the first three electrons. A much larger amount of energy is required to remove the fourth electron. This implies that after three electrons are removed, the elements have a stable electron configuration with a filled energy level. Group 13 elements have three electrons in their outer energy level and attain noble gas configuration with the removal of these three electrons. Hence, the elements must be in group 13.

Electrons in lower energy levels are more difficult to remove because they are more strongly attracted to the nucleus. Therefore, the electrons of element X must be in a lower energy level than those in element Y, and X is in a lower period. Similarly, element Y must be in a lower period than element Z. Therefore, the elements listed in order of increasing period number are X, Y, Z.

Electronegativity

electronegativity the ability of an atom to attract electrons, while being bonded to another atom

Electronegativity is an important concept that has implications for chemical bonding and the properties of different compounds. Electronegativity describes an atom's ability, when combined with another element, to attract electrons.

As the atomic number (Z) increases, the number of protons in the nucleus increases. Thus, the nucleus becomes increasingly positive across a period. Across a period, each electron in the same energy level experiences a stronger attraction to the nucleus. The atomic radius decreases due to the increased nuclear charge across a period. This results in the trend that electronegativity increases across a period (**Figure 3.2.7**).

As we know, going down a group, atoms have successively more electron energy levels and therefore increasing atomic radii. Electrons in the outer energy level are further away from the nucleus, which is consistent with a weaker attraction between the positive nucleus and the negative electrons for larger atoms. Increased shielding by the greater number of inner electrons



Worksheet

Ionisation energy and electronegativity

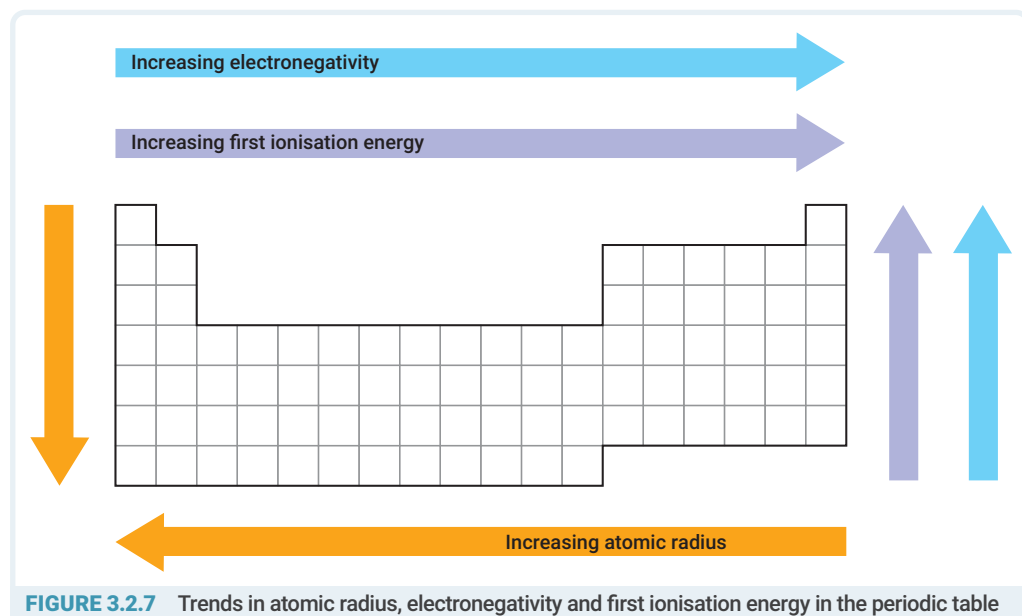


FIGURE 3.2.7 Trends in atomic radius, electronegativity and first ionisation energy in the periodic table

also reduces the attraction of the outer electrons to the nucleus. This means electronegativity decreases down a group.

Electronegativity can be measured using the Pauling scale of electronegativity. The Pauling scale of electronegativity is a relative scale that includes every element, and ranges from approximately 0.7 (francium) to 3.98 (fluorine). This information can be found in the *Formula and data book*.

The periodic table can be used to determine which of two elements is more electronegative. For example, oxygen is more electronegative than sulfur because it is above it in group 16. Chlorine is also more electronegative than sulfur because it is further to the right in period 3.

The most electronegative element is fluorine. The noble gases do not have electronegativity values. This is because their stable and full energy subshells mean that they do not form bonds with other atoms, so their attraction of shared electrons with a bonded atom cannot be measured.



LEARNING CHECK 3.2

DESCRIBING

- 1 **State** how many electrons are in the outer energy levels of:
 - a beryllium
 - b fluorine
 - c phosphorus
 - d sodium
 - e argon.
- 2 **Define** 'first ionisation energy'.

APPLYING

- 3 **Determine** the valency of:
 - a sodium
 - b silicon
 - c fluorine.
- 4 **Determine** whether lithium or fluorine has a bigger atomic radius.
- 5 **Explain** the relationship between valency and electron configuration.
- 6 In the following pairs of elements, **state** which element is more electronegative.
 - a Nitrogen and oxygen
 - b Magnesium and beryllium
 - c Phosphorus and fluorine
 - d Carbon and silicon
- 7 Four elements with atomic numbers 17, 18, 19 and 20 have first ionisation energies of 1257, 425, 596 and 1527 kJ mol⁻¹. **Determine** which ionisation energy belongs to which element.
- 8 **Explain** why atomic radius changes across a period and down a group.
- 9 Chlorine and oxygen are both strongly electronegative elements. On the Pauling scale of electronegativity, chlorine has an electronegativity value of 3.2, and oxygen a value of 3.4. **Suggest** why the value for oxygen is higher.

ANALYSING

- 10 Two elements have the following successive ionisation energies (in kJ mol⁻¹):

P	425	3058	4418	5883	7982
Q	1093	2359	4627	6229	37 838

 - a **Identify** which group on the periodic table each belongs to.
 - b **Justify** your decision.

3.3 Metallic and non-metallic character in the periodic table

Electron affinity

electron affinity the ability of an atom in the gaseous state to accept an electron and form a negative ion

Electron affinity is the ability of an atom in the gas state to accept an electron and form a negative ion. This property of atoms is different from electronegativity because it involves the full addition of an electron, not just the ability to attract electrons.

Moving across a period from left to right, atomic radius decreases as outer electrons and spaces for outer electrons to be added get closer to the positively charged nucleus. As a result, it is easier to add an electron to these atoms and so the electron affinity increases across a period.

Moving down a group, spaces for outer electrons to be added get further away from the nucleus as more energy levels are added. Due to decreased attraction, it becomes harder to add an electron and so the electron affinity decreases.

Properties of metals and non-metals

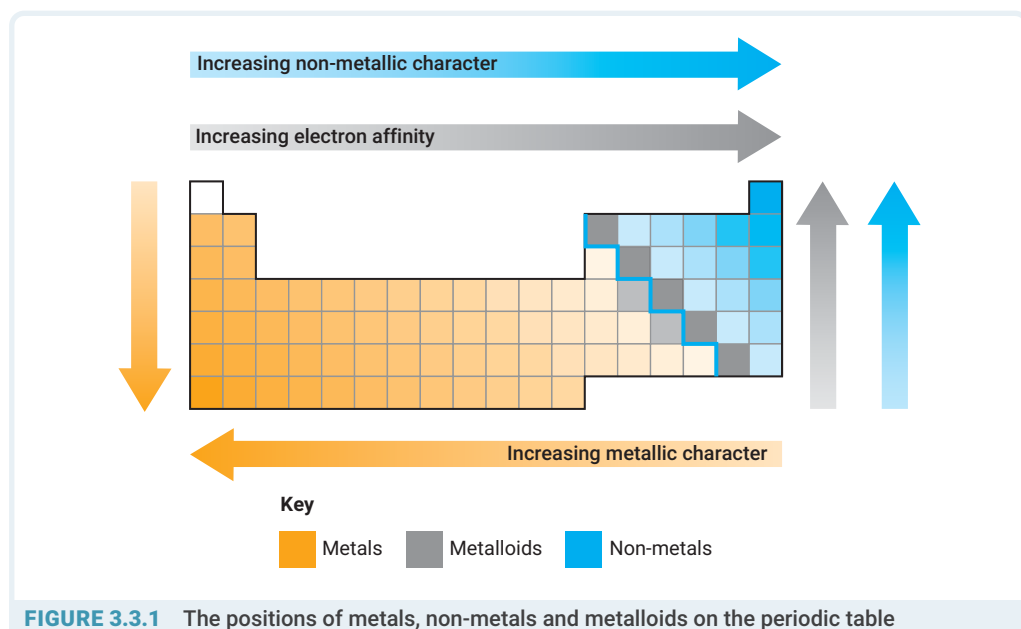
Elements can be classified as metals, non-metals or metalloids. Metals make up 75% of all elements. Metals are shiny in appearance, are good conductors of heat and electricity, and are **malleable** and **ductile**. Non-metals have the opposite properties of metals: they are dull in appearance, do not conduct electricity or heat, and are brittle. Metalloids are a small group of elements that have properties of both metals and non-metals.

Metallic properties are a result of elements behaving like metals and forming metallic bonds. These behaviours are due to the tendency of a metal atom to lose its outer electrons. Atoms that lose outer electrons easily are more metallic and atoms that do not lose outer electrons are less metallic. This is why atoms with higher electron affinity are less metallic and atoms with lower electron affinity are more metallic. These comparative trends are shown in Figure 3.3.1. The 'staircase' of the metalloids, positioned between the metals and the non-metals, is also evident.

As can be seen in **Figure 3.3.1**, there is a close relationship between non-metallic character of an element and electron affinity.

malleable can be repeatedly deformed, or hammered, without shattering

ductile can be drawn out to form a thin wire



Group 1 – the alkali metals

The elements in group 1 are known as the alkali metals. In nature, alkali metals occur in compounds, rather than in their elemental form, because they are very reactive and so form bonds with non-metals easily. They have a relatively low melting point compared with other metals and can be cut with a knife because they are so soft. These characteristics make them unsuitable for construction and fabrication, unlike metals such as iron and aluminium.

The alkali metals are so named because they all react with water to form a solution that is alkaline. This reaction can be quite violent, as in the case of potassium, which reacts spectacularly to form potassium hydroxide and hydrogen gas. The heat generated causes the hydrogen to burn vigorously, causing an explosion with a lilac flame. The reaction is shown in the following word equation:



Moving down group 1, the elements become more reactive. This can be explained by considering the trend for first ionisation energy. All alkali metals react to form a positive ion, by losing their one valence electron. The amount of energy required to remove this electron decreases down the group. This is because, with increasing period number, atoms have more energy levels and correspondingly increased shielding of the outer electrons from the nuclear attraction. This means it is easier it is to remove the valence electron. In summary, the further down the group the element occurs, the lower its first ionisation energy is, and the more its reactivity increases. This trend is true for all metallic groups.

Group 17 – the halogens

The trends and patterns of the elements in group 17, the halogens, provide a good contrast to those in group 1. Each halogen has a distinctive colour. All are non-metals, so they are poor conductors of heat and electricity. The halogens are toxic to life, which in turn makes them useful. For example, chlorine is used to treat swimming pool water and iodine is used as an antiseptic to help prevent wound infections. Importantly, the halogens have different characteristics when bonded with atoms from other groups. For example, halogens bonding with alkali metal atoms form alkali halides, such as sodium fluoride (NaF) and sodium chloride (NaCl). Alkali halides are very different from the halogens because the halogen is now a negative ion. Sodium fluoride is effective in preventing dental decay and sodium chloride is common ‘table’ salt.

The halogens occur naturally in the form of molecules (e.g. F_2 , Cl_2 , Br_2) in which two atoms are bonded together (diatomic molecules) rather than existing as single free atoms. When discussing the properties of halogens, scientists typically refer to their naturally occurring diatomic forms (similarly for hydrogen (H_2), oxygen (O_2) and nitrogen (N_2)). The melting and boiling points of the halogens increase down the group: fluorine and chlorine are gases at room temperature, bromine is a liquid and iodine is a solid. Iodine has an unusual physical property: when it is heated, **sublimation** occurs, rapidly changing it from a grey solid into a purple gas.

Fluorine, the first element in group 17, is the lightest halogen and has the highest electronegativity value of all elements, making it extremely reactive. Almost all other elements form compounds with fluorine. Due to its reactivity, even the naturally occurring form of fluorine (F_2) is extremely difficult to manage in a laboratory and can be stored only in containers made from extremely stable, non-reactive materials such as platinum.

All halogens have a high first ionisation energy, so they do not lose electrons when reacting. Instead, they try to gain a single electron to fill up their outer energy level and thus attain noble gas electron configuration, and form a negatively charged ion. The ability to attract an electron is closely linked to atomic radius and electron affinity.

sublimation the process of a solid turning into a gas when heated, without forming a liquid



FIGURE 3.3.2 Properties of the halogens: all three physical states of the halogens: gas (chlorine), liquid (bromine) and solid (iodine)

Science Photo Library/Andrew Lambert Photography

Fluorine is a small atom, which means that the distance between the positively charged nucleus at the centre of the atom and external electrons is relatively small, so the attraction to the nucleus is strong. Moving down the group, additional internal electrons shield the valence shell from nuclear attraction, so it is felt less strongly and attraction to the nucleus lessens, reducing electron affinity.

PRACTICAL ACTIVITY 3.3.1

REACTIVITY OF THE GROUP 2 METALS – THE ALKALINE EARTH METALS

Introduction

Metals in group 2 of the periodic table are less reactive than those in group 1. When added to hydrochloric acid, the metals react to form the metal chloride, together with hydrogen gas.

Research question

How does the reactivity of group 2 metals differ from each other in hydrochloric acid?

Aim

To demonstrate the relative reactivity of elements within the group 2 metals

Materials per group

- 2 small pieces of magnesium ribbon (highly flammable)
- small piece of calcium (flammable) (Do not use old calcium. Use fresh stock.)
- 50 mL of 1 mol L⁻¹ hydrochloric acid (HCl) (irritant)
- 3 test tubes and a test-tube rack
- splint
- matches



What are the risks in doing this activity?	How can these risks be managed?
	Wear eye protection.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your risk assessment before you proceed.

Procedure

- 1 Fill two test tubes one-quarter full with dilute hydrochloric acid.
- 2 Into one test tube, drop a small piece of magnesium.
- 3 Into the other test tube, drop a small piece of calcium.
- 4 Compare the reactivity of the two metals.
- 5 Drop another piece of magnesium into the first test tube and invert a clean dry test tube over the top to capture the gas produced.
- 6 When the bubbling has stopped, keep the test tube inverted and test the gas by holding a lighted splint to the mouth of the inverted test tube.
- 7 Record what happens.

Results

Record your observations.

Analysis of results

- 1 Compare the reactivities of magnesium and calcium with hydrochloric acid.

Interpretation

- 2 Explain whether the results are consistent with the trends that are observed in the group 2 elements.

LEARNING CHECK 3.3

DESCRIBING

- 1 **Define** 'metalloid'.
- 2 **Describe** the metallic character trend in the periodic table.
- 3 **State** why alkali metals are so named.
- 4 **List** three reasons why chlorine is regarded as a non-metal.
- 5 **Explain** why you expect the reaction of sodium with water would be more, or less, vigorous than that of potassium with water.

APPLYING

- 6 **Identify** why the alkali metals all form an ion with a charge of +1.
- 7 **Distinguish** between electron affinity and electronegativity.
- 8 **Explain** why fluorine is the most electronegative element.
- 9 **Explain** the relationship between first ionisation energy and reactivity in groups 1 and 17.
- 10 **Compare** the reactivity of magnesium in group 2 to that of sodium in group 1.
- 11 Element number 117 was discovered in 2009 and named 'tennessine'. It is placed below astatine, at the bottom of group 17. So far, only a few atoms of tennessine have been obtained. If it were possible to collect a larger sample and investigate it, **predict** the properties it might have.

3.4 Properties of the oxides of the elements in period 3

In examining the metallic character of the elements, you have seen examples of the trends that occur moving across the periodic table. Further trends can be observed by considering the properties of the **oxides** of the elements in period 3.

The formulas of the oxides of the elements vary according to the valency of the different elements in period 3. These can be seen in **Table 3.4.1**. Argon (not shown), together with the other group 18 elements, does not react with oxygen, so it does not form an oxide. Oxygen, which has the electron configuration $1s^2 2s^2 2p^4$, has a valency of 2.

oxide a compound that forms when an element combines with oxygen

TABLE 3.4.1 Formulas of the period 3 oxides

Element	Valency	Oxide formula	Melting point (°C)
Na	1	Na ₂ O	1132
Mg	2	MgO	2852
Al	3	Al ₂ O ₃	2054
Si	4	SiO ₂	1713
P	3	P ₄ O ₆	24
S	2	SO ₂	-75
Cl	1	Cl ₂ O	-20

alkali/base a substance that reacts with acidic substances to form water as one of the products of the reaction

amphoteric a substance that reacts with both alkaline and acidic solutions, so it is classified as neither alkali nor acid

acid a substance that produces hydrogen ions (H^+) in aqueous solution

Table 3.4.1 demonstrates that the melting points vary for these compounds. This means the properties of the compounds are quite different.

While acids and bases will be discussed in detail in Chapter 17, there are some interesting periodic trends relevant to the acidic and basic properties of the oxides that elements form across a period.

Using period 3 as an example, sodium oxide and magnesium oxide are known as alkaline or basic oxides, because as **alkalis/bases** they react with acids to produce water. Both these compounds are soluble in water to form alkaline solutions. Aluminium oxide is insoluble in water; however, it reacts with both **acid** and alkaline solutions and is classed accordingly as an **amphoteric** oxide. The oxides of the non-metals, silicon, phosphorus, sulfur and chlorine are all classified as acidic oxides. Although silicon dioxide does not dissolve in water, it does react when added to an alkaline solution. Phosphorus, sulfur and chlorine oxides are all acidic oxides because they form an acidic solution in water and react with alkaline solutions to form water as one of the products of the reaction.

Moving from left to right across the periodic table, the pattern is easy to observe: the oxides of the elements become less alkaline and more acidic.

LEARNING CHECK 3.4

DESCRIBING

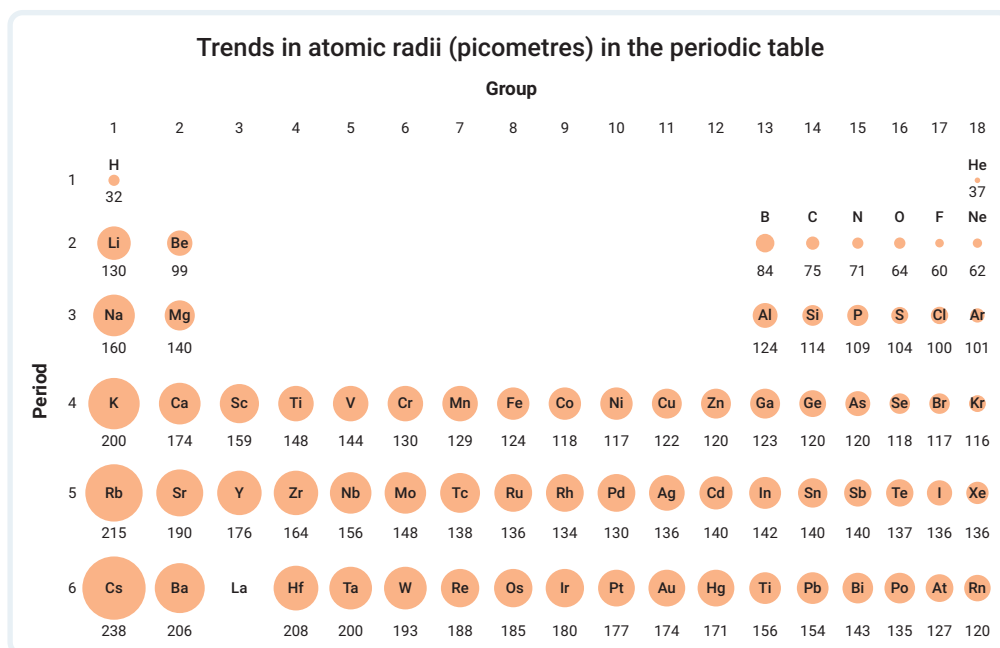
- 1 **Define** 'amphoteric oxide'.
- 2 **Identify** the trend that is observed across period 3 in terms of the acidity of the oxides.
- 3 **Discuss** why magnesium oxide is regarded as alkaline.

APPLYING

- 4 **Classify** the following oxides as acidic, alkaline or amphoteric.
 - a Potassium oxide
 - b Bromine oxide
 - c Boron oxide
- 5 Period 4 has 18 elements, whereas period 3 has only eight. **Compare** the trend in period 4 oxides to that in period 3. **Determine** whether there is a greater or smaller proportion of alkaline oxides.

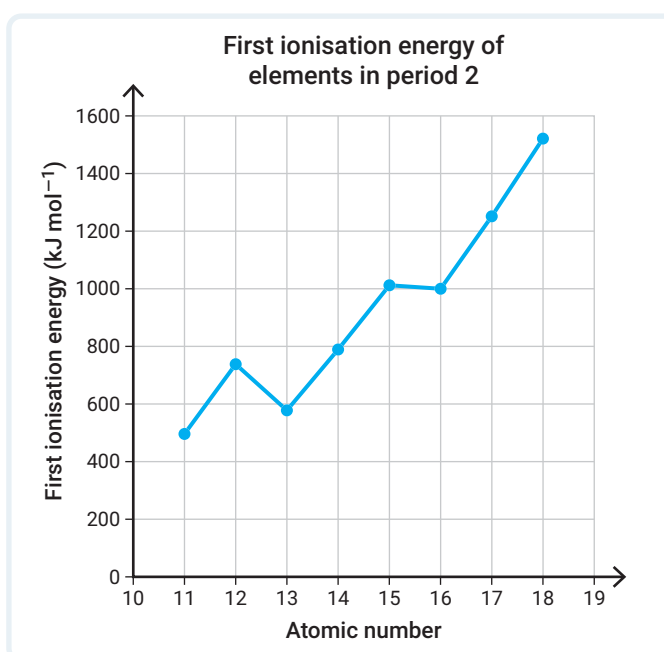
Trends in atomic radius

- Analysis of data for atomic radii, valencies, ionic radii, first ionisation energy and electronegativities can be used to determine trends, patterns and relationships.
- The atomic radius of elements decreases across a period (left to right) and increases down a group.
- The structure of the periodic table in groups and rows summarises relationships in chemical and physical properties of elements and connects these to electron configurations of elements.



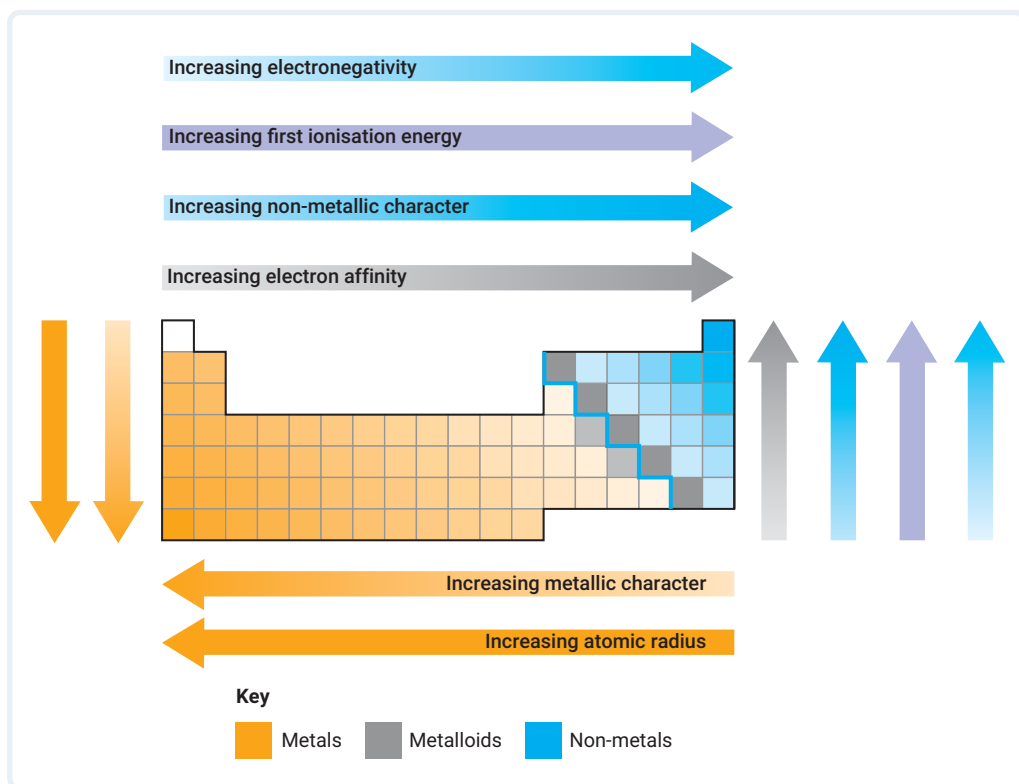
Trends in ionisation energy

- First ionisation energy is the amount of energy required to remove an electron from a neutral atom when it is a gas.
- First ionisation energy increases across a period (left to right) and decreases down a group.
- The energy required to remove additional electrons (after the first electron) increases and can be explained by the electron configuration.



Trends in electronegativity and metallic character

- Electronegativity is an atom's ability to attract an electron to itself.
- Electronegativity increases across a period (left to right) and up a group (from bottom to top).
- The reactivity of metals increases as you move down a group, whereas the reactivity of non-metals decreases as you move down a group.



Period 3 oxides

- An oxide is formed when an element reacts with oxygen; the chemical nature of oxides changes from basic (alkaline) to amphoteric to acidic across a period.

Formulas of the period 3 oxides

Element	Valency	Oxide formula	Melting point (°C)
Na	1	Na ₂ O	1132
Mg	2	MgO	2852
Al	3	Al ₂ O ₃	2054
Si	4	SiO ₂	1713
P	3	P ₄ O ₆	24
S	2	SO ₂	-75
Cl	1	Cl ₂ O	-20

MULTIPLE CHOICE

- The atoms of a certain neutral element contain 12 electrons. Another element with similar properties has atoms containing:
 - 11 electrons.
 - 17 electrons.
 - 20 electrons.
 - 37 electrons.
- Which of the following statements is correct for all metals?
 - All metals have high melting points.
 - All metals are low in reactivity.
 - All metals are hard and strong.
 - All metals conduct electricity.
- Group 13 consists of a series of metallic elements. It is likely that the electronegativity of the group 13 elements:
 - is at its highest for the period 4 elements.
 - is greater than for the group 1 elements.
 - decreases up the group.
 - does not change within the group.
- The element germanium has atomic number 32. It is in group 14 and period 4 of the periodic table. Which combination of properties from the table below is germanium likely to have?

	Conductivity of solid germanium	Nature of germanium oxide
A	High	Alkaline
B	None	Amphoteric
C	Low	Acid
D	Low	Amphoteric

- Which of the following shows the correct trends in properties, from left to right across the periodic table?

	First ionisation energy	Atomic radius	Electronegativity
A	Decreases	Increases	Decreases
B	Increases	Increases	Increases
C	Increases	Decreases	Increases
D	Decreases	Decreases	Decreases

[illegible]

-
- A graph showing the ionisation energy of a Group 2 element as the charge on the ion increases. The y-axis is labeled 'Ionisation energy' and the x-axis is labeled 'Charge on ion' with values 1, 2, 3, 4, and 5. Two curves are shown: a solid line labeled 'X' and a dashed line labeled 'Y'. Both curves show a step-wise increase in ionisation energy at each integer charge. The solid line 'X' starts at a low value for charge 1, increases at charge 2, and then has a very large jump at charge 3 to a high plateau. The dashed line 'Y' starts at a slightly higher value than 'X' for charge 1, increases at charge 2, and then has a smaller jump at charge 3, reaching a plateau below that of 'X'.

SHORT RESPONSE

11. **Explain** why the first ionisation energy increases as you move across the periodic table.
12. The following table shows the relationship between some properties. Copy and complete the table by **identifying** whether the quantity in the left column is greater than (>), equal to (=) or less than (<) the quantity in the right column. Indicate your responses by placing a sign (>, = or <) in each row of the central column.

Radius of the K atom		Radius of the K ⁺ ion
Number of valence electrons in a Ca atom		Number of valence electrons in a Ba atom
Radius of an oxygen atom		Radius of a sulfur atom

CROSS-CHAPTER QUESTION

13. **Compare** the reactivity of the alkali metals and halogens considering electron configuration.

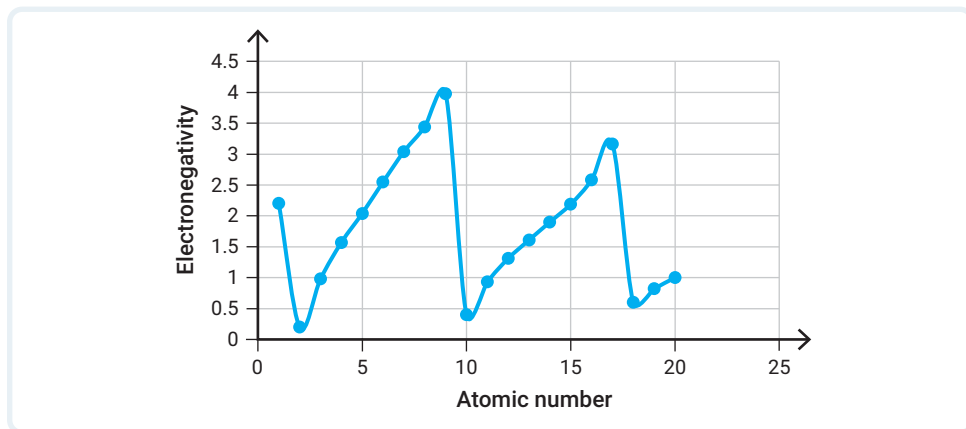
DATA ANALYSIS

14. The melting points of period 2 elements are given in the table below.

Element	Melting point (°C)
Li	180
Be	1277
B	?
C	3727
N	-210
O	-218
F	-219
Ne	-248

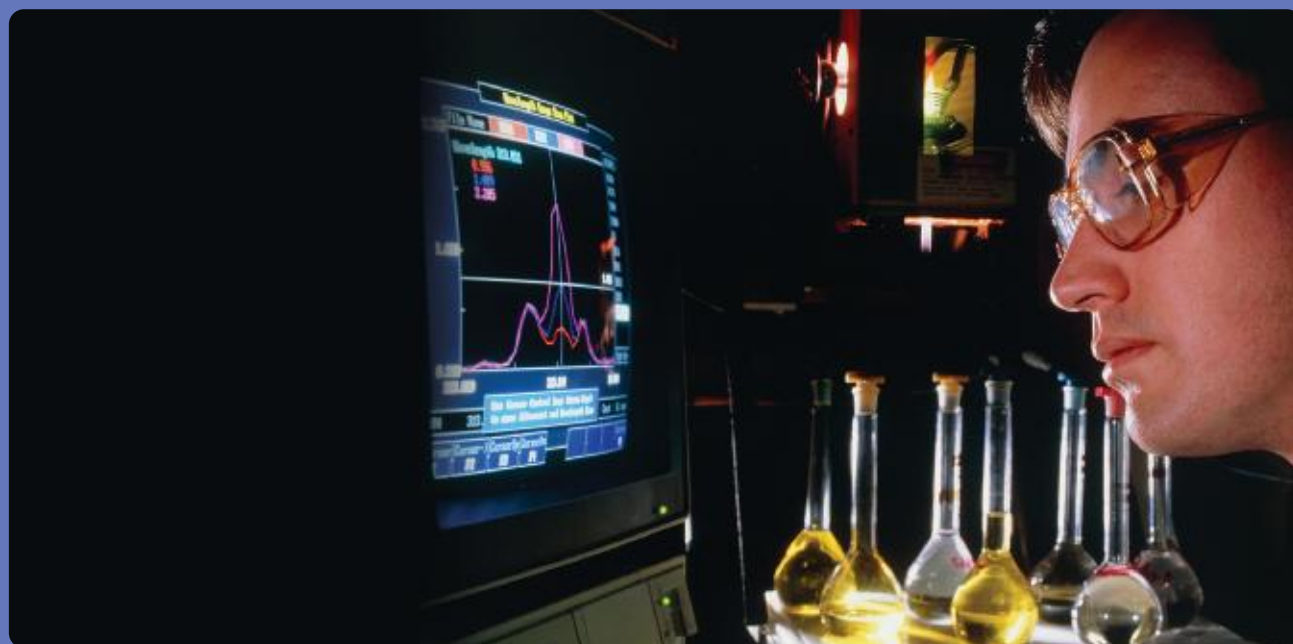
- a Plot a graph of melting point vs atomic number using the data provided.
- b i Use the graph to **predict** the melting point of boron.
ii **State** any assumptions you made in determining part i.
- c **Compare** the data in this table with the data in Table 3.1.1 and generalise trends in melting points for periods 2 and 3.

15. The graph below shows how electronegativity changes with increasing atomic number for the first 20 elements.



- Identify** the element with the highest electronegativity.
- Explain** why this element has the highest electronegativity value.
- Use data from the graph to **identify** the trend in electronegativity that occurs down a group of the periodic table and **explain** why this trend occurs.
- There are three elements on the graph that have very low electronegativity values. If the graph were extended to elements of higher atomic number, **predict** at what atomic number the next low point would occur? **Explain** your thinking.

Analytical techniques



Geoff Tompkinson/Science Photo Library

SYLLABUS DOT POINTS

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 INQUIRY

Investigate:

- flame tests to identify elements
- mass spectra and isotopes*
- atomic absorption spectroscopy (AAS) and the concentration of aqueous metallic ions.*

*Note: Simulations may be used.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

A variety of analytical instrumentation helps scientists to determine what kind of elements or compounds are present in a solid, liquid or gas; how much of each element or compound is present; and the physical and chemical structure of the material. Analytical instruments need to have high sensitivity, low limits of detection, good selectivity, good time resolution, accuracy and reproducibility. The analytical data provided by such instruments, which are often field-based devices, provides scientists with valuable information in areas such as atmospheric and oceanographic testing; environmental and production analysis for agriculture; mining; metallurgy; chemical analysis for pharmaceutical and scientific industries; and forensic investigation.

Practicals

- Flame tests
- Constructing a calibration curve to measure the concentration of copper(II) sulfate

Worksheets

- Analysing mass spectra
- Emission spectra

 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ An ion is formed when an atom loses or gains electrons.
- ✓ Isotopes are atoms of the same element that have different numbers of neutrons and therefore different mass numbers.
- ✓ The relative atomic mass of an element is the average mass of one atom of an element, relative to the mass of carbon-12.
- ✓ The average mass of an atom of an element is calculated from the mass and abundance of isotopes of that element.
- ✓ Electrons exist in discrete energy levels.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ identify that mass spectrometry is a quantitative analysis technique
- ✓ identify that mass spectrometry involves the formation of positive ions, deflection of the ions according to their mass, and detection of the separated ion streams
- ✓ analyse mass spectra to identify isotopes and their percentage composition
- ✓ calculate relative atomic mass from mass spectrum data
- ✓ identify and compare absorption and emission line spectra
- ✓ conduct flame tests and interpret the results
- ✓ explain what the lines on absorption and emission spectra represent
- ✓ compare flame tests to emission spectra
- ✓ describe the Bohr model of an atom and explain why the Bohr model was accepted
- ✓ analyse absorption and emission line spectra to identify unknown elements
- ✓ analyse atomic absorption spectroscopy data to determine the concentration of metallic ions in solution.

4.1 Mass spectrometry

mass spectrometry
a technique used to determine the atomic or molecular composition of a sample and its isotopic composition

isotopic composition
a list of the different isotopes present in a sample of an element and the relative amount of each isotope that is present

Mass spectrometry is a technique that can tell chemists a lot about the structure of a substance such as the specific identity and amount of atoms or molecules in a sample of material. In particular, when analysing a pure element, mass spectrometry can be used to determine **isotopic composition**. An isotopic composition tells you which isotopes are present in an element and the percentage of each isotope present. An advantage of mass spectrometry is that it only requires a small amount of material but can provide a wealth of information about that material. Mass spectrometry has many uses from radioactively dating fossils to detecting drugs in sports supplements. Mass spectrometers are also used in many locations, including satellites where they can be used to identify particles carried by the solar wind.

Mass spectrometry is carried out using a mass spectrometer (**Figure 4.1.1**). During mass spectrometry, the sample is bombarded with high-energy electrons or laser radiation. This causes electrons in the outer energy levels to be knocked out of the atom, leaving the atom with an overall positive charge because it now has more protons than electrons (**Figure 4.1.2**). All atoms can be turned into positively charged ions by this method, even elements that would not normally become positively charged.

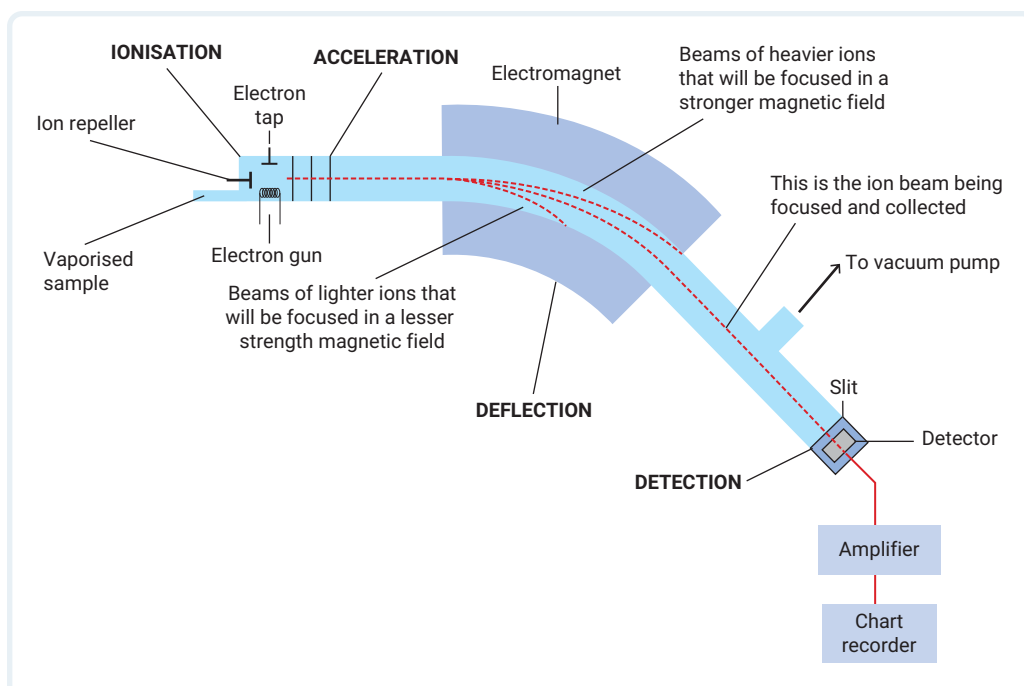


FIGURE 4.1.1 The operation of a mass spectrometer: the five stages in the process are vaporisation, ionisation, acceleration, deflection and detection.

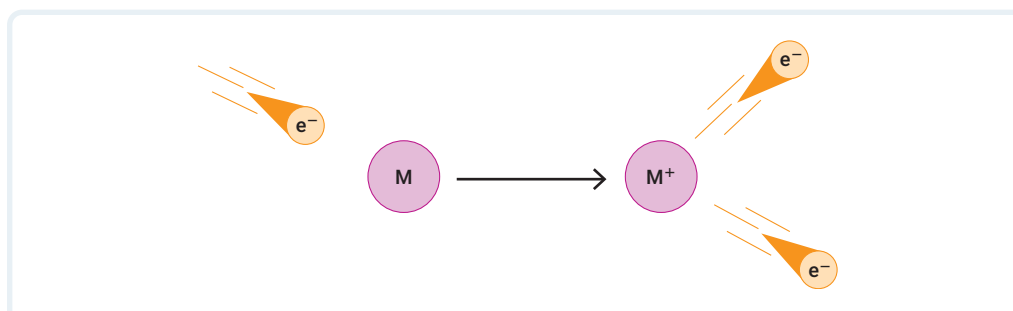


FIGURE 4.1.2 The formation of a positive ion by electron bombardment: M represents an atom or a molecule.

The positive ions in the sample are then accelerated by electric fields so they all move at high speeds. The ions pass through a magnetic field, where they undergo deflection according to their masses. Lighter ions are deflected more by the magnetic field; heavier ions undergo less deflection.

Detectors measure the number of ions that strike them. This information is transformed into graphical form – a mass spectrum – as can be seen in [Figure 4.1.3](#). This graph shows the presence of ions with different masses and their relative abundance. The x-axis of the spectrum is usually the mass-to-charge ratio of the ions, but given that the charge on each ion is usually +1, $\frac{\text{mass}}{\text{charge}}$ is equal to the relative mass of the particle.

If the sample consists of a mixture of isotopes, then each isotope has a different mass-to-charge ratio and appears as a different peak on the **mass spectrum**. The height of the peak indicates the abundance of that particle.



Worksheet
Analysing mass spectra

mass spectrum a graph of the information from a mass spectrometer, showing the number of particles of different mass (or, more precisely, mass-to-charge ratio) that are present in a sample and the relative abundance of each

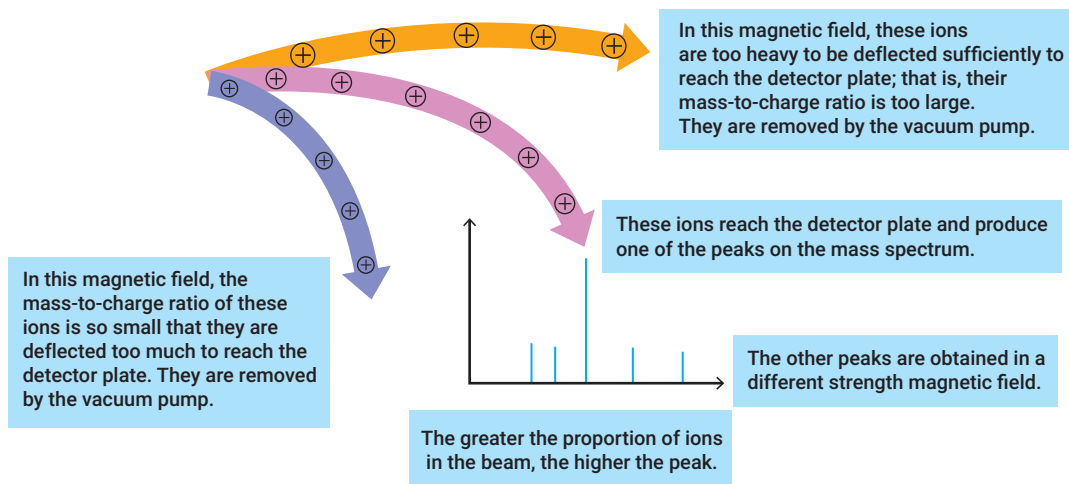


FIGURE 4.1.3 Focusing and collecting an ion stream in a mass spectrometer

For example, a sample of carbon that contains two isotopes, carbon-12 and carbon-13, will show two positive ions, each with a different mass. **Figure 4.1.4** shows that an ion with an atomic mass of 12 is present in the greatest **concentration** and there is a small amount of an ion with an atomic mass of 13. These are the two isotopes of carbon present in the sample.

concentration a measure of the amount of solute dissolved in a given amount of solvent

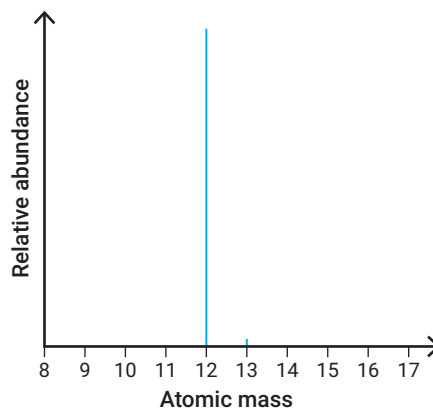


FIGURE 4.1.4 The mass spectrum of carbon

WORKED EXAMPLE 4.1.1

Consider the mass spectrum of molybdenum in **Figure 4.1.5**.

- Determine the number of isotopes of molybdenum and their relative atomic masses.
- Which isotope is present in the highest concentration?
- Which isotopes are present in equal amounts?

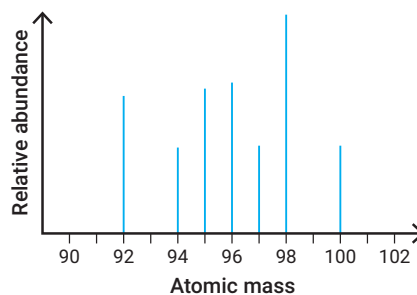


FIGURE 4.1.5 The mass spectrum of molybdenum

ANSWERS

- a 1 Identify the number of peaks in the mass spectrum.**
There are seven peaks in the mass spectrum, so there are seven isotopes of molybdenum.
- 2 Determine the atomic mass of each peak.**
They have atomic masses of 92, 94, 95, 96, 97, 98 and 100.
- b Identify the highest peak in the mass spectrum.**
The isotope with a mass of 98 has the highest concentration in the sample. This isotope has the highest peak in the mass spectrum.
- c Identify the peaks with the same height.**
The isotopes at 94, 97 and 100 are present in equal amounts. Their peaks are the same height.

LEARNING CHECK 4.1

DESCRIBING

- 1 Identify** the name of the graph generated during mass spectrometry.
- 2 Explain** why isotopes of different mass can be separated by a mass spectrometer.

ANALYSING

- 3 Figure 4.1.6** shows the mass spectrum of an element.
- a Identify** the isotopes present in this element.
- b Create** a table showing the isotopes of the element and their relative abundances.
- c Calculate** the relative atomic mass of this element and use this information to **identify** the element.

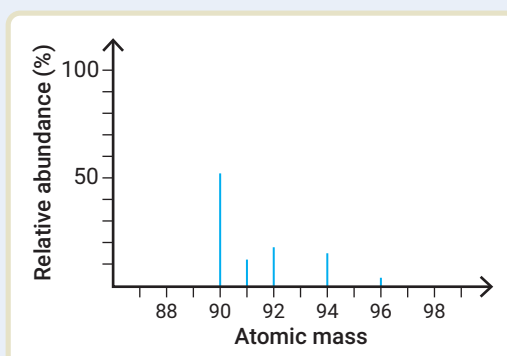


FIGURE 4.1.6 The mass spectrum of an element

spectroscopy a method of analysis performed by subjecting a sample to electromagnetic radiation and observing the frequencies absorbed or emitted

spectroscope a device used to separate light into its component wavelengths

line spectrum a pattern of lines showing the wavelengths that have been absorbed or emitted

ground state the lowest energy electron configuration of an atom

4.2 Emission spectroscopy

Spectroscopy is the process of analysing light. A **spectroscope** or spectrometer is a device used to separate light into its component wavelengths. The absorption or emission of light by an element produces a **line spectrum** that is unique for every element, so these spectra can be used to identify and analyse substances for their elemental composition.

Movement of electrons

As you learnt in Chapter 1, within an atom, electrons occur in specific energy levels around the nucleus. When all the electrons are at their lowest possible energy levels, the atom is in its **ground state**.



Worksheet
Emission spectra

When an atom absorbs energy, such as heat or light, the electrons in the energy levels around the nucleus absorb this extra energy and move to higher energy levels. An electron may move up one or more levels but cannot move part of the way between levels, so the amount of energy they absorb must be equal to the difference between one energy level and another (**Figure 4.2.1**).

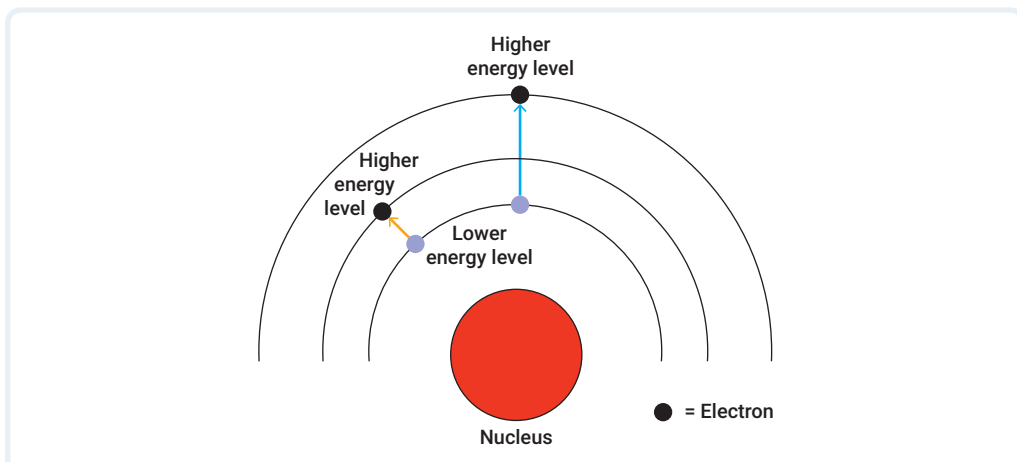


FIGURE 4.2.1 When electrons absorb energy, they can move to a higher energy level.

excited state an altered electron configuration from the ground state that occurs when electrons absorb additional energy

An atom with electrons in higher energy levels than its regular ground state is said to be in an **excited state**. Atoms with electrons in the excited, higher energy levels are unstable. After a very short time, less than one-millionth of a second, the electrons move back down to their original energy levels. As they do, they release the energy that they previously absorbed. This is shown in **Figure 4.2.2**.

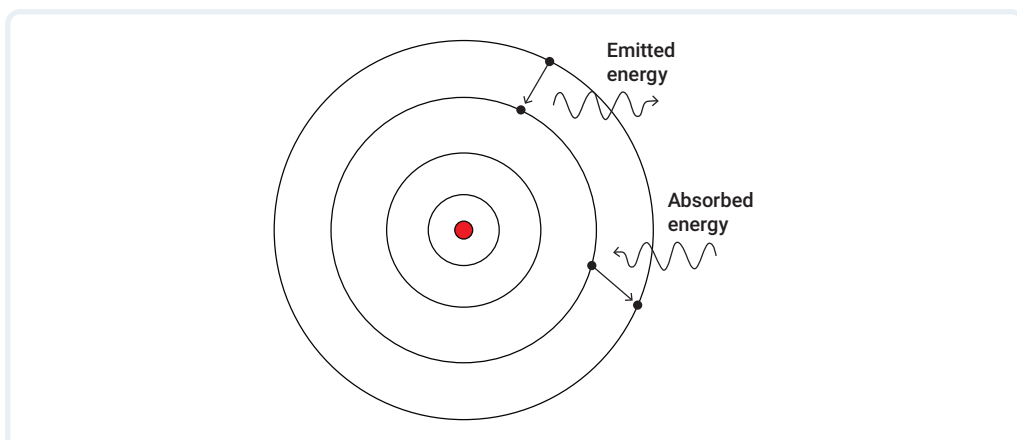


FIGURE 4.2.2 Absorption and emission of light by atoms due to electron movement

photon a particle representing a specific quantity of energy

wavelength a property of light related to the length of the wave, which can give light properties such as colour

This energy is emitted as **photons** of light. Each transition between two energy levels represents a specific amount of energy, which equates to a photon of light with a specific **wavelength**. An atom will absorb and emit photons of different wavelengths because of the different transitions between different energy levels available for its electrons when excited.

Flame tests

Some metal ions, when excited, emit light with a wavelength that corresponds to a specific colour of light that we can see. In a flame test, a sample is placed in a high-energy flame

such as that produced by a Bunsen burner. Some of the metal ions in the sample regain their valence electrons to become neutral atoms. The colours produced in the flame are due to the electrons of the neutral atom falling back to their ground state, as shown in **Figure 4.2.2**. Flame tests can be used to qualitatively analyse unknown substances to determine the metal ions present based on the colours they emit.

While flame tests are simple to conduct, it is sometimes difficult to distinguish between elements that produce similar colours. For example, barium produces a pale green flame (**Figure 4.2.3**), whereas copper produces a green flame, and zinc produces a whitish green flame.

Table 4.2.1 lists the flame colours of some elements. Magnesium and silver do not produce coloured flames.

TABLE 4.2.1 Flame colours of some elements

Ion	Flame colour
Na^+	Strong persistent orange
K^+	Lilac (pink)
Sr^{2+}	Red
Ca^{2+}	Brick red (orange-red)
Ba^{2+}	Pale green (apple green)
Cu^{2+}	Blue (halides); green (others)
Pb^{2+}	Light blue-grey
Fe^{2+}	Gold when very hot, bright blue or green turning to orange-brown
Fe^{3+}	Orange-brown



FIGURE 4.2.3 Barium produces a green flame colour.

PRACTICAL ACTIVITY 4.2.1

FLAME TESTS

Introduction

Some element ions give a characteristic colour when placed into the flame of a Bunsen burner. These different colours can be used to identify different positive ions.

Research question

What colours do different elements produce in a flame test?

Aim

To observe and compare the colours of flames produced by different elements

Materials

- solid samples (0.1 g of calcium nitrate ($\text{Ca}(\text{NO}_3)_2$), calcium chloride (CaCl_2), sodium nitrate (NaNO_3), sodium chloride (NaCl), potassium nitrate (KNO_3), potassium chloride (KCl), strontium nitrate ($\text{Sr}(\text{NO}_3)_2$), strontium chloride (SrCl_2))
- 20 mL of 1.0 mol L^{-1} hydrochloric acid (HCl)
- 20 mL distilled water
- $2 \times 100 \text{ mL}$ beakers
- platinum or nichrome wire loop
- Bunsen burner
- matches



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
The Bunsen burner and equipment will get hot and could cause burns.	Turn off the Bunsen burner or turn it to a yellow flame when not in use. Handle hot objects with care and do not place them directly onto bench tops; use a heatproof mat.
Use of chemicals	

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the two listed risks to identify specific hazards involved with each. Ask your teacher to check your risk assessment before you proceed.

Procedure

- 1 Place approximately 20 mL of 1.0 mol L⁻¹ HCl into a 100 mL beaker.
- 2 Place approximately 20 mL of distilled water into another 100 mL beaker.
- 3 Dip the wire loop into the distilled water, and then into the acid to clean it.
- 4 Hold the wire loop in the blue part of the Bunsen burner flame to remove any chemicals left on it. When the flame burns a normal blue colour, the wire loop is clean.
- 5 Dip the wire loop into the solid sample of calcium nitrate and hold it in the blue part of the Bunsen burner flame. Record the colour or colours observed.
- 6 Repeat steps 3–5 for all the solids, ensuring you record the colour of each flame.

Results

Construct a table of your results, showing the chemical name and the colour of the flame. Compare your results with those of other groups, because observation of colour can differ between individuals.

Interpretation

- 1 Describe any patterns you can identify in your results.
- 2 Compare your results and observations with those of other groups and with the information in Table 4.2.1.
- 4 Explain why the same elements have the same colour flame when a flame test is conducted.
- 5 Explain why you may not have seen the expected colours in this experiment.
- 6 Write a conclusion linking elements and their flame colours.

Evaluation

- 7 Describe any problems you had determining the colours.

Analysing emission spectra

emission spectrum

a spectrum showing coloured lines or bands due to specific wavelengths of electromagnetic radiation being emitted as excited electrons move from higher energy levels back to their ground state levels

Each element has a unique set of energy levels, so when electrons move between them it involves absorption and emission of different amounts of energy. This means that every element emits a unique series of wavelengths of light – its **emission spectrum**.

A spectrometer can separate the emitted light into its component wavelengths to produce an emission spectrum. **Figure 4.2.4** shows the operation of the spectrometer. The emitted light is composed of multiple wavelengths, which are dispersed through a prism and shone onto film (or digital array detector) to produce an emission spectrum. If you were to look through a spectrometer at the coloured flame produced in flame tests, you would see a series of coloured lines on a black background, that correspond to the emission spectrum of the element being observed.

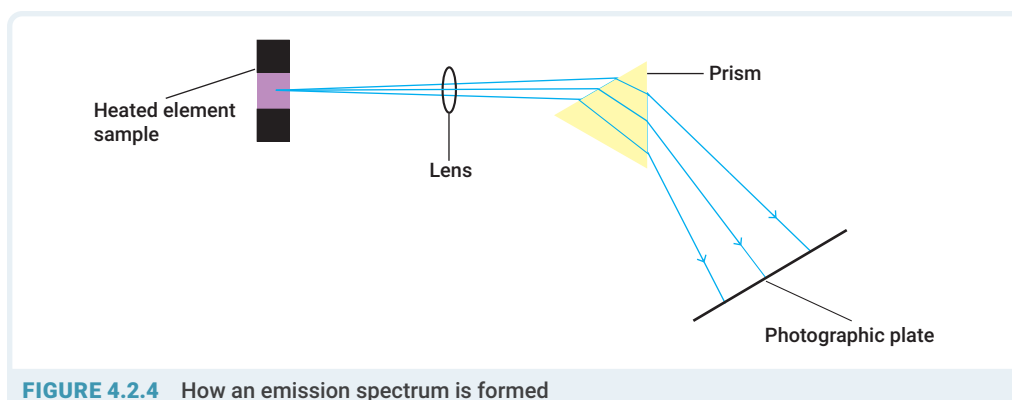


FIGURE 4.2.4 How an emission spectrum is formed

The spectra of some common elements are shown in **Figure 4.2.5**. The coloured lines are the wavelengths of light emitted as the result of the electrons falling back into their original energy levels from an excited state. Scientists can identify the elements in an unknown sample by comparing the lines in the emission spectrum to the spectra of known elements, such as those shown in **Figure 4.2.5**. This is because the emission spectra for a particular element will always be the same.

Hydrogen emission spectrum and atomic structure

As you learnt in Chapter 1, scientists' understanding of the structure of the atom has developed as new evidence from experiments has been incorporated into atomic models. Although Rutherford's model of the atom as a central nucleus surrounded by a cloud of negatively charged electrons had become widely accepted, there was still no explanation for the stability of an atom. Why didn't it just collapse?

Also, by the early 20th century, the emission spectra of many elements had been studied and measured. However, there was no explanation for why individual elements had unique spectra and what the particular wavelengths represented. Neils Bohr collected data on the hydrogen emission spectrum and proposed a theory that quantitatively explained this emission spectrum. Bohr calculated the wavelengths of different energy transitions for hydrogen, as seen in **Figure 4.2.6**, and compared these values to the wavelengths in the emission spectrum. The wavelengths of four lines in the visible region of the emission spectrum of hydrogen match the values calculated by Bohr.

In 1913, Bohr used data related to calculations for energy transitions and matching lines in emission spectra to propose that electrons move around the nucleus in fixed orbits that have only certain discrete energies. When the electrons absorb energy, they move to orbits further from the nucleus and they emit energy when they fall back to their lower energy orbit. This became known as the **Bohr model**.

Although Bohr's quantitative interpretation received scientific acceptance for the spectrum of hydrogen and a few other elements, his calculations could not be used to quantitatively interpret more complex spectra.

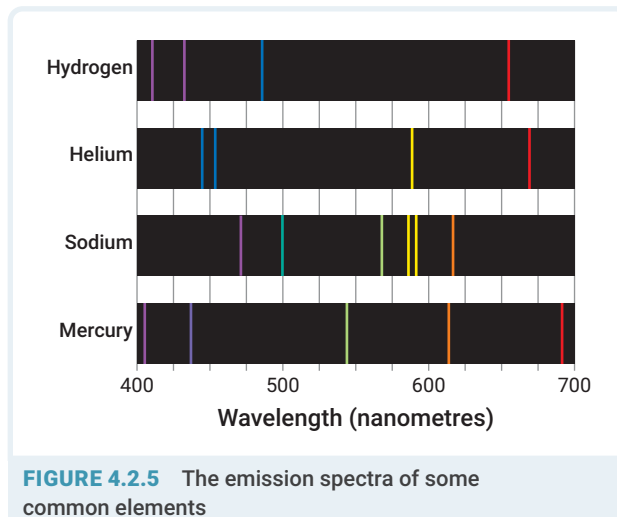


FIGURE 4.2.5 The emission spectra of some common elements

Bohr model a model of the atom that considers electrons as particles orbiting the nucleus in specific energy levels

In 1923, Edwin Schrödinger developed an equation that could be used to interpret more complex spectra, and in 1933 he received the Nobel Prize in Physics for his work on atomic structure.

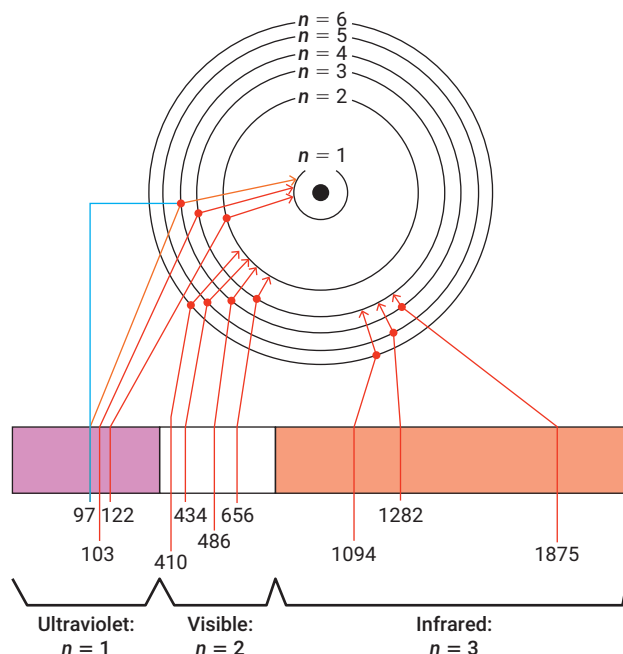


FIGURE 4.2.6 A series of electron energy levels represented as $n = 1$, $n = 2$ and $n = 3$. It shows the different electron transitions that can occur in a hydrogen atom.

LEARNING CHECK 4.2

DESCRIBING

- 1 **Explain** 'ground state' and 'excited state' in reference to electrons in an atom.
- 2 **Outline** how electrons can absorb and emit energy.
- 3 **State** the Bohr model of the atom.
- 4 **Explain** why Bohr's model of the atom was accepted.
- 5 **Describe** why there are many distinct lines in an emission spectrum.

APPLYING

- 6 **Identify** two elements that produce a distinct flame colour and explain why.
- 7 **Explain** why all atoms of an element will emit the same set of wavelengths of light when excited by intense heat.
- 8 **Explain** why sodium and magnesium have different emission spectra.

ANALYSING

- 9 **Figure 4.2.7** shows the emission spectra for several known elements and one unknown element. **Identify** the unknown element, giving reasons for your answer.

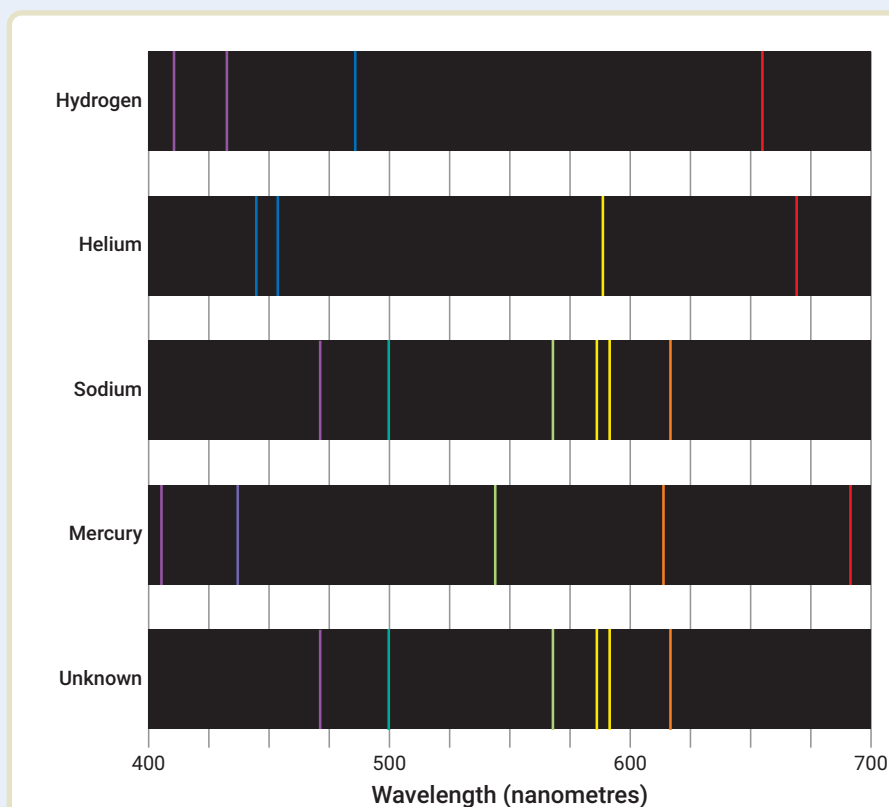


FIGURE 4.2.7 Emission spectra of some known elements and an unknown element

4.3 Absorption spectroscopy

Instead of recording the emission of light as electrons fall from excited to ground state energy levels, we can also monitor the absorption of light and energy as the electrons move to higher energy levels. Absorption spectroscopy is the technique that does this, and like emission spectroscopy, it can be used to identify the presence of certain elements, and their concentration in certain samples. The wavelengths of the lines on an **absorption spectrum** are exactly the same as those on the emission spectrum of the same element. **Figure 4.3.1** shows the absorption and emission spectra for hydrogen. If you compare these spectra, you will see the same number of lines in the same places (wavelength). The main difference is that emission spectra show coloured lines on a black background, whereas absorption spectra show black lines on a coloured background. However, all other aspects of the spectra are the same for the same element.

absorption spectrum a spectrum showing dark lines or bands due to specific wavelengths being absorbed as electrons move from ground state energy levels to higher energy levels

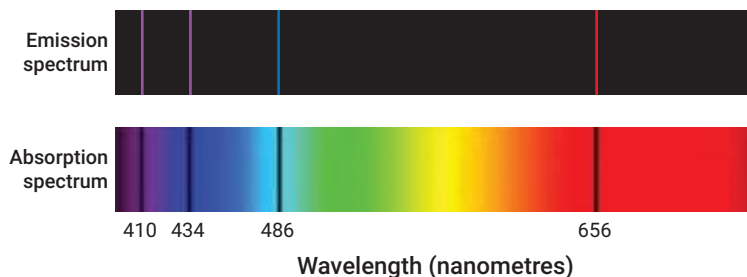


FIGURE 4.3.1 The emission and absorption spectra of hydrogen

Science Photo Library/Carlos Clarivan

Applying atomic absorption spectroscopy

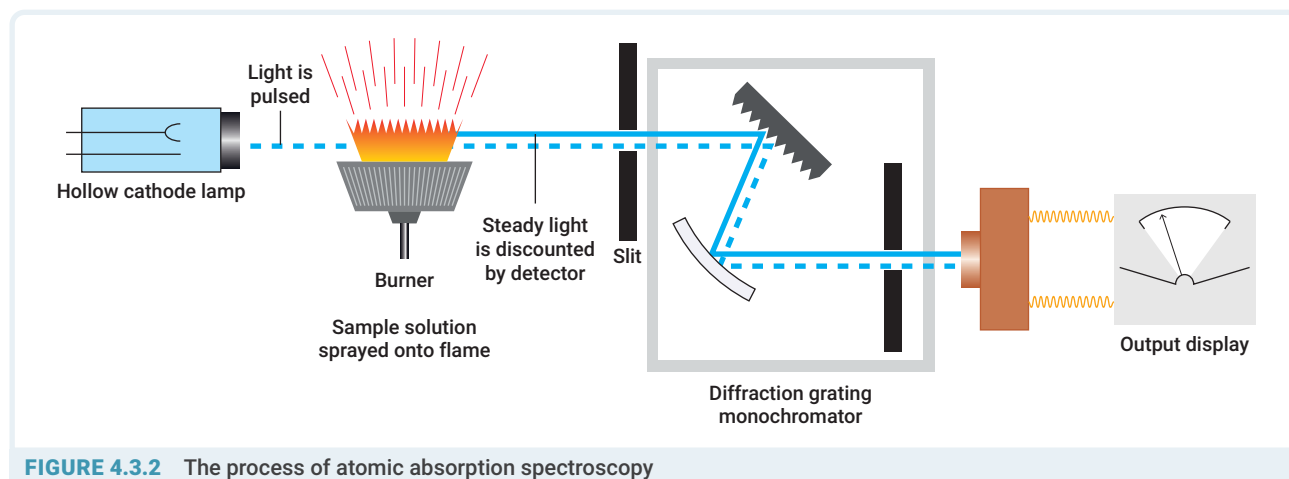
Analysing the wavelengths that make up an absorption or emission spectrum is a form of **qualitative analysis** – the spectrum can help identify the elements in the sample being tested.

Atomic absorption spectroscopy, or AAS, is a modified version of absorption spectroscopy that can be used to perform **quantitative analysis**; that is, to find the amount of an element present. It is usually used to measure the concentration of metal ions in a solution. If the metal being analysed is known, the wavelength of light that the metal absorbs during excitement is known. When light is shone through a heated vaporised sample, the absorbance of the specific wavelength of that light can be measured. The amount of absorbance is directly related to the amount of metal ion present.

This analytical technique uses an atomic absorption spectrometer. **Figure 4.3.2** outlines the technique.

qualitative analysis
a technique that determines the identity of a sample, rather than the quantity

quantitative analysis
a technique that measures the amount of a substance present in a sample



AAS can be used only to measure the quantity of a specific element. This is because the lamp for this process is made from the same element that is being tested. If we want to test for the amount of zinc, then we use a zinc lamp filament. Similarly, if mercury is being tested, then the lamp is filled with mercury vapour.

An electric current is passed through the filament or via electrodes through a gaseous sample of the element. The electric current heats the filament to very high temperatures, or electron bombardment due to the current flowing in the gas excites the atomic vapour. The

excited atoms then emit light (see section 4.2). When the lamp contains a filament or vapour of a single element, the light emitted has only the unique set of wavelengths particular to that element.

The sample being tested is **atomised**, which changes the substances it contains into gaseous atoms. When the light from the lamp passes through the atomised sample, only the element being tested for will absorb the light from the lamp. This is because its atoms have the same energy levels as the atoms that emitted the light from the lamp. Other elements in the vaporised sample will not absorb this light because the energy levels of all other atoms are different and their electrons cannot absorb the energies of the light present.

The light passes through the sample and is focused through a slit before entering a **monochromator**. The monochromator selects just one wavelength of the light for analysis by the detector. The detector measures the intensity of the light, which is then displayed as a number. This number, the **absorbance value**, is not a concentration but a measure of the amount of light that has been absorbed by the sample.

To measure the quantity of an element present, the absorbance of the sample is compared to that of a known concentration by constructing a **calibration curve**. First, several known concentrations of the element are prepared and their absorbances are measured by atomic AAS. Then, a calibration curve of absorbance against concentration is plotted. The concentration of an unknown sample can then be compared and determined by **interpolation** once its absorbance is measured.

atomise heating a substance until it exists as a vapour of gaseous atoms

monochromator a filter used to ensure that only radiation of a specific wavelength is shone at a sample

absorbance value a measurement of the relative amount of radiation that is absorbed by a sample

calibration curve a graph of absorbance values of a series of samples of known concentrations used to determine the unknown concentration of a sample

interpolation to read or construct a new data point that has not been measured but is within the range of measured data

WORKED EXAMPLE 4.3.1

To determine the concentration of mercury in a sample of fish, the absorbances of some mercury samples of known concentration were measured by atomic absorption spectroscopy. **Table 4.3.1** shows the results obtained. The fish sample was then analysed. Its absorbance value was 0.57. Note that Chapter 14 also discusses measures of concentration.

Determine the concentration of mercury in the fish sample.

TABLE 4.3.1 Measurements of absorbance of known concentrations of mercury

Mercury concentration (parts per million, ppm)	Absorbance
2.0	0.15
4.0	0.30
6.0	0.46
8.0	0.61
10.0	0.74

ANSWER

1 Construct a calibration curve using a computer spreadsheet or graphics calculator.

The mercury concentration [Hg] in units of ppm is the independent variable, and the absorbance (A, arbitrary units) is the dependent variable. Figure 4.3.3 displays the result.

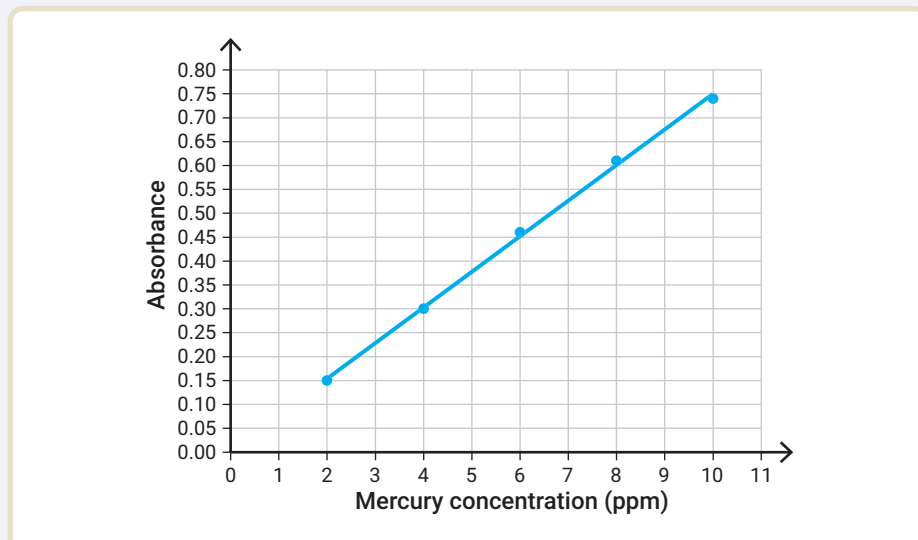


FIGURE 4.3.3 Calibration curve for mercury concentration

2 Use the calibration curve to determine the concentration of mercury in the fish.

Reading from the graph, an absorbance of 0.57 equates to a concentration of approximately 7.6 ppm.

Alternatively, we can use the linear relationship demonstrated by curve to calculate the concentration.

Since the relationship is: $A = 0.0745[\text{Hg}] + 0.005$

On rearrangement, this yields:

$$[\text{Hg}] = \frac{A - 0.005}{0.0745}$$

Hence:

$$A = 0.57$$

$$[\text{Hg}] = 7.58 \text{ ppm}$$

Rounded, $[\text{Hg}] = 7.6 \text{ ppm}$ to be consistent with precision of data.

PRACTICAL ACTIVITY 4.3.1

CONSTRUCTING A CALIBRATION CURVE TO MEASURE THE CONCENTRATION OF COPPER(II) SULFATE

Introduction

A solution of copper(II) sulfate (CuSO_4) is blue and the intensity of the colour directly relates to the solution's concentration. Several methods can be used to determine the intensity of the colour. For example, a simple light meter measures the intensity of light passing through a solution. The more light that is absorbed, the more concentrated the solution is. However, a colorimeter shines light of a specific wavelength through a sample and measures an absorbance value similarly to an atomic absorption spectrometer.

In this experiment, you will measure the intensity of light passing through different concentrations of copper(II) sulfate solutions of known concentration and construct a calibration curve. You will then be given a copper(II) sulfate sample of unknown concentration and use your calibration curve to determine the solution's concentration.

Research question

Can a calibration curve be used to accurately determine the unknown concentration of a copper(II) sulfate solution?

Aim

To determine the concentration of a copper(II) sulfate solution by constructing and using a calibration curve

Materials

- light source and light meter, colorimeter or colorimeter data probe with data logger or laptop
- 35 mL (approximately) of 1.0 mol L^{-1} of CuSO_4
- distilled water
- 50 mL beaker
- 10 mL measuring cylinder
- clear glassware to hold sample



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Use of chemicals	
Use of light meter or data logger	

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the two listed risks to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- Collect approximately 35 mL of 1.0 mol L^{-1} copper(II) sulfate from your teacher.
- You will use a piece of clear glassware to hold your copper(II) sulfate sample. Measure 10 mL into this piece of glassware or fill a provided container to overflowing.
- Take a reading of the light absorbance of the sample. The absorbance value will be provided as a read-out on your light meter or colorimeter.
- Measure 8 mL of the 1.0 mol L^{-1} copper(II) sulfate solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.8 mol L^{-1} copper(II) sulfate solution.
- Repeat steps 2 and 3 to determine an absorbance reading for this sample.
- Measure 6 mL of the 1.0 mol L^{-1} copper(II) sulfate solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.6 mol L^{-1} copper(II) sulfate solution.
- Repeat steps 2 and 3 to determine an absorbance reading for this sample.
- Measure 4 mL of the 1.0 mol L^{-1} copper(II) sulfate solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.4 mol L^{-1} copper(II) sulfate solution.
- Repeat steps 2 and 3 to determine an absorbance reading for this sample.
- Measure 2 mL of the 1.0 mol L^{-1} copper(II) sulfate solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.2 mol L^{-1} copper(II) sulfate solution.
- Repeat steps 2 and 3 to determine an absorbance reading for this sample.
- Collect 10 mL of the sample of unknown concentration from your teacher.
- Repeat steps 2 and 3 to determine an absorbance reading for this sample.

Analysis of results

- Construct a table showing your known concentrations and their absorbances. Include the result for the unknown concentration in this table.
- Construct a fully labelled calibration curve for your known results.
- Use the graph to determine the concentration of the unknown sample by interpolation.

- 4 Did your graph pass through the origin (0,0)? Why should you expect it to? Suggest reasons why your graph might not do this.

Interpretation

- 5 Compare your results with those of other groups. Did you all get the same answer? If possible, create a table showing a class set of results. Account for any differences between the results.

Evaluation

- 6 If your teacher can tell you the correct concentration of the unknown sample, discuss the accuracy of your results (i.e. how close your results were to the true value).
- 7 Identify one error that may have affected your results. Discuss its effect on the results. Identify the error as random or systematic.
- 8 How could this experiment be extended? Think of a real-world example.

LEARNING CHECK 4.3

DESCRIBING

- 1 **Outline** the process of atomic absorption spectroscopy.
- 2 **Describe** how to construct a calibration curve for the process of atomic absorption spectroscopy.
- 3 **Discriminate** between absorption and emission line spectra.
- 4 **Explain** why the lamp in an atomic absorption spectrometer contains a filament or vapour of the same element as the element being tested.

ANALYSING

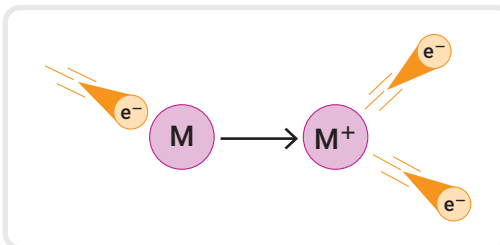
- 5 Cadmium is useful in small amounts but dangerous in large amounts. A sample of paint was tested for its cadmium content. Samples of known concentrations of cadmium were analysed by atomic absorption spectroscopy and the results in the following table were obtained.

Cadmium concentration (mg L ⁻¹)	Absorbance
0.0	0.000
1.0	0.038
2.0	0.082
3.0	0.120
4.0	0.160
5.0	0.200
6.0	0.240
Unknown	0.110

- a **Construct** a calibration curve for the known concentrations of cadmium.
 - b From the curve, **determine** the concentration of cadmium in the sample of paint.
- 6 **Describe** some ways that atomic absorption spectroscopy can help scientists and students learn more about the elements in materials, or how elements can be identified. Use examples in this section for ideas, or search the internet for other uses.

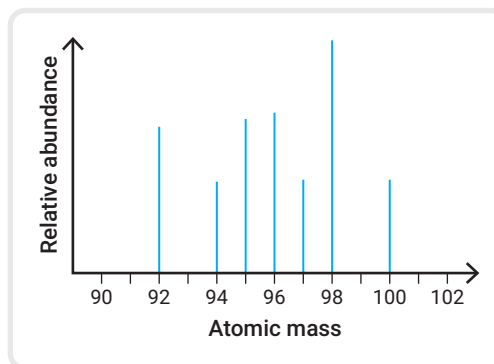
Mass spectrometry

- During mass spectrometry, samples are bombarded with electrons. This causes electrons to be knocked out of the outer shells of the atoms in the sample.
- The process results in the formation of positively charged ions.



Mass spectra

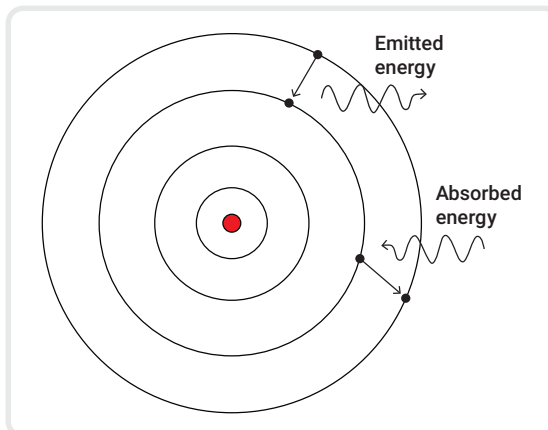
- The positive ions are detected in the mass spectrometer and a mass spectrum is generated.
- A mass spectrum of an element shows the isotopes present in a sample of that element and the relative abundance of each isotope.
- The data in a mass spectrum of an element can be used to calculate the relative atomic mass of that element.



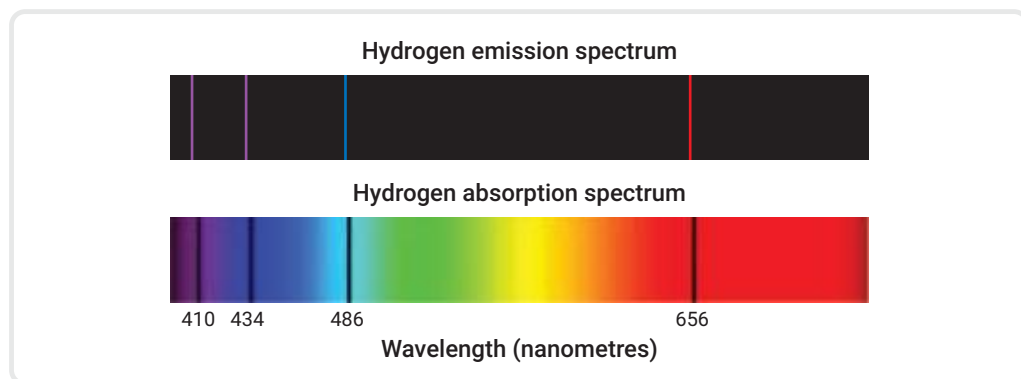
Flame tests, absorption and emission spectra

- Electrons can absorb energy and move to a higher energy level (excited state). When they return to their original energy level, they release energy, which is emitted as light.
- Flame tests and absorption and emission spectroscopy rely on electron transfer between energy levels.
- Flame tests can be used to identify some metals by the colour of their flame.
- Different metals emit different colour flames.

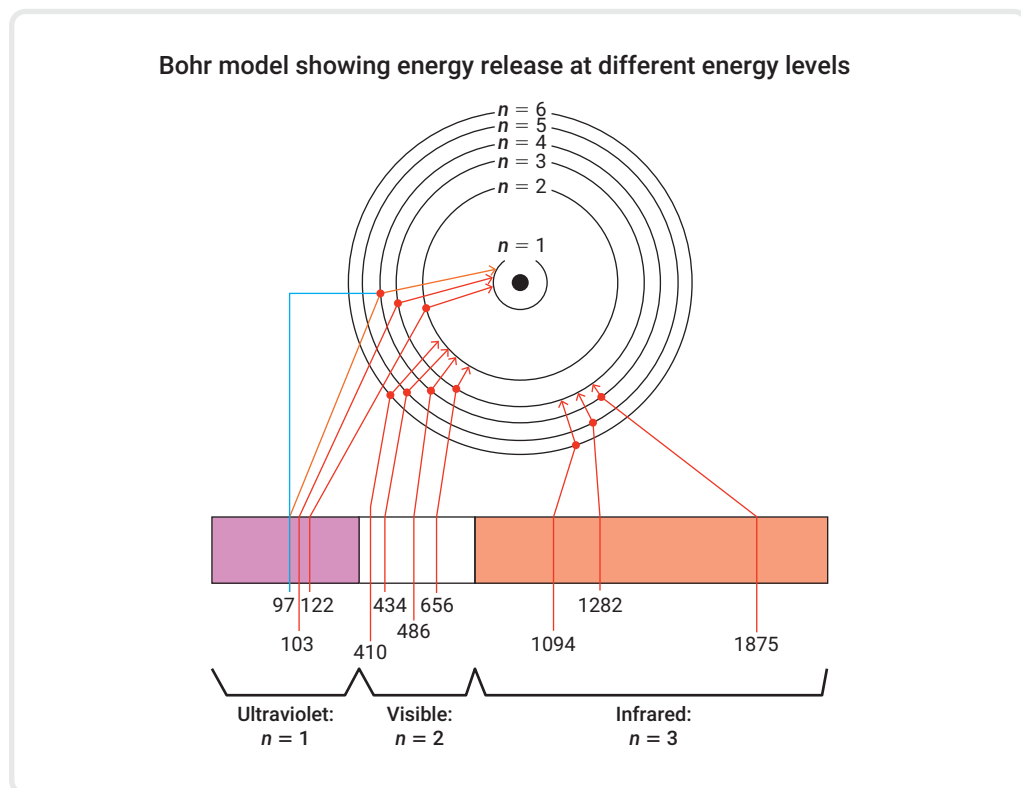
Science Photo Library/Martyn F. Chillmaid



- Atomic absorption spectroscopy can be used to determine the concentration of metallic ions in solution.
- Absorption and emission line spectra of an element show exactly the same number of spectral lines at the same wavelength.



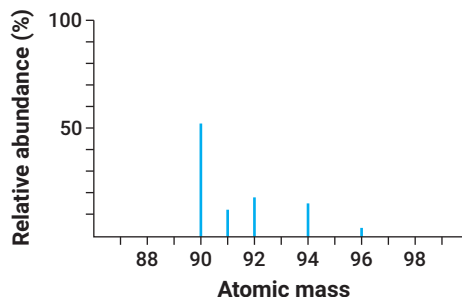
- The Bohr model of the atom states that electrons exist in discrete energy levels, and electrons absorb and release energy as they transition between energy levels.



MULTIPLE CHOICE

1. Which one of the following species would *not* be deflected by a magnetic field?
A Proton
B Neutron
C Electron
D Ion
2. In a mass spectrometer, the relative masses of particles are determined by using the fact that:
A the velocity at which the particles hit the detector can be accurately determined.
B the force with which the particles hit the detector can be measured.
C the path followed by the particles depends on their charge and mass.
D larger particles are deflected more by a magnetic field.
3. In terms of isotopic analysis, the purpose of a mass spectrometer is to:
A determine the isotopic molar masses of a chemical compound.
B identify the existence of isotopes of an element and determine the masses of the different isotopes.
C identify the ratio of relative atomic masses of several chemical compounds.
D identify the existence of isotopes of an element and their abundances, to determine relative atomic masses of a chemical compound.
4. The lamp used in an atomic absorption spectrometer produces radiation:
A that will cause an electron in the element being analysed to be excited to a higher energy state.
B over a wide range of energies.
C with energies in the infrared part of the spectrum that will excite electrons to higher states.
D that will cause an electron in the element being analysed to move to a lower energy state.
5. A major advantage of atomic emission spectroscopy over flame tests is:
A the samples don't have to be heated.
B the samples are not destroyed by the analysis technique.
C more than one element in a sample can be identified.
D compounds as well as elements can be identified.
6. Scientist Niels Bohr is credited with:
A proposing the atom consists of a nucleus and electron cloud.
B producing an accurate emission spectrum for hydrogen.
C calculating the wavelengths of the lines in an emission spectrum of hydrogen.
D proposing that electrons exist in discrete energy levels around the nucleus.

Questions 7 and 8 refer to the following mass spectrum of an element.



7. Which of the following statements related to the spectrum is *incorrect*?
 - A The element has five isotopes.
 - B The most abundant isotope has an atomic mass of 90.
 - C It is assumed all the ions produced have the same charge.
 - D The element is thorium.
8. What is the approximate relative atomic mass of the element?
 - A 90
 - B 91
 - C 92
 - D 93
9. Which of the following metal ions could be used in fireworks to produce a deep red coloured light?
 - A Sr^{2+}
 - B Ba^{2+}
 - C Mg^{2+}
 - D Na^{+}
10. The flame colours of three compounds are given below.

Compound	Flame colour
Sodium chloride	Yellow
Calcium chloride	Brick-red
Potassium chloride	Lilac

A crushed sample of an unknown chemical was tested by dipping a piece of clean platinum wire into it, then inserting the wire into the hottest part of a Bunsen burner flame. Three students each conducted the same test and made the following observations:

Student 1: The flame was reddish.

Student 2: The flame was brownish red then, after a short time, became blue again.

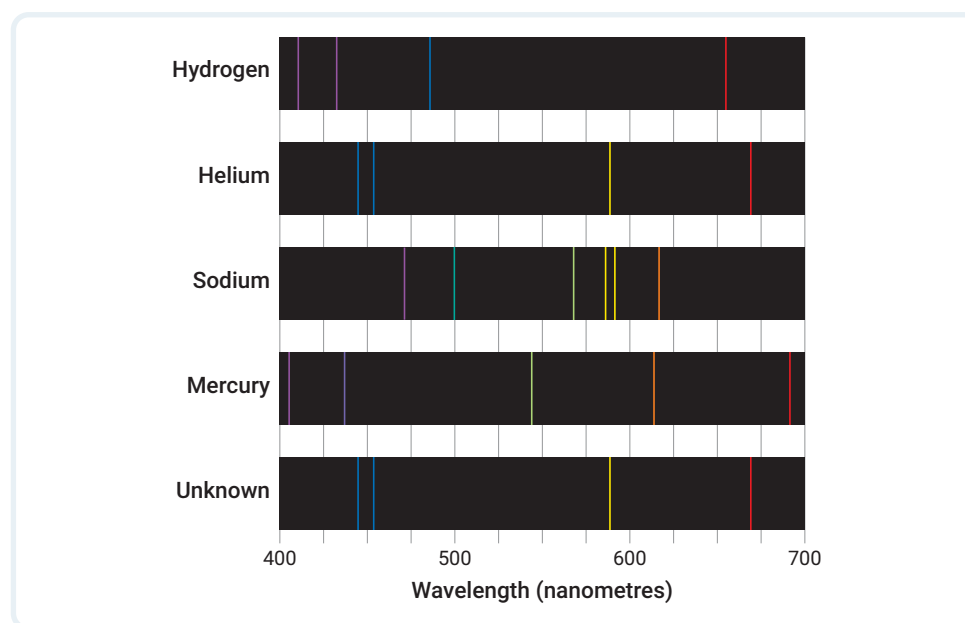
Student 3: The flame became dark red when the wire was inserted into the Bunsen burner flame.

Based on the student's observations, it could be concluded that the sample contains a large portion of:

- A chlorine.
- B calcium chloride.
- C calcium.
- D potassium chloride and sodium chloride.

SHORT RESPONSE

11. The line emission spectra of several common elements and one unknown element are shown below.



- a **Identify** the unknown element.
 - b The atomic absorption spectrum of a second unknown element, at wavelengths 400–650 nm, had absorption lines at 613, 545, 437 and 405 nm.
 - i **Describe** the appearance of the absorption spectrum.
 - ii **Explain** why the emission spectra shown above can be used to identify an element from its absorption spectrum.
 - iii **Identify** the second unknown element.
12. **Explain** why the Bohr model of the atom was accepted.

CROSS-CHAPTER QUESTION

13. Boron has two isotopes: ^{10}B and ^{11}B .

Relative abundance (%)	Mass number
20	10
80	11

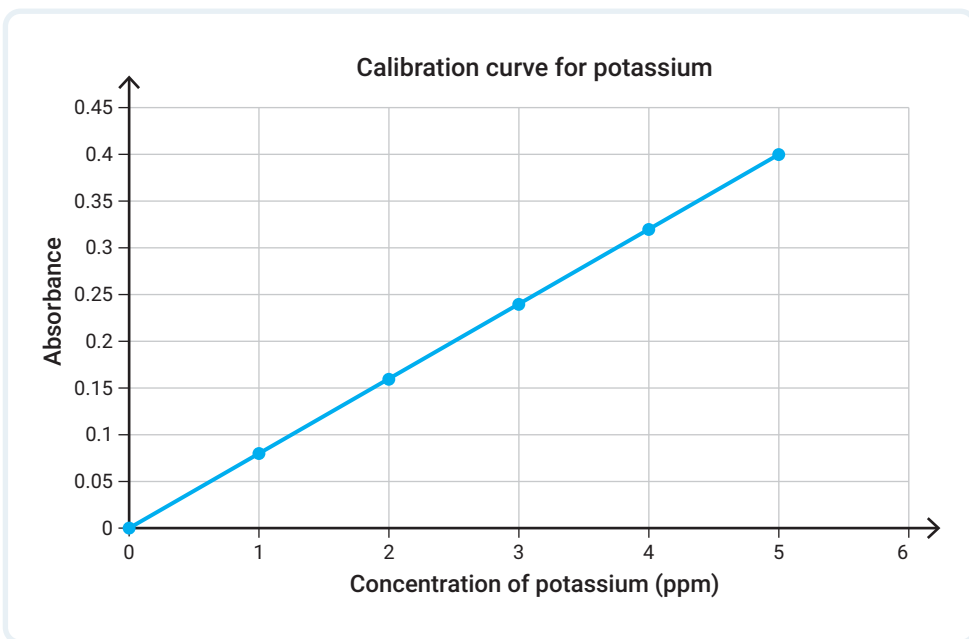
- State** the electron configuration of boron.
- Apply** nuclear symbol notation to the boron-11 species produced by the mass spectrometer.
- Calculate** the relative atomic mass of boron.
- Construct** a diagram to show the mass spectrum of boron.

DATA ANALYSIS

14. Analyse data

Bananas are an important source of the metal potassium, which is essential in trace quantities in virtually all animals. A 3.50g sample of bananas was pulped and dissolved in 20.0 mL of acid. The mixture was filtered and diluted to 100.0 mL with distilled water. A small quantity of this solution was then passed through an atomic absorption spectrometer along with a number of solutions of known potassium concentration. The calibration curve generated by this experiment is shown below. The absorbance of the diluted banana solution was recorded as 0.19.

- Explain** why it was necessary to generate a calibration curve for various potassium concentrations in this experiment.
- Use the calibration curve to **determine** the concentration of potassium in the diluted banana solution.
- Determine** the concentration of potassium in the undiluted acid solution in parts per million (ppm).



15. Analyse data

The level of zinc in some food was labelled on the container. Students analysed samples of the food by atomic absorption spectroscopy to determine whether the level given on the label was correct. Samples of known concentration of zinc were analysed and the following data was obtained.

Zinc concentration (mg L ⁻¹)	Absorbance
1	0.03
2	0.06
5	0.16
10	0.32
15	0.48

- Explain** why the presence of a metal, such as calcium, in the food would not have interfered with this process.
- Construct** a calibration curve with concentration of zinc on the horizontal axis and absorbance on the vertical axis.
- The sample of food gave an absorbance reading of 0.36. From the graph, **determine** the concentration of zinc in the food.
- The label on the container stated that the level of zinc in the food did not exceed 7.5 mg L⁻¹. **Explain** whether this statement was true.

Introduction to bonding



Jurik Peter/Shutterstock.com

SYLLABUS DOT POINTS

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 transitional 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.
- Discriminate between the terms *empirical formula*, *molecular formula* and the *formula unit*.





- Determine Lewis (electron dot) structure of molecules and ions showing all valence electrons for up to four pairs for each atom.
- Identify the numbers of bonding and lone pairs of electrons around each atom in a molecule.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Although there are 118 known elements, an infinite number of different substances are made from these elements. To form these new substances, atoms bond together to form compounds with their own unique properties. Understanding how this happens is fundamental to our ability to create new substances.

Worksheets

- Ionic and covalent bonds
- Bonding crossword
- Comparing chemical bonding



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ The atom has a basic structure, which involves subatomic particles in the nucleus and the outer shell(s).
- ✓ Valence electrons occur in the outermost energy level or shell of an atom.
- ✓ Ions are formed when atoms lose or gain electrons.
- ✓ The electron configuration of an atom can be identified from the periodic table.
- ✓ The first ionisation energy and electronegativity help inform the type of bond formed.
- ✓ At a high level, elements can be classified as metals, metalloids and non-metals.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ identify that the chemical bonds that form between atoms depend on the electrons in the valence shell and the position of the element on the periodic table
- ✓ identify the three types of bonds formed as due to transfer or sharing of electrons
- ✓ recognise that all types of bonds involve electrostatic attraction between oppositely charged substances
- ✓ recognise that metallic bonds form because of the low ionisation energies of metals
- ✓ recognise that ionic bonds are formed by the transfer of electrons from a metal to a non-metal where there is a significant difference in electronegativities
- ✓ identify the charges on ions formed from elements based on the position of the element on the periodic table, excluding the transition metals
- ✓ illustrate examples of multiple ions and ionic compounds formed by transition metals
- ✓ recognise that covalent bonds are formed by the sharing of electrons of two non-metals that have a small difference in electronegativities
- ✓ name and write formulas for ionic and molecular compounds
- ✓ represent ionic and molecular compounds using Lewis structures.
- ✓ identify bonding and lone pairs of electrons in Lewis and structural diagrams of molecules
- ✓ explain the difference between empirical and molecular formulas and formula unit.

element a substance consisting of only one type of atom

compound a substance consisting of atoms of two or more elements chemically bonded together in a fixed ratio

chemical bond a force of attraction between atoms that holds them together to form compounds or molecules

molecule a group of atoms covalently bonded together, representing the smallest unit of an element or a compound that can take part in a chemical reaction

5.1 Chemical bonding

Although there are 92 naturally occurring elements on Earth, the vast majority do not exist in large quantities in an uncombined state. **Elements** exist more commonly in **compounds** because they form **chemical bonds** with atoms of other elements. In addition, several elements exist as **molecules**, rather than single atoms. For example, hydrogen exists as hydrogen gas (H_2) rather than as a single H atom. These atoms form bonds to form a lower-energy, more chemically stable state with a full valence shell.

Forming chemical bonds can involve the transfer or sharing of electrons. The bonded atoms are held together by the electrostatic force of attraction between positive (the nucleus of an atom) and negative charges (electrons).

Whether an atom forms chemical bonds by the transfer or sharing of electrons depends on the arrangement of the electrons in the atom, particularly the valence electrons, as well as the element (or elements) it combines with. For example, lithium, which has a single valence electron, loses its valence electron to have a full outer shell. As such, lithium is generally involved in chemical bonds formed by transferring electrons. The number of electrons that are

gained, lost or shared by atoms in order to achieve a full valence shell reflects the atom's valency. The valency of an atom helps us determine the number of single bonds the atom can form. For example, nitrogen has five valence electrons and therefore requires three electrons to achieve a full outer shell. As such, nitrogen has a valency of 3 and can form three covalent bonds.

The three types of chemical bonds that occur because of the transfer or sharing of electrons are:

- metallic bonds, which involve the electrostatic attraction between metal ions and a 'sea' of delocalised electrons (Note: When bonds form between atoms of different metallic elements, **alloys**, or amalgams, are produced, rather than compounds.)
- ionic bonds, which involve the electrostatic attraction between positive metal atoms that have lost electrons and non-metal atoms that have gained electrons
- covalent bonds, which involve electrostatic attraction between shared electrons and nuclei of non-metal atoms.

Figure 5.1.1 shows the relationship between metals and non-metals, and the type of bonding in which they participate.

alloy a mixture of two or more metals, which has properties of each of the constituent metals



Weblink
Chemical bonding

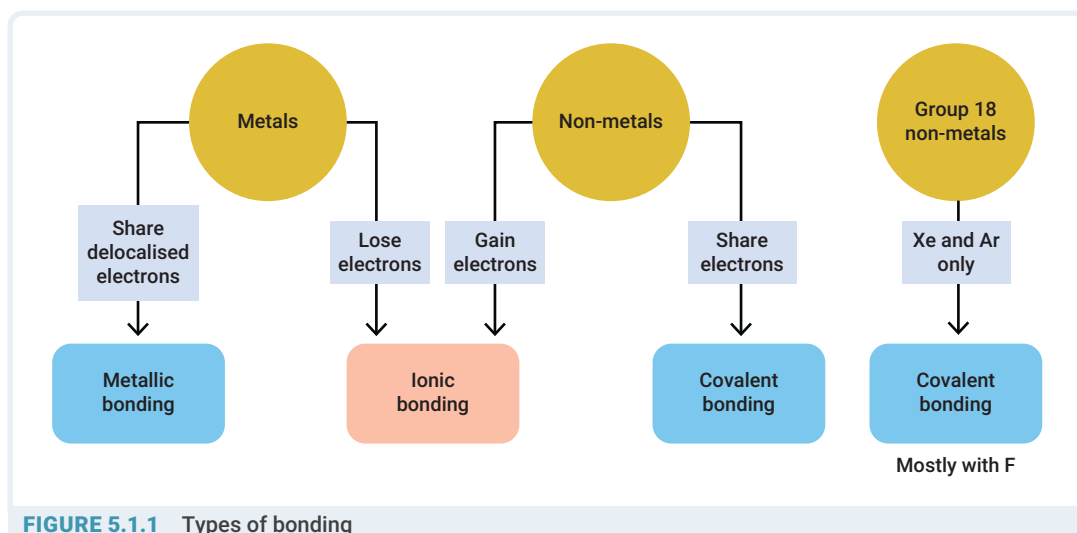


FIGURE 5.1.1 Types of bonding

LEARNING CHECK 5.1

DESCRIBING

- 1 **Identify** the three types of chemical bonds that hold atoms together.
- 2 **Recall** what determines the type of chemical bond that is formed between atoms.

APPLYING

- 3 **Explain** why chemical bonds form between atoms.
- 4 **Explain** why elements in the same group of the periodic table bond in the same way.
- 5 **Classify** the bonding in each of the following solids as ionic, metallic or covalent.
 - a Magnesium
 - b Water
 - c Barium chloride
 - d Phosphorus triiodide
 - e Lithium sulfide

5.2 Ionic bonding

Although metals are common, they are not usually found naturally as pure metals. This is because metal atoms generally tend to lose electrons to obtain a filled outer shell and readily combine with elements that tend to gain electrons to form compounds.

For example, sodium metal is very reactive and is never found naturally in its metallic state. When sodium is brought together with chlorine gas, it forms the compound sodium chloride (Figure 5.2.1), which is more commonly known as table salt (Figure 5.2.2).



Weblink
Ionic bonding

Science Photo Library



FIGURE 5.2.1 Sodium metal and chlorine gas react violently.



FIGURE 5.2.2 Crystals of sodium chloride are completely different in form from the constituent elements sodium and chlorine.

Science Source/ER Degginger

ion an atom that has become charged through the gain or loss of one or more electrons

cation a metallic atom that has become positively charged by losing one or more electrons

anion a non-metallic atom that has become negatively charged by gaining one or more electrons

Ionic bonding involves the transfer of electrons. When atoms lose or gain electrons to attain a full outer shell, they form a charged particle called an **ion**. The ions that form due to the loss of one or more electrons are positive (+) and are called **cations**. Ions that form due to the gain of one or more electrons are negative (–) and are called **anions**.

Sodium chloride is an ionic compound made up of a combination of a metal and non-metal. The sodium atom with the electron configuration 2,8,1 has a tendency to lose one electron to become like neon (2,8), whereas the chlorine atom with the configuration 2,8,7 tends to gain one electron to become like argon (2,8,8). The sodium atom becomes a positively charged sodium ion and is represented as Na^+ . It now has 11 protons and 10 electrons and an overall +1 charge. The chlorine atom becomes a negatively charged chloride ion and is represented as Cl^- . It now has 17 protons and 18 electrons and an overall –1 charge. The resulting compound, sodium chloride (NaCl), is electrically neutral because the positive and negative charges balance each other (Figure 5.2.3).

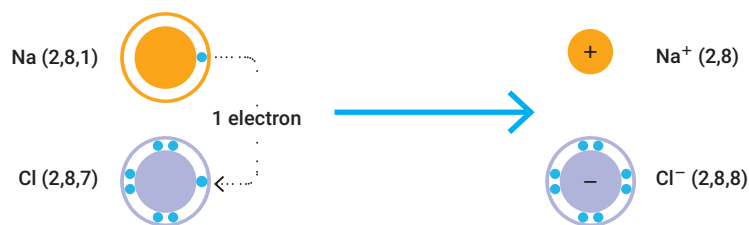
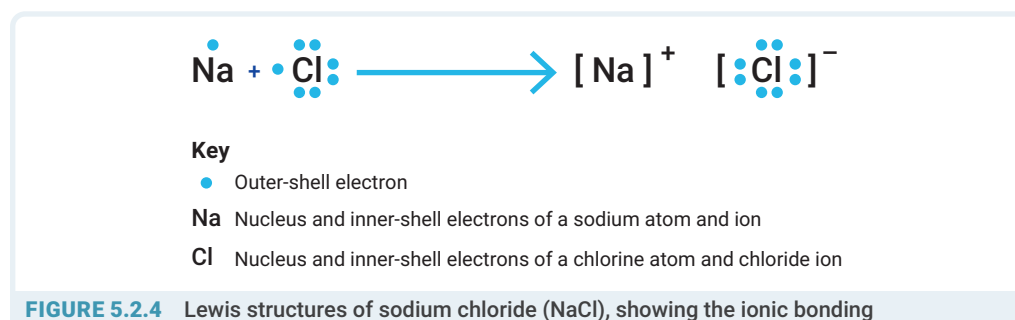


FIGURE 5.2.3 The formation of sodium chloride involves electron transfer.

Roland Smith

When ionic compounds form, the total number of electrons lost and the total number of electrons gained must be equal. This means that the total positive charge on the cation(s) must equal the total negative charges on the anion(s), so the formula of an ionic compound will always represent an electrically neutral substance.

In Figure 5.2.3, only the valence electrons in the outer shell of the atoms have been drawn. This reaction can also be represented using **Lewis structures**. These diagrams are often more convenient to use because the valence electrons can be represented as dots (or crosses), as shown in Figure 5.2.4.

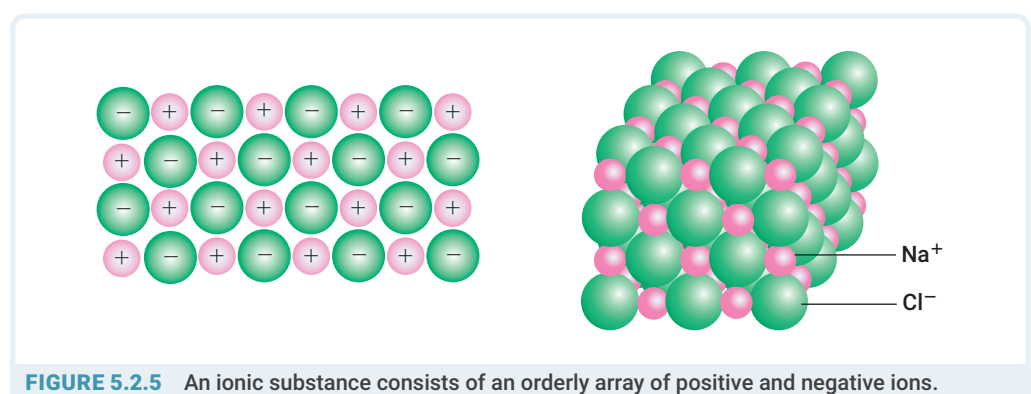


Lewis structure a representation of the electron arrangement of atoms, using dots or crosses to represent electrons in their shells; also referred to as a Lewis dot structure, electron dot structure or electron dot formula

There is a strong electrostatic attraction between the positive cations and the negative anions and this attraction results in a chemical bond being formed. This type of chemical bond is called an **ionic bond**. Ionic bonding is due to the electrostatic force of attraction between oppositely charged ions.

Ionic compounds have a crystalline 3D lattice structure in which the positive and negative ions are arranged in an orderly fashion with every positive ion being surrounded by negative ions and every negative ion being surrounded by positive ions (Figure 5.2.5). The electrostatic attraction between pairs of oppositely charged ions extends throughout the whole crystal. There are no separate units of NaCl, just a continuous lattice of cations and anions held together by ionic bonds.

ionic bond the mutual, non-directional, electrostatic attraction of positively and negatively charged ions, which holds them together in a lattice in a compound



The chemical formula represents the ratio of ions, rather than the actual number of ions, because there are no discrete molecules in an ionic compound. A formula that gives the simplest whole-number ratio of atoms or ions is called an **empirical formula**. More specifically, in any ionic compound, the smallest unit can be referred to as a **formula unit**. It shows the overall ratio of positive and negative ions. For example, the formula unit for sodium chloride has a ratio of 1 sodium cation : 1 chloride anion.

Magnesium chloride is another example of an ionic compound. It has a different ratio of cations to anions, because the magnesium atom with the electron configuration 2,8,2 has a tendency to lose two electrons. Therefore, the formula unit for magnesium chloride is MgCl_2 because its overall ratio is 1:2. Figure 5.2.6 demonstrates the formation of magnesium chloride.

empirical formula a chemical formula showing the simplest ratio of elements in a compound rather than the total number of atoms in the compound

formula unit a chemical formula showing the overall ratio of positive and negative ions in an ionic compound

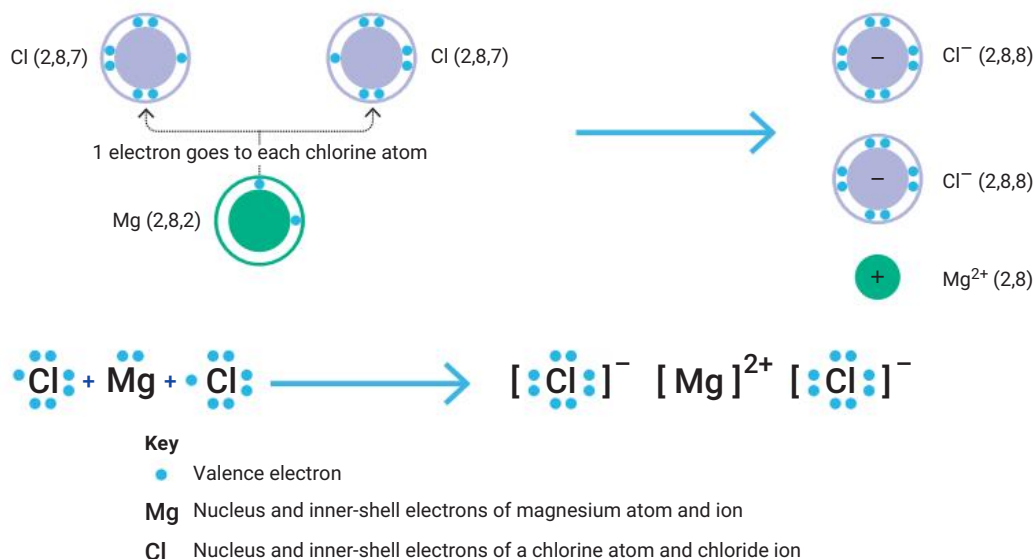


FIGURE 5.2.6 The formation of magnesium chloride involves electron transfer; it can also be shown as Lewis structures.

Ions and the periodic table

For the main group elements, the periodic table can be used to deduce the charge of an ion because the number of electrons an element loses or gains depends on the number of valence electrons. Many transition metal elements (in groups 3–12) can form multiple ions because of the small energetic difference between the s and d subshells. This means they can lose both the 's' electrons and a variable number of 'd' electrons (**Figure 5.2.7**).

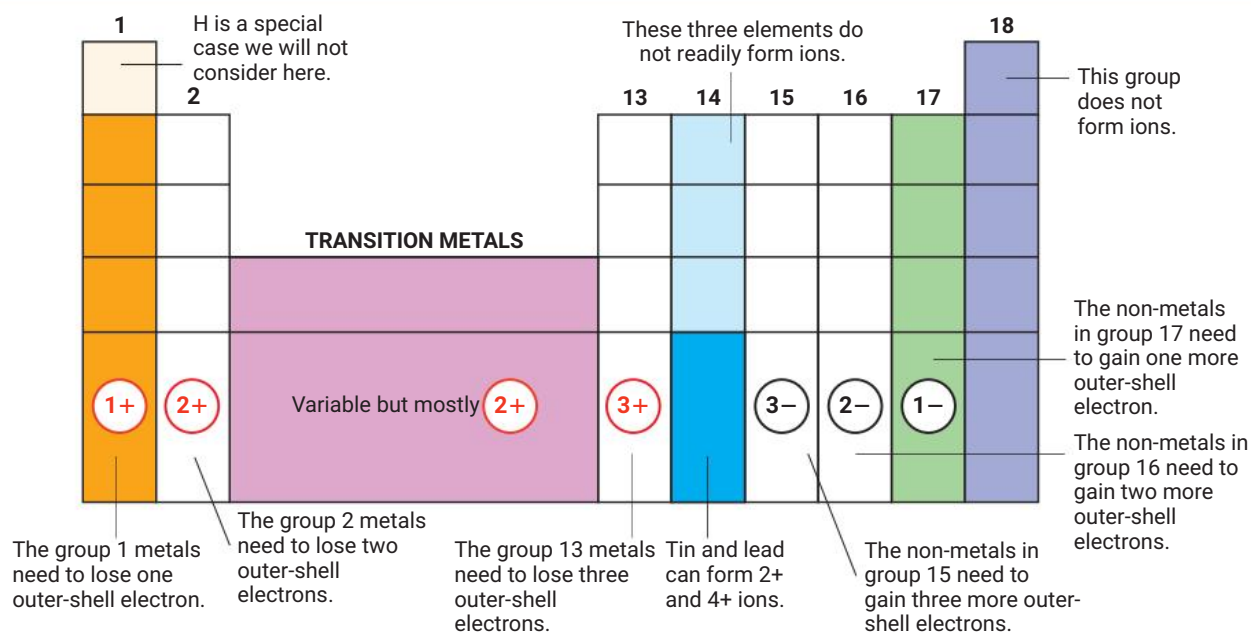


FIGURE 5.2.7 The charges of ions formed from elements according to their group on the periodic table

You will recall learning about ionisation energy and electronegativity in Chapter 3. Both help determine if an ionic bond will form between two atoms. Metals, on the left side of the periodic table, have a low first ionisation energy and low electronegativity (see Figure 3.2.7, page 52) and lose electrons very easily to form cations. Non-metals, located on the right side of the periodic table, have high first ionisation energy and high electronegativity. This means that it takes a lot of energy to remove electrons from the valence shell of a non-metal, but it will attract electrons towards itself, readily forming anions. For example, sodium has a first ionisation energy of 502 kJ mol^{-1} and an electronegativity of 0.9 compared to chlorine which has a first ionisation energy of 1257 kJ mol^{-1} and an electronegativity of 3.2. This explains the violent reactivity that occurs when an atom of sodium comes into contact with a molecule of chlorine and why the electron is easily transferred from the sodium atom to the chlorine atom. Notice the significant difference between the electronegativities of sodium and chlorine ($3.2 - 0.9 = 2.3$). This large difference exists for all ionic compounds.

Names of ionic compounds

The International Union of Pure and Applied Chemistry (IUPAC) has the responsibility for making rules for naming elements and other chemical substances. In the names and formulas of ionic compounds composed of only two elements, the positively charged ion is always listed first and the negatively charged ion is listed second. The name of the positive ion is the same as the name of the element from which it was formed. However, in the name of the negative ion, the end of the non-metal parent element is replaced with ‘-ide’. For example, chlorine becomes chloride, oxygen becomes oxide, and bromine becomes bromide. Sodium chloride, magnesium oxide and potassium bromide are correct ionic compound names. These ionic compounds are made of **monatomic ions**.

monatomic ion an ion consisting of one atom

Transition metals have slightly different electron structures from other metals. As a result, many transition metals can have more than one possible charge. A simple naming system is used to show which ion is present. For example, the metal iron (Fe) forms both Fe^{2+} and Fe^{3+} ions. In the past, the ending of the name was used to indicate the charge of the ion. The one with a lower charge ended in ‘-ous’ and the one with a higher charge ended in ‘-ic’. But now there is a much easier system in which the charge is shown as a Roman numeral, as shown in **Table 5.2.1**. Note that there is no gap between the parenthesis and the symbol of the element. The former names are still sometimes used in the chemical industry.

TABLE 5.2.1 Examples of naming transition metal ions

Ion	Fe^{2+}	Fe^{3+}	Cu^{+}	Cu^{2+}
Current naming convention	Iron(II)	Iron(III)	Copper(I)	Copper(II)
Former name	Ferrous	Ferric	Cuprous	Cupric

Two different ionic compounds form between iron and chlorine: iron(II) chloride (FeCl_2) and iron(III) chloride (FeCl_3). The particular compound is indicated by its name and the formula. Similarly, copper also forms two compounds with chlorine: copper(I) chloride (CuCl) and copper(II) chloride (CuCl_2).

In many ionic compounds, the positive ion or the negative ion (or both) consist of two or more atoms that are strongly bonded together and act as a single entity. These ions are called **polyatomic ions**. Polyatomic ions share their outer-shell electrons and the charge on polyatomic ions is spread over the whole ion. As shown in **Table 5.2.2**, the names of the compounds involving polyatomic ions follow the convention discussed.

polyatomic ion a group of atoms that chemically bond together and have an overall charge, which can be positive or negative; they can take part in ionic bonding

TABLE 5.2.2 Some examples of common polyatomic ions and their valencies

Name of ion	Formula	Valency	Example of a compound	Formula of example compound
Ammonium	NH_4^+	+1	Ammonium chloride	NH_4Cl
Hydroxide	OH^-	-1	Iron(III) hydroxide	$\text{Fe}(\text{OH})_3$
Nitrate	NO_3^-	-1	Silver nitrate	AgNO_3
Sulfate	SO_4^{2-}	-2	Copper(II) sulfate	CuSO_4
Carbonate	CO_3^{2-}	-2	Calcium carbonate	CaCO_3
Phosphate	PO_4^{3-}	-3	Sodium phosphate	Na_3PO_4

When referring to ions and ionic compounds, the valency of an ion is related to its charge. This is shown in Table 5.2.2. Note: When multiple polyatomic ions are required, the formula is enclosed in brackets with the number of ions outside of the brackets as shown for iron(III) hydroxide above.



Weblink
Matchmaker:
ionic bonding

Formulas of ionic compounds

There is one simple rule when determining the formulas of ionic compounds: ionic compounds have no net charge. This means the total number of positive charges equals the total number of negative charges. When determining the formula, you need to determine the ratio of ions that will achieve this.

WORKED EXAMPLE 5.2.1

Write the formula for calcium sulfide.

ANSWER

1 Determine the ions present.

Calcium sulfide is composed of calcium ions and sulfide ions.

2 Determine the charge on the ions.

The ions present are Ca^{2+} and S^{2-} .

3 Compare the charges and determine the ratio that gives equal positive and negative charges.

These ions have the same-sized charge. So, one sulfide ion balances one calcium ion and the total positive charge equals the total negative charge. The total charge on the molecule is equal to zero.

	Calcium cation	Sulfide anion	
Charge of ion	+2	-2	
Number of atoms	1	1	Total charge on the molecule
Total ion charge	+2	-2	0

4 Write the formula.

The charges on the ions are not shown in the overall chemical formula.
The formula is CaS .

WORKED EXAMPLE 5.2.2

Write the formula for cobalt(III) oxide.

ANSWER

1 Determine the ions present.

Cobalt(III) oxide is composed of cobalt ions and oxide ions.

2 Determine the charge on the ions.

The ions present are Co^{3+} and O^{2-} .

3 Compare the charges and determine the ratio that gives equal positive and negative charges.

These ions have different-sized charges. So, two cobalt ions balance three oxide ions and the total positive charge equals the total negative charge.

	Cobalt cation	Oxide anion	
Charge of ion	+3	-2	
Number of atoms	2	3	Total charge on the molecule
Total ion charge	+6	-6	0

4 Write the formula.

The charges on the ions are not shown in the overall chemical formula.

The formula is Co_2O_3 .

WORKED EXAMPLE 5.2.3

Write the formula for ammonium sulfate.

ANSWER

1 Determine the ions present.

Ammonium sulfate is composed of ammonium ions and sulfate ions.

2 Determine the charge on the ions.

The ions present are NH_4^+ and SO_4^{2-} .

3 Compare the charges and determine the ratio that gives equal positive and negative charges.

These ions have different-sized charges. So, two ammonium ions balance one sulfate ion and the total positive charge equals the total negative charge.

	Ammonium cation	Sulfate anion	
Charge of ion	+1	-2	
Number of atoms	2	1	Total charge on the molecule
Total ion charge	+2	-2	0

4 Write the formula.

The charges on the ions are not shown in the overall chemical formula.

The formula is $(\text{NH}_4)_2\text{SO}_4$.

(Note: When there are multiple polyatomic ions, the formula is enclosed in brackets with the number of ions outside of the brackets.)

LEARNING CHECK 5.2

DESCRIBING

- 1 **Define** 'ionic bonding'.
- 2 **Recall** the overall charge on an ionic compound. How is this achieved?

APPLYING

- 3 **Explain** why metals generally form cations and non-metals generally form anions when these two types of substances react together.
- 4 Draw the Lewis structure for:
 - a barium
 - b fluorine
 - c barium fluoride.
- 5 **State** the charge that would be expected for an ion formed from an atom with an electron configuration of:
 - a $1s^2 2s^2 2p^6 3s^2$
 - b $1s^2 2s^2 2p^3$.
- 6 **Explain** why hydrogen can form both a positive ion and a negative ion, and give the formula of each ion.
- 7 **Deduce** the name and give the formula for the ions of the following elements.
 - a Bromine
 - b Sulfur
 - c Barium
 - d Potassium
 - e Nitrogen
- 8 **Explain** why the formula of an ionic solid is an empirical formula and does not give the actual number of ions present in the compound.
- 9 **Identify** the names of the following ionic compounds.
 - a KI
 - b $BaCl_2$
 - c CaH_2
 - d PbO_2
 - e Na_3PO_4
 - f $Zn(NO_3)_2$
- 10 Write the formulas for the following ionic compounds.
 - a Magnesium nitride
 - b Aluminium oxide
 - c Mercury(II) sulfide
 - d Ammonium hydroxide
 - e Copper(I) carbonate
 - f Zinc phosphate
- 11 **Explain** why the formula of:
 - a sodium bromide is NaBr, not Na_2Br or $NaBr_2$
 - b aluminium fluoride is AlF_3 , not AlF or Al_3F or Al_2F_6 .
- 12 **Deduce** the valence of element X in each of the following compounds.
 - a X_2S_3
 - b $X(NO_3)_2$

13 Deduce the valence of Y in each of the following compounds.

- a MgY_2
- b $(\text{NH}_4)_3\text{Y}$

ANALYSING

14 The formulas (electrovalencies) of some common ions are shown in the following table.

Name	Formula	Name	Formula	Name	Formula	Name	Formula
Sulfate	SO_4^{2-}	Iron	Fe^{3+}	Sulfide	S^{2-}	Phosphate	PO_4^{3-}
Calcium	Ca^{2+}	Iodide	I^-	Lithium	Li^+	Barium	Ba^{2+}

Select from the ions listed in the table to answer the following questions.

- a Write a formula for an ionic compound that will have the formula XY_3 .
- b Write a formula for an ionic compound that will have the formula X_3Y_2 .
- c Write the formulas of two ionic compounds that have the formula XY .
- d Name two ions that have the same electron configuration.

5.3 Metallic bonding

Metals make up a large percentage of the periodic table. The valence electrons in metal atoms are loosely held, as shown by their low ionisation energies. To attain a **noble gas electron configuration**, metal atoms tend to lose electrons, which are negatively charged, to form a positively charged cation. This process requires energy.

In a metal, atoms are surrounded on all sides by other metal atoms, which also tend to lose electrons. These electrons that are detached from their atoms are called **delocalised electrons** and are able to move freely through regions of the metallic structure. When the electrons break away from their atom, they leave behind a positively charged cation. The delocalised electrons are free to move randomly between atoms within the three-dimensional arrangement and are often referred to as a 'sea of electrons' (Figure 5.3.1). The electrostatic force of attraction between the positively charged cations and the negatively charged delocalised electrons results in **metallic bonds** that hold the structure together.

noble gas electron configuration an arrangement of electrons that is identical to one of the elements in group 18, known as the noble gases

delocalised electrons electrons that are not held in a fixed orbit around one nucleus, but are shared by many atoms

metallic bond the electrostatic force of attraction between cations and the delocalised electrons that are free to move within the regular lattice structure of a metal

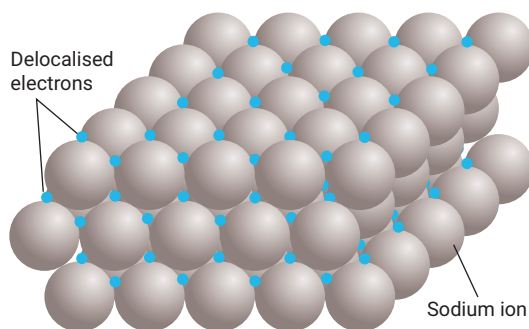


FIGURE 5.3.1 A model of a section of the metallic lattice of sodium showing the sodium cations in a 'sea of electrons'



Weblink
Metallic bonding

Metallic bonds are non-directional because the electrons are free to move between cations. This movement of electrons explains why metals can conduct electricity when in the solid state.

Overall, a piece of metal is uncharged because the total positive charge of the cations is equal to the total negative charge of the delocalised electrons.

LEARNING CHECK 5.3

DESCRIBING

- 1 **Identify** the major characteristics of the valency electrons in a metal.
- 2 **Define** 'metallic bonding'.

APPLYING

- 3 **Explain** why metals have delocalised electrons and why you should not stick a metal fork into a power socket.
- 4 The following passage describes the formation and structure of a metal. **Identify** the correct terms from the list below to complete the following passage.

free	positive	ions	fixed
valence	attraction	electrons	lattice
delocalised	metal	ionised	

The metallic structure consists of a regular array of ___a___ metal ___b___ surrounded by a sea of ___c___ electrons. The metal ions remain in ___d___ positions forming a lattice structure while the valence electrons are ___e___ to move throughout the ___f___ of positive metal ions.

In the formation of the metallic structure, metal atoms lose their ___g___ electrons and are said to be ___h___. This ionisation process forms positive ___i___ ions and free ___j___. The stability of the metallic structure comes from the strong ___k___ between the positive metal ions and the sea of electrons.



Worksheet
Ionic and covalent bonds

covalent bonding

bonding between two or more atoms where electron pairs are shared between the valence shells

bonding pair a pair of electrons shared between two atoms

5.4 Covalent bonding

Like all elements in the periodic table (except the noble gases), non-metals form bonds to attain a full valence shell. They can achieve this by accepting an electron from a metal to form an ionic bond. However, non-metal atoms can also share electrons to attain a filled valence shell. The electrostatic force of attraction between the shared electrons and the positively charged nuclei holds the atoms together and is called **covalent bonding**. Covalent bonding is also referred to as 'directional bonding' because there is a direct line between the bonded particles. An example of covalent bonding is shown for a hydrogen molecule in **Figure 5.4.1**.

Consider the hydrogen molecule in Figure 5.4.1. Each hydrogen atom has one valence electron, so it needs one more to fill its valence shell. When two hydrogen atoms come close enough together, the single electron of each atom is attracted to both nuclei and drawn into the space between them. The electron shells of the individual atoms begin to merge, and the two nuclei start to share the two electrons. Each hydrogen atom can now be considered to have two electrons and a filled valence shell. This shared pair is called a **bonding pair**.

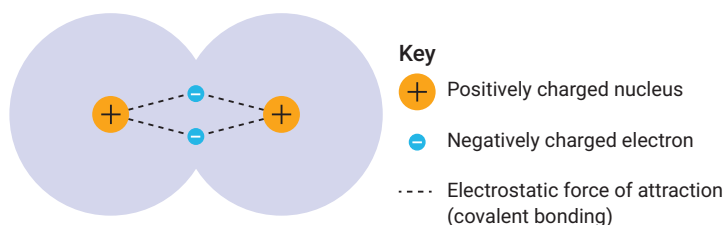


FIGURE 5.4.1 Covalent bonding in a molecule of hydrogen (H_2)

Covalent bonds typically form between atoms with a small difference in electronegativity because these atoms tend to share electrons relatively equally, leading to a more stable bonding situation. For example, the difference in electronegativity between carbon and oxygen in the carbon monoxide molecule is 0.8. In this case, the electrons are shared, but they are slightly more attracted to the more electronegative oxygen atom. This sharing of electrons enables both atoms to achieve more stable electronic configurations. This will be further covered in Chapter 11.

Representing covalent bonds

A covalent bond can be represented in different ways (**Figure 5.4.2**). Fluorine forms the diatomic molecule F_2 by sharing one pair of electrons. We can use a Lewis structure to show shared and unshared electrons. The pairs of unshared valence electrons are called **non-bonding pairs** or **lone pairs**. Another way to show one pair of shared electrons is by a dash. One dash means one shared (bonding pair) of electrons or one covalent bond. The valence structure shows the non-bonding pairs of electrons and the bonding pair as a dash. The structural formula leaves off the non-bonding pairs and just shows the dash between the elements.

Lewis structure	Valence structure	Structural formula
Key • Valence electron	•• Non-bonding pair — Bonding pair	F Nucleus, inner-shell electrons and non-bonding pairs

FIGURE 5.4.2 Representing covalent bonds

Covalent bonds and the periodic table

Covalent bonding occurs when both elements forming a given substance need to gain electrons to attain a noble gas configuration, such as H_2 , H_2O and CO_2 .

Hydrogen, metalloids and non-metals, except the noble gases, tend to form covalent bonds. The number of covalent bonds an atom forms is given by the valency of the element, which is directly related to its position in the periodic table. **Figure 5.4.3** demonstrates how, in the Lewis structure formulas of some simple molecules, the number of covalent bonds of the main atom corresponds with the number of electrons needed to attain a noble gas configuration.



Weblink
Covalent bonding game

Worksheet
Bonding crossword

non-bonding or lone pair electrons a pair of electrons located on an individual atom, which are not shared



Worksheet
Comparing chemical bonding

Valence electrons	H•	•B•	•C•	•N•	•O•	•F•
Compound with H ×	H×H	$\begin{array}{c} \text{H} \\ \times \\ \text{B} \times \text{H} \\ \times \\ \text{H} \end{array}$	$\begin{array}{c} \text{H} \\ \times \\ \text{H} \times \text{C} \times \text{H} \\ \times \\ \text{H} \end{array}$	$\begin{array}{c} \text{H} \\ \times \\ \text{N} \times \text{H} \\ \times \\ \text{H} \end{array}$	$\begin{array}{c} \text{O} \times \text{H} \\ \times \\ \text{H} \end{array}$	$\begin{array}{c} \text{F} \times \text{H} \end{array}$
Number of covalent bonds of the main atom	1	3	4	3	2	1

Key • Single electron × Contribution to the bonding pair

FIGURE 5.4.3 Lewis structures of some simple molecules

Multiple covalent bonds

multiple covalent bond a bond formed when two atoms share more than one pair of electrons; a double bond consists of two shared pairs of electrons and a triple bond of three shared pairs

double covalent bond a covalent bond consisting of two shared pairs of electrons

triple covalent bond a covalent bond consisting of three shared pairs of electrons

Sometimes, when atoms come together, they share more than one pair of electrons. When this occurs, a **multiple covalent bond** is formed.

Consider what happens when an oxygen atom bonds to another oxygen atom to form O_2 or when a nitrogen atom bonds to another nitrogen atom to form N_2 .

Each oxygen atom has a valency of 2, so it can form two covalent bonds. When two oxygen atoms come together, two pairs of electrons are drawn into the region between the two nuclei and their outer shells partly merge. This means there are four bonding electrons holding the nuclei closely together, so this will be a stronger bond than that formed by one pair. Because two pairs of electrons are shared, the bond is called a **double covalent bond**.

Nitrogen has a valency of 3, so it can form three covalent bonds. When two nitrogen atoms come together, they share three pairs of electrons and form a **triple covalent bond**. This bond is even stronger than a double bond and is the main reason why nitrogen gas is relatively unreactive, because a lot of energy is needed to break the triple bond in a chemical reaction. **Figure 5.4.4** shows how double bonds (e.g. in O_2) and triple bonds (e.g. in N_2) can be represented.

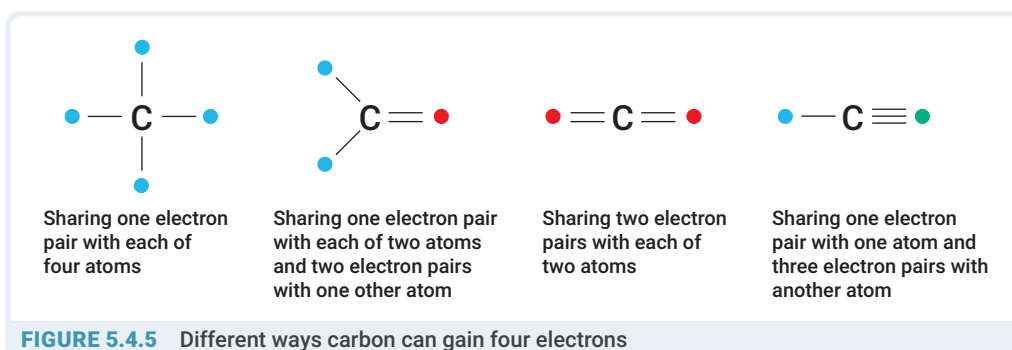
Molecular formula	Lewis structure	Valence structure	Structural formula
O_2	$\begin{array}{c} \bullet \bullet \bullet \bullet \\ \text{O} \bullet \bullet \bullet \text{O} \\ \bullet \bullet \bullet \bullet \end{array}$	$\begin{array}{c} \bullet \bullet \bullet \bullet \\ \text{O} = \text{O} \\ \bullet \bullet \bullet \bullet \end{array}$	$\text{O} = \text{O}$
N_2	$\begin{array}{c} \bullet \bullet \bullet \bullet \\ \text{N} \bullet \bullet \bullet \text{N} \\ \bullet \bullet \bullet \bullet \end{array}$	$\begin{array}{c} \bullet \bullet \bullet \bullet \\ \text{N} \equiv \text{N} \\ \bullet \bullet \bullet \bullet \end{array}$	$\text{N} \equiv \text{N}$

FIGURE 5.4.4 Different ways of representing double and triple covalent bonds



Weblink
Single and multiple
covalent bonds

Carbon, in group 14, has a valency of 4 so can form four covalent bonds. It cannot form a quadruple covalent bond because this type of bond cannot exist. A quadruple covalent bond would require four pairs of electrons to be located between the two nuclei. This cannot happen because the repulsive forces between the negative electrons becomes greater than the attractive forces between the electrons and the nuclei, preventing a fourth pair of electrons in the bond. Carbon can achieve its four bonds in many ways, as shown in **Figure 5.4.5**.



Names and formulas of covalent molecules

As you have already learnt, the valency of an element is an indication of the number of bonds it can form. Generally, this can be deduced from the periodic table as shown in Table 3.2.1. However, as with the transition metals, several non-metallic elements can display more than one valency in covalent compounds. For example, in carbon monoxide (CO), carbon has a valency of 3 and forms a triple bond with oxygen, whereas in carbon dioxide (CO₂), carbon has a valency of 4 and forms two double bonds with two oxygen atoms. For this reason, the names of covalent compounds need to indicate the valency. The IUPAC convention for naming covalent compounds differs from the naming of ionic compounds.

The following rules apply to covalent compounds made up of only two elements.

1. Use the element name for the first element and change the end of the name of the second element to 'ide'; for example, hydrogen fluoride.
2. Name the element that is further to the left in the periodic table first.
3. If both elements are in the same group, name the element lower down the group first; for example, silicon carbide.
4. An exception to rules 2 and 3 is when oxygen is bonded to Cl, Br or I. In these compounds, oxygen is named last. The rationale is that oxygen is more electronegative than Cl, Br or I.
5. Indicate the number of atoms of each type by using prefixes in front of each part of the name (although 'mono' may be omitted from the first named element). For example, CO is carbon monoxide, CO₂ is carbon dioxide and N₂O₅ is dinitrogen pentoxide. **Table 5.4.1** provides a list of the first 10 prefixes used for naming covalent compounds.

TABLE 5.4.1 Naming covalent molecules

Prefix	Mono	Di	Tri	Tetra	Penta	Hexa	Hepta	Octa	Nona	Deca
Number of atoms	1	2	3	4	5	6	7	8	9	10

Once the name of the compound is given, writing the formula is straightforward. The order is given by the name, and the prefix of the element becomes the subscript for that element in the formula. For example, diphosphorus pentoxide is P₂O₅ and sulfur hexafluoride is SF₆.

Sometimes compounds are referred to by their common names rather than their systematic names. For example, H₂O, dihydrogen monoxide, is called water; NH₃, nitrogen trihydride, is called ammonia; and CH₄, carbon tetrahydride, is called methane.

The **molecular formula** of a compound represents the actual number of atoms of each element that are combined in each molecule of a compound. Molecular formulas can be simplified to empirical formulas. For example, glucose has the molecular formula C₆H₁₂O₆ and the empirical formula CH₂O. It is important to use the molecular formula for covalent compounds because some may have the same empirical formula or their molecular formula is the empirical formula. For example, methanal has the molecular formula of CH₂O.

molecular formula a chemical formula that gives the total number of atoms of each element in a molecule

LEARNING CHECK 5.4

DESCRIBING

- 1 **State** the method used to determine which element is written first in a covalent compound.
- 2
 - a **Define** 'covalent bond'.
 - b **Identify** how many electrons share in the formation of a:
 - i single bond
 - ii double bond
 - iii triple bond.
- 3 Consider the molecule CH_3Cl . **Identify** how many electrons there are in the valence shell of each:
 - a carbon atom
 - b hydrogen atom
 - c chlorine atom.
- 4 **Determine** the molecular formulas for the covalent molecules formed from:
 - a oxygen and fluorine
 - b boron and hydrogen.
- 5 **Explain** why the formula for methane is CH_4 , not CH_3 or CH_5 .

APPLYING

- 6 Write the formulas for:
 - a sulfur dioxide
 - b dinitrogen pentoxide
 - c carbon tetrachloride
 - d nitrogen trifluoride
 - e silicon tetrabromide.
- 7 Name the following compounds.
 - a N_2O
 - b NCl_3
 - c SO_3
 - d H_2S
 - e N_2O_4
- 8 **Draw** Lewis structure formulas and structural formulas for the following substances.
 - a Br_2
 - b HF
 - c NF_3
 - d CS_2
 - e HCHO
- 9 **Identify** the bonding and non-bonding electron pairs in SiH_3 .



ANALYSING

10 Three Lewis structures are shown below.

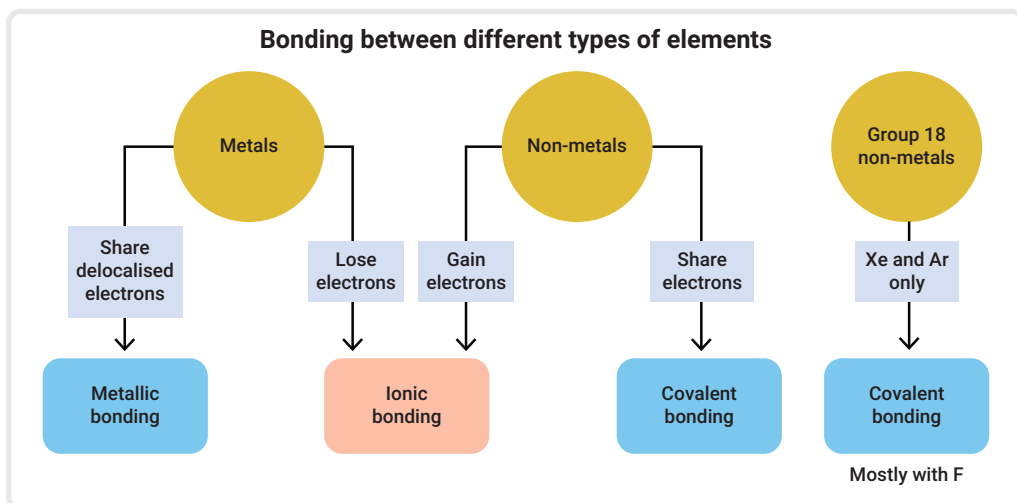


Copy and complete the following table by answering the questions by referring to the three structures.

	Compound A	Compound B	Compound C
To which group in the periodic table does X belong?			
Suggest an identity for X.			
Suggest a possible molecule with this structure.			

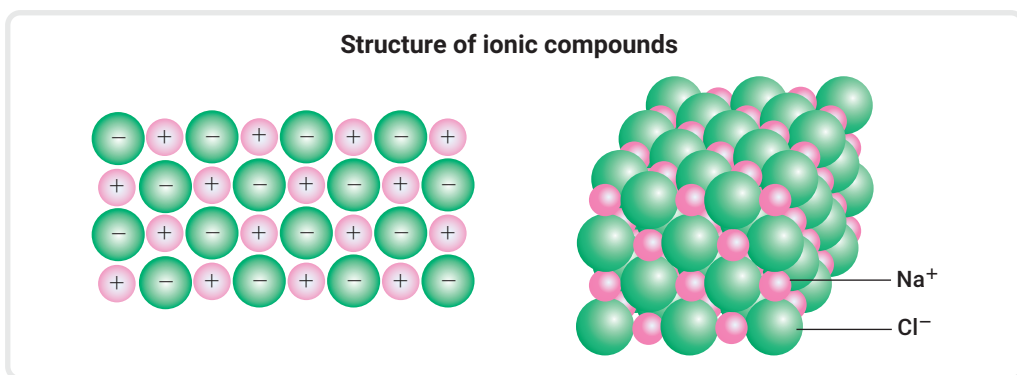
CHAPTER SUMMARY

Types of bonding



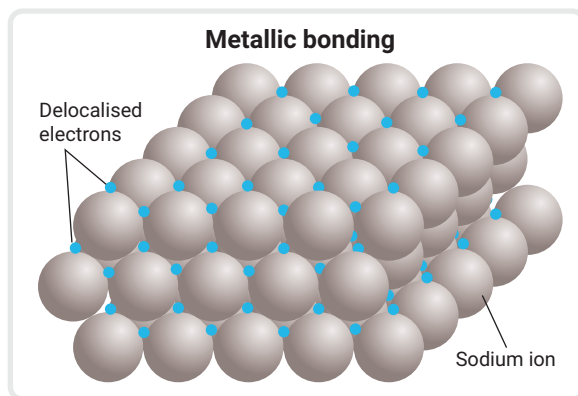
Ionic bonding

- Ionic bonding occurs between a metal and a non-metal and results from the electrostatic attraction of a cation and an anion.



Metallic bonding

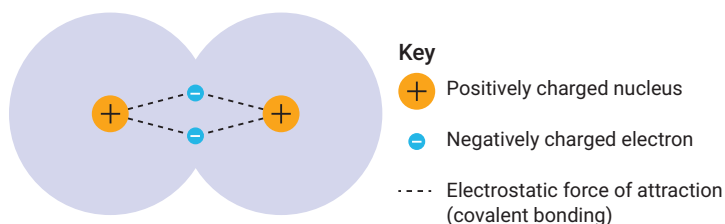
- Metallic bonding occurs between metals.
- A metallic bond is the attraction of metal cations to the surrounding delocalised electrons.



Covalent bonding

- Covalent bonding occurs between two non-metals and results from the sharing of electrons between non-metallic atoms.
- Valency is a measure of the number of bonds that an atom can form.

Covalent bonding between two atoms



- Lewis structures show the valence electrons of atoms or ions. This can be used to identify the numbers of bonding and non-bonding (lone) pairs of electrons around each atom in a molecule.

Lewis structure showing valence electrons



Key • Valence electron

Naming metallic, ionic and covalent compounds

Naming and writing chemical formulas

Metal	Ionic compound	Covalent compound
- If made up of one type of metal, it will take the name of the metal. - If made up of more than one type of metal, it is an alloy.	- Cation: keeps name of the element. - Anion: non-metal name ends in 'ide'. - The total charges of the cations and anions in the ionic formula need to equal zero. - Always use empirical formula. - Transition metals use Roman numerals in their names. - Polyatomic ions use their names. - Polyatomic ions use parentheses if more than one ion in chemical formula. - Never use prefixes.	- Keep the name of the first element. - Always use prefixes. - mono (never on first element) - di - tri - tetra - penta - hexa - hepta - octa - nona - deca - Second element name always end in 'ide'. - Never use empirical formula unless asked to.

CHAPTER EXAM

MULTIPLE CHOICE

- A particular element has three electrons in its valence shell. This element will form:
 - a metallic lattice when pure and covalently bonded molecules when combined with chlorine.
 - an ionic lattice when pure and a covalent lattice when bonded to chlorine.
 - an ionic lattice when pure and covalently bonded molecules when combined with chlorine.
 - a metallic lattice when pure and an ionic lattice when combined with chlorine.
- Identify** the maximum number of covalent bonds that carbon will form.
 - 1
 - 2
 - 3
 - 4
- Identify** which of the following substances does not contain ions that are free to move.
 - NaCl(aq) – a solid dissolved in water
 - Na(s)
 - NaCl(l)
 - Na(l)
- In the Lewis structure of the OF_2 molecule, what is the number of lone pairs of electrons around the central oxygen atom?
 - 1
 - 2
 - 3
 - 4
- Identify** which particle is a negative ion.

	Electrons	Neutrons	Protons
A	9	10	9
B	10	14	12
C	16	16	16
D	18	18	16

- Determine** which of the following correctly represents the ions that are in the ionic compound whose formula is $\text{Fe}_2(\text{SO}_4)_3$.
 - Fe^{2+} and SO_4^{2-}
 - Fe^{3+} and SO_4^{2-}
 - Fe^{2+} and SO_4^{3-}
 - Fe^{2+} and SO_3^{2-}
- Element Q forms a compound with the formula $\text{Q}_2(\text{SO}_4)_3$. What is the formula for the compound that Q forms with chlorine gas?
 - Q_3Cl_2
 - Q_3Cl
 - QCl_3
 - Q_2Cl_3

8. Atom X has six valence electrons. Atom Y has seven valence electrons. What is the bonding type and most likely formula for the compound when the two atoms react?

	Bonding type	Formula
A	Ionic	X_2Y
B	Ionic	XY_2
C	Covalent	X_2Y
D	Covalent	XY_2

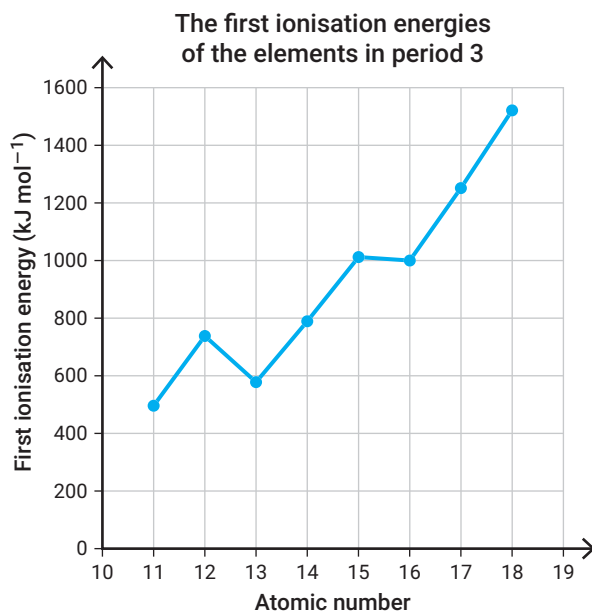
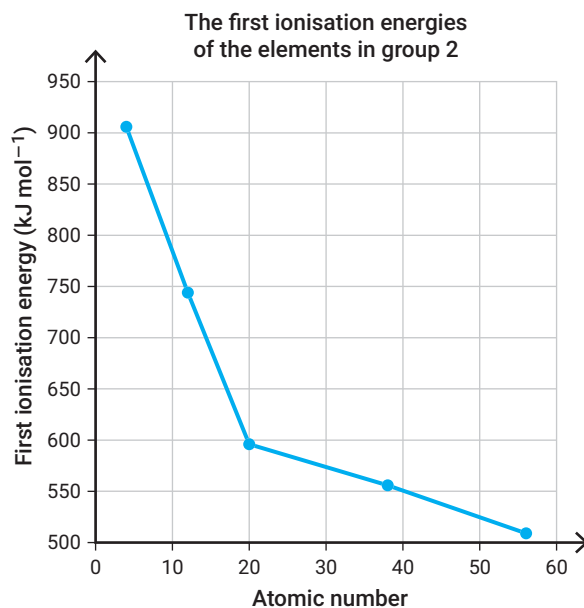
9. When the ionic compound Ba_3P_2 is formed:
- A six electrons are shared.
 - B six electrons are moved from the barium atoms to the phosphorus atoms.
 - C six electrons move from the phosphorus atoms to the barium atoms.
 - D three electrons move from the barium atoms to the phosphorus atoms.
10. The element X reacts with oxygen to form the compound X_2O_3 , whereas element Y reacts with hydrogen to form the compound H_2Y . What is the most likely formula of a stable compound formed by element X and element Y?
- A XY
 - B X_2Y
 - C X_2Y_3
 - D X_3Y_2

SHORT RESPONSE

11. Determine the names of the following compounds.
- a Ag_2O
 - b N_2O_3
 - c SF_4
 - d NH_4Br
 - e $Mg(OH)_2$
 - f NO
12. Write the correct formula for each of the following compounds.
- a Boron trichloride
 - b Calcium carbonate
 - c Silicon dioxide
 - d Dinitrogen tetroxide
 - e Potassium sulfide
 - f Lead(IV) oxide

CROSS-CHAPTER QUESTION

13. The following graphs show the first ionisation energies of some elements of the periodic table.



- a Compare** the first ionisation energies of the elements in group 2 with the first ionisation energies of the elements in period 3.
- b Describe** the relationship between ionisation energy and the likelihood of a metal ion taking part in ionic bonding.

DATA ANALYSIS

14. Analyse data

The ionisation energies of some metals are shown below.

Element	First ionisation energy (kJ mol^{-1})
Lithium (Li)	526
Sodium (Na)	502
Magnesium (Mg)	744
Aluminium (Al)	584

Based on these values, **identify** the metal that would most easily take part in ionic bonding with chlorine.

15. Analyse data

A scientist wanted to determine the nature of substances in a mixture. They conducted multiple tests, and identified the charges of four different substances. The substances were found to be monatomic ions.

Substance	Ion charge
P	+1
Q	-1
R	-2
S	+2

- a Identify** which substance(s) can take part in covalent bonding.
- b Identify** which substance(s) can take part in metallic bonding.
- c** Write the formula for ionic compounds that could be formed if the above substances were able to react.

SCIENCE AS A HUMAN ENDEAVOUR

Syllabus dot point

Explore the use of radioisotopes for carbon dating and radiotherapy and Marie Curie's contribution to research on radioactivity.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Exploring radioisotopes and radioactivity

It's hard to imagine not having an X-ray to determine if your arm is broken or to check for tooth decay. The term 'radioactivity' and the use of radioactivity for medical diagnosis and treatment can be attributed to the famous scientist Marie Curie.

Marie Curie

Marie Curie (1867–1934) is recognised as one of the greatest scientists of her time. She and her husband, Pierre Curie, were pioneers in the field of radioactivity. Marie developed methods to extract elements from a pitchblende mixture and after months of tiring work, discovered the element polonium. Later that same year, the Curies announced the discovery of another element, which they named radium. To describe the behaviour of elements that produced invisible rays, Marie invented the word 'radioactivity', which is based on the Latin word for 'ray'.

While investigating uranium, Marie identified that the quantity of uranium was directly related to its radioactivity. This led her to hypothesise that the radiation must come from the atom itself. This hypothesis was important in advancing the model of the atom. Ernest Rutherford built on the work of the Curies and proved that an atom was divisible.

Marie pioneered the use of radium as a treatment for cancer. She also developed mobile radiography units and participated in their use in field hospitals during World War I.

Marie and Pierre, along with Henri Becquerel, were awarded the Nobel Prize in Physics in 1903 for their work on radioactivity. In 1911, Marie Curie was awarded the Nobel Prize in Chemistry for the discovery of the elements radium and polonium,

for the isolation of radium, and for the study of the nature and compounds of this element. She was the first woman to be awarded a Nobel Prize and the first person to be awarded two Nobel Prizes. Today, only four people have been awarded two Nobel Prizes.

Medical use of radioisotopes

In medicine, **radioisotopes** are used in two ways – to diagnose disease and in **radiotherapy**, to treat specific illnesses. When an isotope is ingested by a patient, the radiation it emits produces a map of where it ends up in the body, revealing areas of unusual structure or activity, which can be an indicator of disease.

One of the most common radioactive isotopes used for diagnosis is technetium-99. This isotope has a short half-life of about 6 hours, which means technetium-99 breaks down very quickly and does not remain in the body. Its many uses include imaging the kidneys, liver, lungs and brain. The isotope attaches to a specific pharmaceutical (a compound used as a medical drug) and is then transported to the required place in the body, where it decays, emitting radiation that is detected and measured.

Ewing Galloway/Alamy Stock Photo



FIGURE 1 Marie Curie was a pioneer in the field of radioactivity.

radioisotope a radioactive isotope of an element that is unstable and emits radiation as it decays

radiotherapy the use of radiation to kill cancer cells and shrink tumours

This is usually analysed by computer to produce images to help doctors diagnose their patients (**Figure 2**).

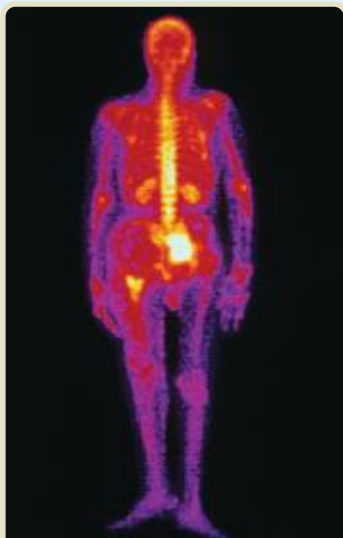


FIGURE 2 This image, produced by the radiation emitted by technetium-99, shows thyroid cancer that has spread (metastasised).

Radioisotopes are used in the treatment of diseases, especially cancer. This can be done either internally or externally. External treatment involves aiming a beam of radiation at the specific part of the body where the cancer is located. Internal treatment involves ingesting a drink or tablets containing the radioisotope. Cancer cells divide very rapidly so they absorb the radioisotope at high levels. The radiation emitted by the isotope as it decays damages the cells.

Carbon dating

Carbon dating (also referred to as carbon-14 dating or radiocarbon dating) is a method used to determine the age of an object that was once living.

All living things are made of carbon-based compounds. The main isotope of carbon, C-12, is stable, but the C-14 isotope is radioactive. Over time, C-14 undergoes radioactive decay as one of its neutrons is converted to a proton and a negatively charged **beta particle** is emitted. The new nucleus has seven protons and seven neutrons, so C-14 has been converted into nitrogen, N-14.

This knowledge is used by scientists to carbon date archaeological samples of biological origin, such as bone and wood. Plants and animals incorporate both isotopes into their cells; therefore, both isotopes are maintained in living organisms in about the same proportion as in the atmosphere. When an organism dies, the carbon-14 in its tissues decays to nitrogen-14 at a known rate and is no longer replaced. By comparing the ratio of carbon-12 to carbon-14 in a sample, scientists can determine the age of an object or a fossil (**Figure 3**).

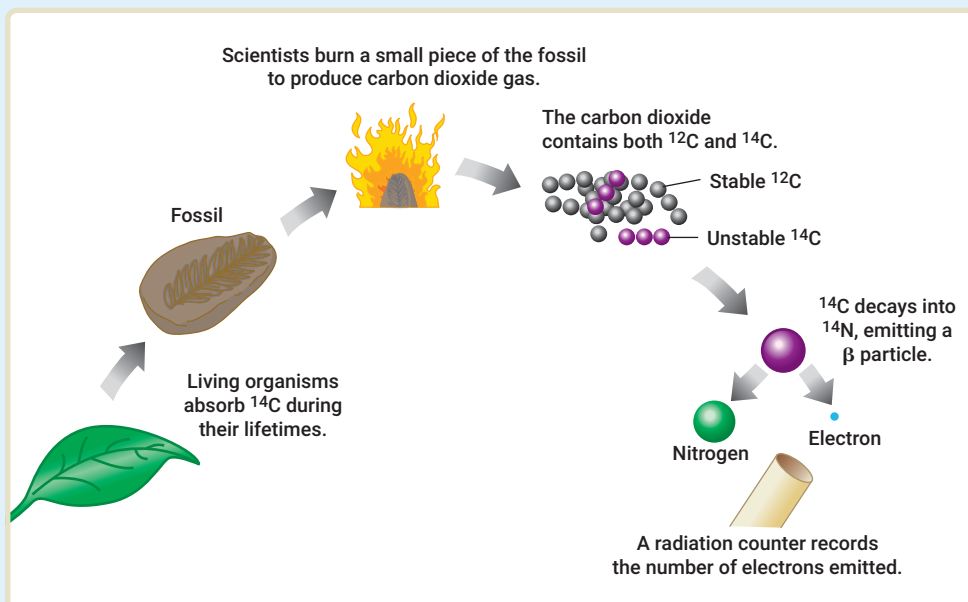


FIGURE 3 Carbon dating. Carbon-14 decays at a known rate. Scientists can determine the age of a fossil by measuring the amount of carbon-14 it contains.



Weblink
Using radioisotopes

carbon dating a method used to determine the age of a once-living object by comparing the ratio of carbon-12 to carbon-14 in the object

beta particle an electron or positron (which has the same mass as an electron but the opposite charge) emitted when an unstable nucleus decays

Compounds and mixtures



leomagdala/Shutterstock.com

SYLLABUS DOT POINTS

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.

SCIENTIFIC INQUIRY

Investigate:

- the separation of mixtures based on physical properties.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Pure substances may be either elements or compounds and they have distinct physical and chemical properties. Elements cannot be separated into simpler substances. Compounds can be separated into simpler substances by chemical processes but cannot be separated by physical processes. Mixtures can be separated into simpler substances by physical processes. The physical process used depends on the part or parts of the mixture that needs to be kept.

Practical

- Extracting water from copper(II) sulfate solution

Worksheets

- Separating methods
- Physical separation techniques
- Separating techniques crossword
- Properties of pure substances and mixtures



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Elements are made up of one type of atom.
- ✓ Compounds are made up of two or more different elements chemically combined in a fixed ratio.
- ✓ Mixtures are made up of two or more different substances that are not chemically combined.
- ✓ Distillation is one of the many techniques that can be used to separate mixtures.
- ✓ Physical properties are observable characteristics of a substance.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ identify whether a substance is pure or a mixture
- ✓ use the properties of homogeneous and heterogeneous mixtures to distinguish between the two types of mixtures
- ✓ relate separation techniques to properties
- ✓ identify measurable properties of pure substances
- ✓ describe how changing the amounts of substances in a mixture affects its properties
- ✓ compare the properties of pure substances and mixtures
- ✓ interpret data to determine whether a sample is a pure substance or a mixture.

pure substance
composed of only
one substance

6.1 Classifying substances

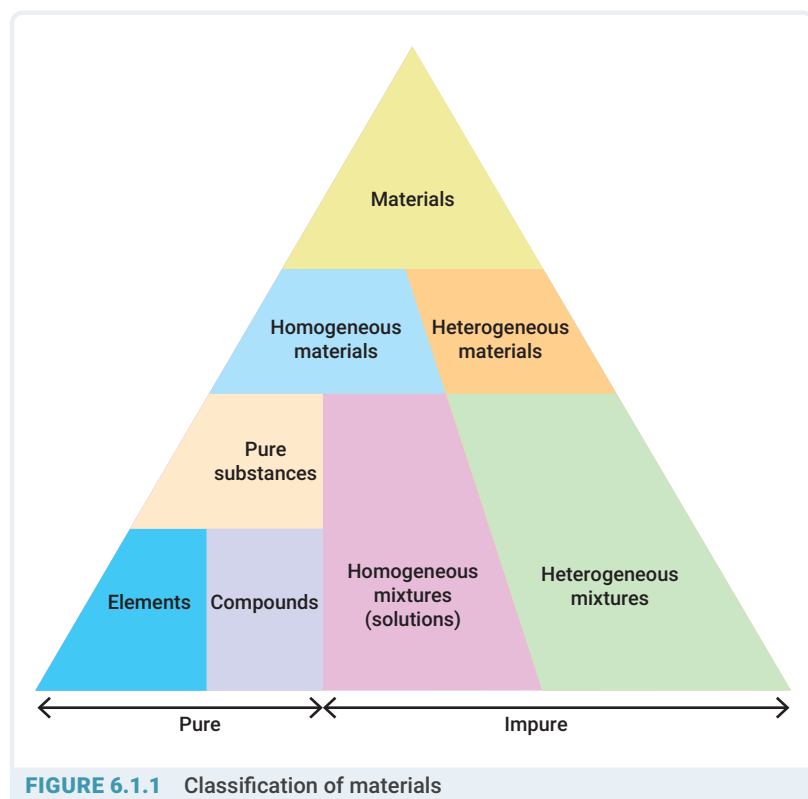


FIGURE 6.1.1 Classification of materials

Materials can be organised into two main categories: pure substances and mixtures. **Pure substances** are composed of only one substance and can be either elements or compounds. Although elements cannot be broken down into simpler substances, compounds can be broken down into their constituent elements using chemical processes. As highlighted previously, elements are considered to be fundamental building blocks of all materials (Figure 6.1.1).

Because every sample of a pure substance has the same composition, under the same conditions, pure substances exhibit the same physical and chemical properties. Some of these physical properties are:

- melting and boiling points
- electrical and thermal conductivity
- strength
- density
- solubility.

Some chemical properties are:

- reactivity
- stability
- toxicity
- flammability.

Throughout human history, we have used substances initially because of their physical properties, without always being aware of their chemical properties. For example, early airships were filled with hydrogen because of its low density. Unfortunately, the high flammability of hydrogen was not taken into consideration and this contributed to the *Hindenburg* disaster on 6 May 1937 at the Naval Air Station, Lakehurst, New Jersey, United States (Figure 6.1.2).

Mixtures are a combination of two or more pure substances. The pure substances within a mixture retain their distinct properties and, as a result, mixtures exhibit varying physical and chemical properties depending on their composition. Unlike pure substances, mixtures can be separated by a variety of different physical separation techniques – some of which will be described later in this chapter. Table 6.1.1 summarises the key similarities and differences between pure substances and mixtures.



FIGURE 6.1.2 The *Hindenburg* disaster occurred because early airships were filled with light, but flammable, hydrogen.

Alamy Stock Photo/Science History Images

mixture a combination of two or more substances



Weblink
Pure substances and mixtures

TABLE 6.1.1 Differences between pure substances and mixtures

Feature	Pure substance	Mixture
Composition	Made of one substance and therefore has a fixed composition	Made up of many different substances, i.e. has variable composition
Properties	Constant	Vary depending on the composition
Separation	Elements cannot be separated into different substances by physical or chemical processes, but compounds can be chemically separated into their constituent elements.	Can be separated by physical processes. Chemical separation results in compounds different from the original composition.
Examples	Gold, water, diamond	Milk, air, sand

LEARNING CHECK 6.1

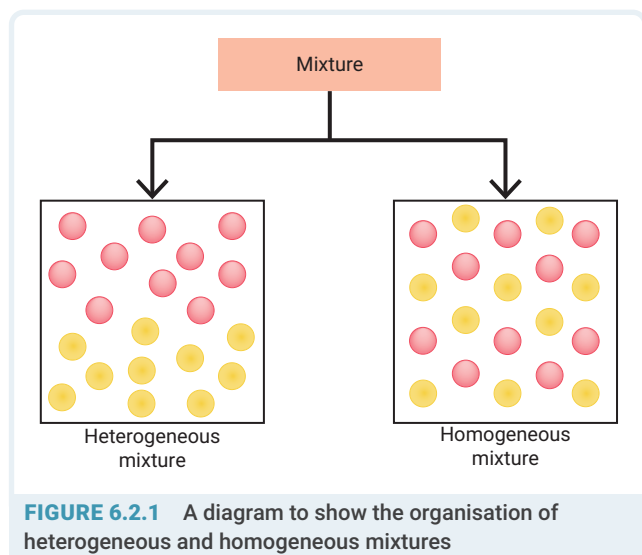
DESCRIBING

- 1 A student made the following statement about a pure substance:
'A pure substance can only be made from one type of element.'
 - a **Identify** the error in the statement.
 - b Rewrite the statement so that it accurately describes a pure substance.
- 2 **Identify** two features that can be used to differentiate a mixture from a pure substance.

APPLYING

- 3 A sample of substance A is known to have two different elements present. Upon further analysis, it was noted that the properties of the substance were consistent throughout all samples. **Classify** this substance as a pure substance or a mixture. **Justify** your answer.
- 4 A student was given two samples – one sample of pure gold and one sample of beach sand. **Describe** the differences in how easy it would be to separate each sample into its individual components.

6.2 Heterogeneous and homogeneous mixtures



homogeneous mixture
a mixture in which substances are uniformly distributed throughout

heterogeneous mixture
a mixture in which substances are non-uniformly distributed throughout



Syllabus link
Solutions are covered in further detail in Chapter 14.



Weblink
How to identify heterogeneous and homogeneous mixtures

As shown in Figure 6.1.1, mixtures can be further classified depending on how the particles in the mixture are distributed. In **homogeneous mixtures**, the particles within the mixture are uniformly distributed whereas in **heterogeneous mixtures**, particles are non-uniformly distributed (**Figure 6.2.1**).

Solutions are one of the most common types of homogeneous mixtures, particularly those that have liquid solvents. Solids, liquids and gases can be dissolved in liquid solvents to produce solutions. For example:

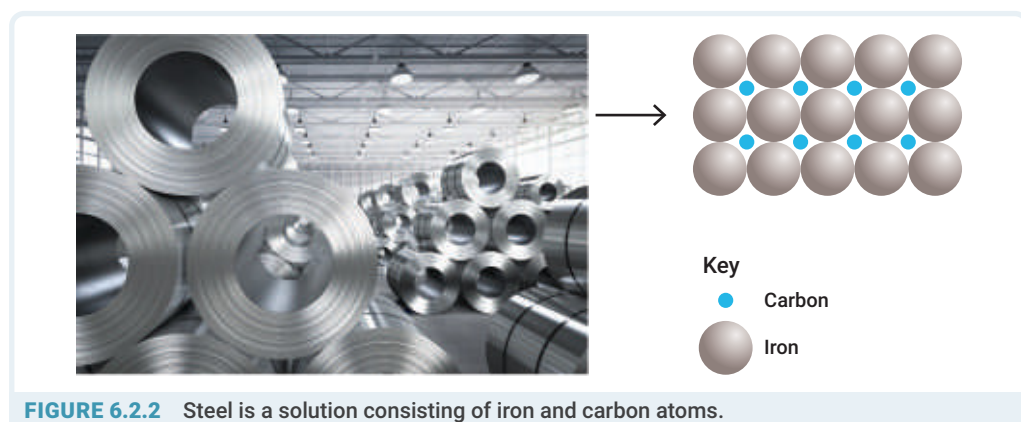
- salt water is a liquid solution in which solid salt is dissolved in water
- rubbing alcohol, sometimes used in cleaning products, is a solution in which liquid alcohol is dissolved in water
- soda water is a solution in which a gas, carbon dioxide, is dissolved in water.

Although liquid solutions are most widely recognised, solutions also occur in various phases. For example:

- air is a gas solution that is a homogeneous mixture of different gases such as nitrogen, carbon dioxide and oxygen
- steel is a solid solution that is a homogeneous mixture of iron and carbon (**Figure 6.2.2**).

Like homogeneous mixtures, heterogeneous mixtures can be found in various forms. For example:

- granite is a type of rock that is composed of different minerals including mica and quartz
- concrete is a mixture of substances including water, sand and cement.



Distinguishing between homogeneous and heterogeneous mixtures

It is sometimes possible to identify whether a mixture is homogeneous or heterogeneous by looking at the arrangement and relative size of the particles. Homogeneous mixtures appear uniform throughout, meaning that they have a consistent composition regardless of the

section of the mixture being observed. For example, salt water looks the same throughout. In contrast, the components of heterogeneous mixtures are often visible and distinguishable from one another because they are not evenly distributed. For example, heterogeneous mixtures like salad dressing can be visually identified by the presence of distinct phases or visible particles that do not mix uniformly (Figure 6.2.3).

Some heterogeneous mixtures such as milk can appear to be uniform like homogeneous mixtures. However, upon closer inspection with a microscope, the non-uniformity of milk becomes more obvious (fat globules suspended in serum) (Figure 6.2.4).

It can be difficult to determine whether a mixture that appears transparent is homogeneous or heterogeneous. In these instances, we can make use of the Tyndall effect. When light is shone through a heterogeneous mixture, a beam of light will be visible within the mixture as it is reflected off particles within the mixture. This is not observable in a homogeneous mixture (Figure 6.2.5).



FIGURE 6.2.3 Salad dressing is a heterogeneous mixture in which the components are visually distinguishable.

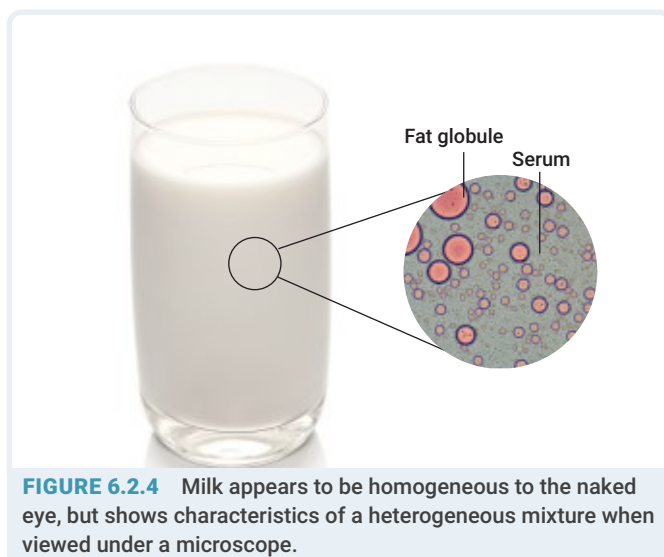


FIGURE 6.2.4 Milk appears to be homogeneous to the naked eye, but shows characteristics of a heterogeneous mixture when viewed under a microscope.

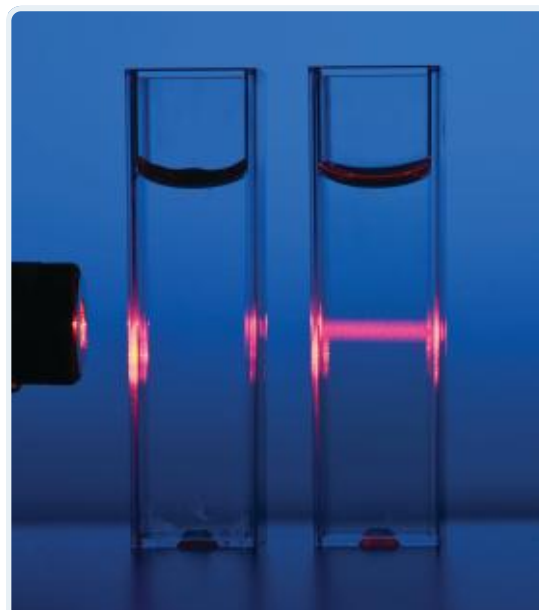


FIGURE 6.2.5 The Tyndall effect in homogeneous and heterogeneous mixtures. The tube on the left shows light passing through a homogeneous mixture, whereas a visible beam is evident in the heterogeneous mixture on the right.

Table 6.2.1 summarises the key similarities and difference between homogeneous and heterogeneous mixtures.

TABLE 6.2.1 Summary of the differences between homogeneous and heterogeneous mixtures

Feature	Homogeneous mixture	Heterogeneous mixture
Appearance	Substances are not visually distinguishable from each other and appear uniform throughout.	Substances are often visually distinguishable.
Composition	Substances are uniformly distributed.	Substances are unevenly distributed.
Examples	Cleaning solutions, steel, salt water	Blood, milk, muddy water

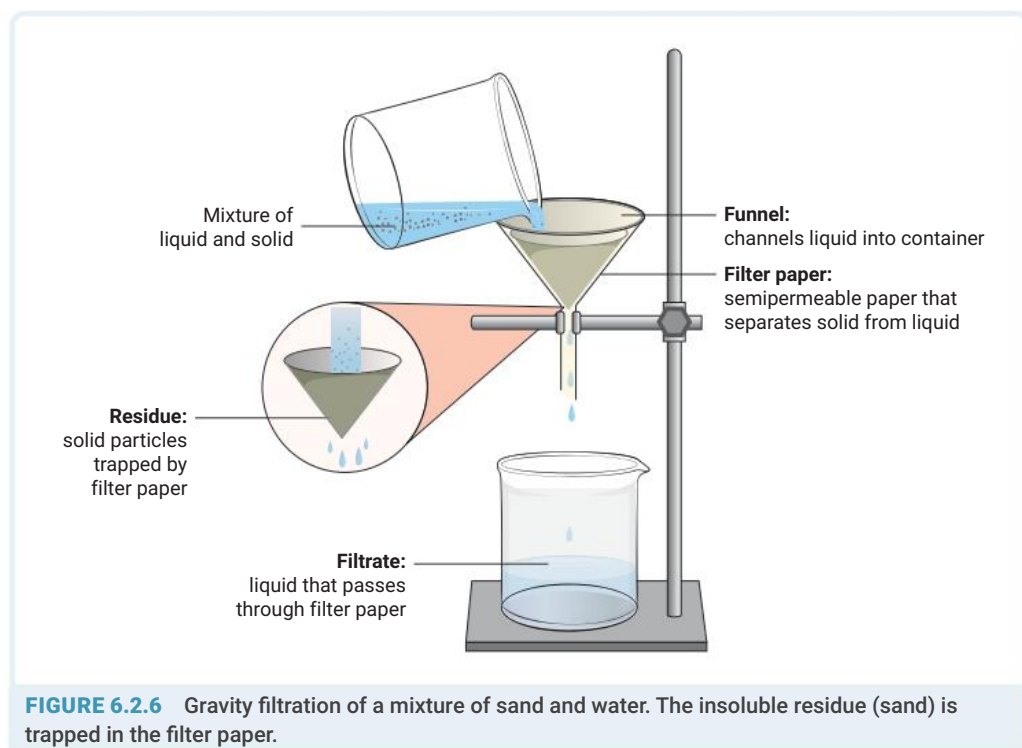
Separating mixtures

Unlike pure substances, mixtures can be separated by physical processes into pure substances or even simpler mixtures. Physical separation techniques make use of the differences in properties of the substances present in the mixture.

A heterogeneous mixture of two solids, such as sand and table salt, can be separated by particle size. This could involve a simple technique called **sieving**, where the mixture is poured through a sieve and the smaller particles pass through while the larger particles are trapped in the sieve. Sieving is often used in cooking. The mixture could also be separated by dissolving the substance in water and using **gravity filtration** (Figure 6.2.6). In the example shown, the insoluble sand (residue) will be left in the filter paper and the soluble table salt will travel through the filter paper into the filtrate.

sieving a technique that separates particles in a mixture based on particle size

gravity filtration a technique that uses gravity to separate a solid–liquid mixture by passing the mixture through filter paper



distillation a separation technique of a liquid mixture based on boiling point; the mixture is heated and the vapour produced is condensed

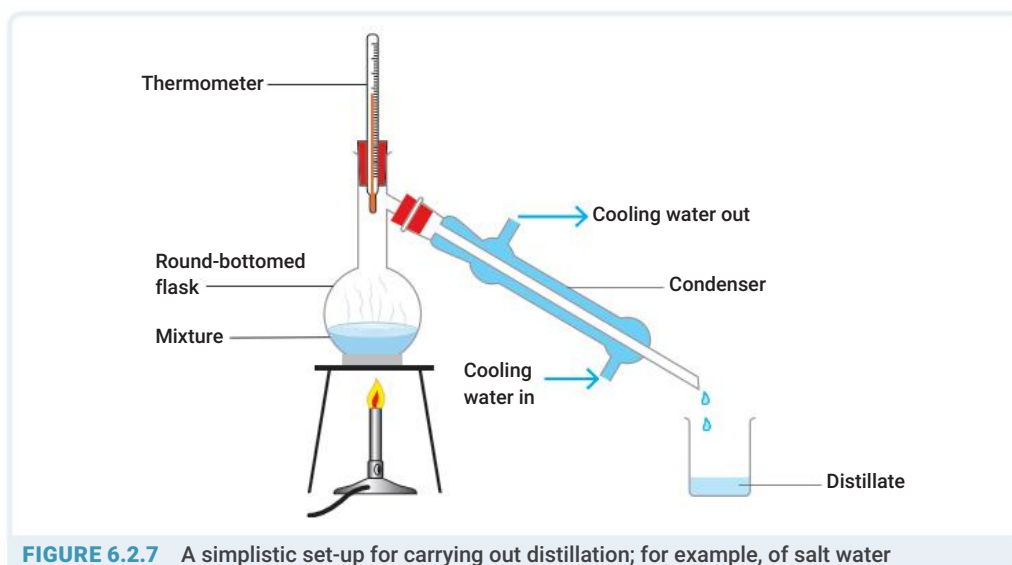
In homogeneous mixtures such as liquid solutions, the substances in the mixture can be separated based on differences in boiling point in a process known as **distillation**. In the example in Figure 6.2.7, salt water is boiled (in the round-bottomed flask) to give off pure water vapour, which is cooled (in the condenser) and collected as pure water liquid (the distillate). This is an example of simple distillation.

Simple distillation is most effective in separating liquid mixtures where the difference in boiling point of the substances is at least 70°C. The substance with the lower boiling point distils and the substance with the higher boiling point remains in the flask. In the case of salt water, the salt remains in the flask. For mixtures in which there are smaller differences in boiling points of the components, fractional distillation is a more effective technique.

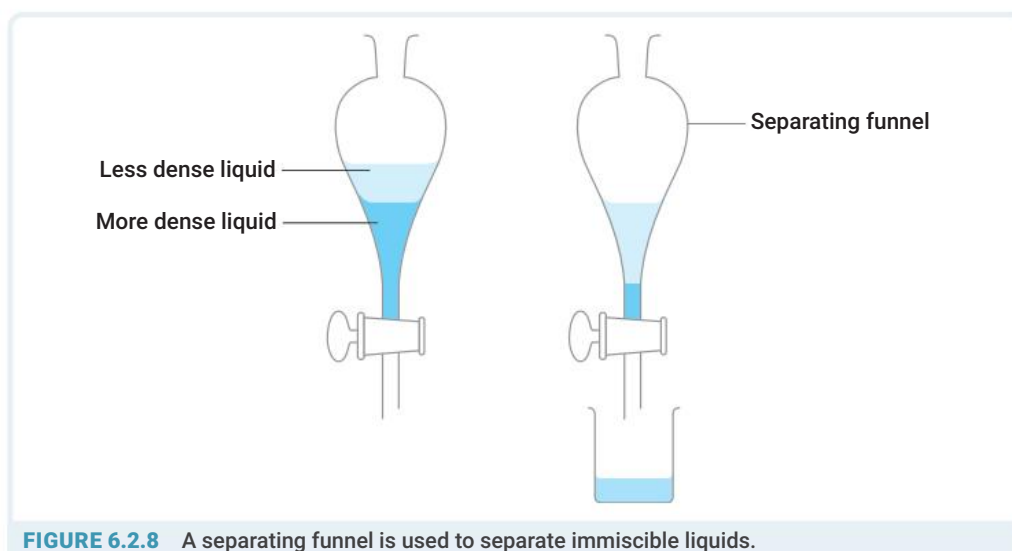
Distillation is a versatile technique and is used for various purposes. For example, fractional distillation is used to obtain petrol and many other useful fuels from crude oil. In areas where fresh water is scarce, distillation can be used to purify seawater to provide clean drinking water. Purifying seawater in this way requires a lot of energy and therefore, unfortunately, it is not a viable long-term solution for providing drinkable water.



Weblink
Fractional distillation



If a mixture is made up of immiscible liquids (liquids that are unable to dissolve in each other, e.g. water and oil), they can be separated by using a separating funnel (**Figure 6.2.8**). In this technique, the denser liquid leaves the funnel first, leaving the less dense liquid in the separating funnel.



Weblink
Separating funnels

Worksheets
Separating methods

Physical separation
techniques

Separating techniques
crossword

Table 6.2.2 summarises some of the different separation methods and the properties used for separation.

TABLE 6.2.2 Different separation methods and the properties used to separate the mixture

Separation method	Type of mixture	Property
Sieving	Solids and solids or solids and liquids	Differences in particle size
Filtration	Solids and liquids	Differences in particle size and state
Simple distillation	Liquids and liquids or solids and liquids	Large differences in boiling point
Fractional distillation	Liquids and liquids	Small differences in boiling point
Separating funnel	Liquids that are unable to be dissolved in each other	Differences in density

PRACTICAL ACTIVITY 6.2.1

EXTRACTING WATER FROM COPPER(II) SULFATE SOLUTION

Introduction

Copper(II) sulfate is sometimes used in fungicides and herbicides. When mixed with water, copper(II) sulfate forms a homogeneous mixture called a solution.

In this experiment, you will be recovering the water from a copper(II) sulfate solution by distillation.

Research question

How can physical properties such as boiling point be used to separate mixtures via distillation?

Aim

To separate water from a copper(II) sulfate solution

Equipment

• 20 mL of 0.5 M copper(II) sulfate	• 50 mL measuring cylinder	• rubber connection tubing	• heatproof mat
• 2 × 100 mL conical flasks	• 2 pieces of glass tubing with rubber stoppers	• tripod	• matches
• test tube		• gauze mat	• tongs
		• Bunsen burner	• retort stand
			• clamp and boss head



What are the risks in doing this experiment?

Exposure of copper(II) sulfate to skin

Exposure of copper(II) sulfate to aquatic life

How can you manage these risks to stay safe?

Wear PPE gear, including glasses, gloves and a lab coat. Wash hands thoroughly if contact is made with skin.

Dispose of chemicals in the waste bucket or container provided.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

- 1 Place a Bunsen burner on a heatproof mat.
- 2 Place a tripod on the heatproof mat so that it sits above the Bunsen burner. Place the gauze mat on top of the tripod.
- 3 Measure 20 mL of copper(II) sulfate solution and place it in a conical flask.
- 4 Insert the rubber stopper that is connected to the glass tubing into the conical flask containing copper(II) sulfate solution.
- 5 Attach one end of the rubber tubing to the glass tubing.
- 6 Place the conical flask containing the copper(II) sulfate solution onto the gauze mat on the tripod.
- 7 Secure the neck of the conical flask with the clamp to the retort stand. This will help hold the flask in place.
- 8 Insert the remaining rubber stopper with glass tubing into the empty conical flask. Connect the glass tubing to the other end of the rubber tubing. Leave the empty conical flask on the bench. The set-up should look similar to Figure 6.2.9.

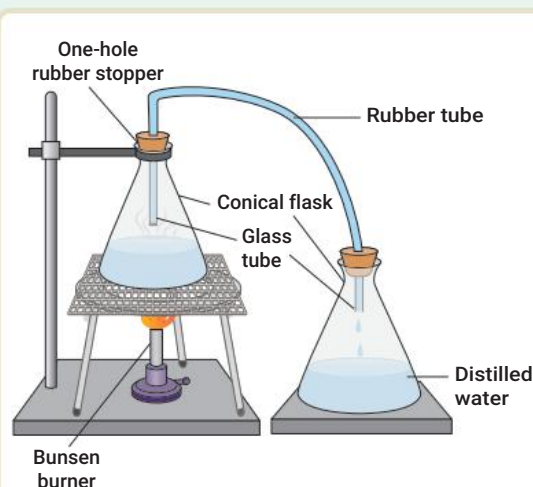


FIGURE 6.2.9 The experimental set-up for the distillation process

- 9 Use a match to light the Bunsen burner and gently heat the flask containing the copper(II) sulfate solution on a blue flame.
- 10 Continue to heat the flask until only a small amount of copper(II) sulfate solution is left. Do not let it dry out.
- 11 Observe the contents of the second conical flask.
- 12 Leave the set-up for a few minutes to allow the materials to cool down before packing up.

Results

Note down any observations made during the experiment.

Analysis of results

- 1 Describe what was happening throughout the experiment.
- 2 Identify the property being used to separate the two components and how this property allows separation.

Interpretation

- 3 Based on the results, which substance, water or copper(II) sulfate, has a higher boiling temperature.

LEARNING CHECK 6.2

DESCRIBING

- 1 **Identify** the key difference between a homogeneous mixture and a heterogeneous mixture.
- 2 **Identify** which of the following statements about a homogeneous mixture is not correct.
 - A Substances are visually distinguishable from each other.
 - B The mixture appears uniform throughout.
 - C At a molecular level, the components are evenly distributed.
 - D A solution is an example of a homogeneous mixture.

APPLYING

- 3 The smoke that is produced from a burning fire contains particulates of organic matter, methane gas, carbon monoxide gas as well as many other chemicals. Would smoke be considered as a homogeneous or heterogeneous mixture? **Explain** your answer.
- 4 A student made a cup of tea using loose tea leaves. **Describe** the process(es) they could use to separate the water, sugar and loose tea leaves from the mixture.
- 5 A student conducted an experiment to classify clear liquids as homogeneous mixtures or heterogeneous mixtures. They were given two beakers of liquid to evaluate as shown in **Figure 6.2.10**.
Based on what is shown in the figure **classify** each sample as a homogeneous or heterogeneous mixture.
- 6 A student was given a liquid mixture containing dissolved magnesium chloride (MgCl_2) and sand. **Identify** the physical properties that would be used to separate this mixture.



FIGURE 6.2.10 The experimental set-up to classify clear liquids

TURTLE ROCK SCIENTIFIC LLC / Science Photo Library

6.3 Properties of pure substances and mixtures

Water and salt water are both colourless, transparent liquids; however, one is a pure substance made up of only water (H_2O) molecules, and the other is a homogeneous mixture of salt dissolved in water. Although it is difficult to visually differentiate between the two, we can determine whether a sample is a pure substance or a mixture by examining its properties, such as boiling point.

At standard atmospheric pressure, the boiling point of pure water is always 100°C , no matter how much water is left in the vessel. However, the boiling point of salt water rises as the concentration of salt in the water increases (Figure 6.3.1). A solution of seawater contains approximately 3.5 g of salt in every 100 g of seawater and has an initial boiling point of 100.64°C . However, as the water is boiled off and the concentration of salt in the remaining seawater increases, the boiling point rises.

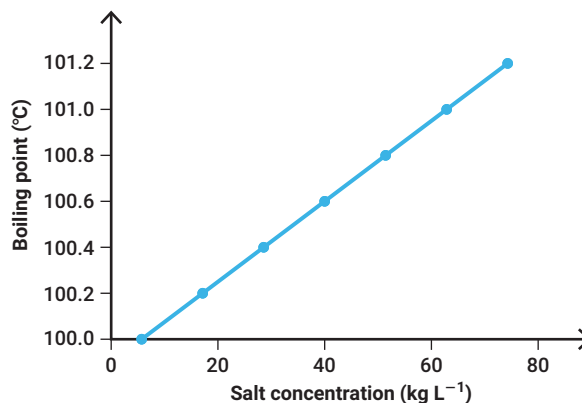


FIGURE 6.3.1 The increase in boiling point of salt water as the concentration of salt increases.

Other properties that are consistent for pure substances but vary for mixtures are melting point, density and reactivity, which for mixtures will vary depending on the composition and the quantities of each substance. For example, the density of seawater varies around the world, mainly because of the different amounts of salt present. The amount of salt present in the water of the Pacific Ocean is just less than 26 ppt (parts per trillion – 1 part per trillion parts of a mixture). However, the Dead Sea, which sits between Israel and Jordan, has a salt composition of more than 300 ppt. The greater amount of salt in the Dead Sea means its water is much denser than that of the Pacific Ocean. The high density of seawater is the reason we can float in the Dead Sea.

At times, combining different elements or compounds can improve the usability of an otherwise pure substance. For example, pure gold is soft and malleable, and can therefore be easily moulded into different shapes. However, gold jewellery can be easily dented and scratched. To increase its strength and hardness, gold jewellery is approximately 90% gold mixed with metals such as copper and silver. This percentage changes depending on the carat of gold used. For example, 18carat gold is approximately 75% pure gold and 25% other metals, whereas 9carat gold is approximately 38% pure gold and 62% other metals. The higher percentage of pure gold in 18carat gold makes it less prone to corrosion and tarnish. Gold used in jewellery is a homogeneous mixture.



Worksheet
Properties of pure
substances and mixtures

LEARNING CHECK 6.3

DESCRIBING

- 1 **Describe** why pure substances and mixtures have different properties.
- 2 **Explain** how you could use boiling point to determine whether a substance is pure or a mixture.

APPLYING

- 3 Two samples of ice were used in an experiment. One sample contained only water and the other sample was a mixture of sugar and water. **Describe** an experiment to determine which ice sample was pure water. (You do not need to include any methodology.)

ANALYSING

- 4 The following table shows the melting points of different compositions of gold.

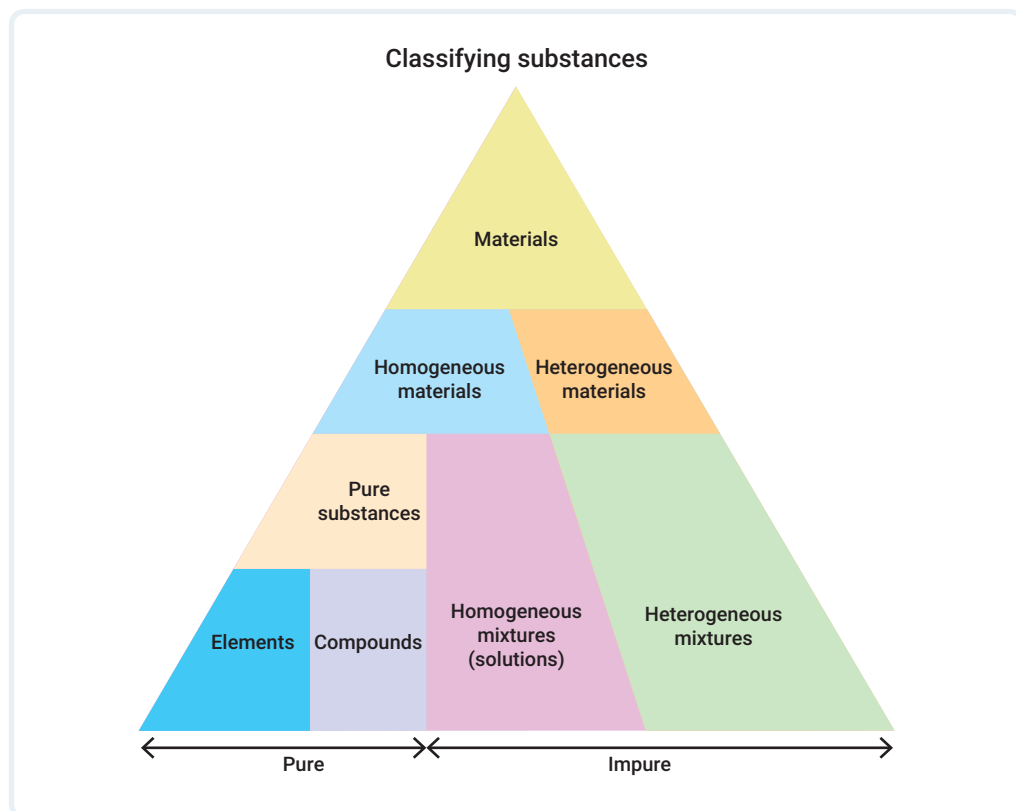
Gold	Percentage gold (%)	Approximate melting point (°C)
Pure gold	100	1063
18 carat gold	75	926
14 carat gold	59	880

- a **Identify** the trend in the melting point as the percentage of gold decreases.
- b A sample of gold was discovered on the ground and scientists wanted to determine if it was pure gold. Upon further investigation, it was found that the melting point was approximately 800°C. **Determine** whether this sample is pure or a mixture. **Explain** your answer.

CHAPTER SUMMARY

Classifying substances

- Materials can be classified into pure substances and mixtures. Mixtures can be further classified into homogeneous and heterogeneous mixtures.



Homogenous vs heterogenous mixtures

- Homogeneous and heterogeneous mixtures have different properties.

Feature	Homogeneous mixture	Heterogeneous mixture
Appearance	Substances are not visually distinguishable from each other and appear uniform throughout.	Substances are often visually distinguishable.
Composition	Substances are uniformly distributed.	Substances are unevenly distributed.
Examples	Cleaning solutions, steel, salt water	Blood, milk, muddy water

Separating mixtures

- Pure substances have distinct properties, whereas mixtures have the properties of the pure substances that they are made of. This also depends on the composition of the mixture.
- Although pure substances can only be separated by chemical means, mixtures can be separated based on differences physical properties.

Separation method	Type of mixture	Property
Sieving	Solids and solids or solids and liquids	Differences in particle size
Filtration	Solids and liquids	Differences in particle size and state
Simple distillation	Liquids and liquids or solids and liquids	Large differences in boiling point
Fractional distillation	Liquids	Small differences in boiling point
Separating funnel	Liquids that are unable to be dissolved in each other	Differences in density

CHAPTER EXAM

MULTIPLE CHOICE

1. What is a key difference between pure substances and mixtures?
 - A Mixtures are made of one element, whereas pure substances are made of more than one element.
 - B Pure substances can only be made from compounds, whereas mixtures can be both compounds or a single element.
 - C Mixtures can be separated into their components by physical processes, whereas pure substances cannot.
 - D Mixtures all exhibit the same properties whereas pure substances have varying properties.
2. Which of the following statements is true about homogeneous mixtures?
 - A Examples include solutions.
 - B They have distinguishable components.
 - C They have an uneven distribution of particles throughout.
 - D They cannot be separated by physical separation techniques.
3. The particles in heterogeneous mixtures:
 - A are all the same size.
 - B are evenly distributed.
 - C can often be visually distinguished.
 - D undergo a chemical reaction with each other.
4. Water (H_2O) is an example of:
 - A a mixture.
 - B an element.
 - C a solution.
 - D a pure substance.
5. A scientist removed a small sample of sodium metal from a larger block of sodium metal. Which of the following statements would be true?
 - A The reactivity of both samples of metal would be the same.
 - B The density of the two samples are different from each other.
 - C The boiling point of the larger sodium sample is greater than the smaller sample.
 - D The melting point of the smaller sample of sodium is slightly higher than that of the larger sample.

Questions 6 and 7 refer to the following information.

A scientist wanted to dilute a solution by adding more water. Instead, they accidentally added a different liquid solution into the original solution. There was no observable change when this occurred.

6. The new substance can be classified as:
 - A a reactant.
 - B a mixture.
 - C an element.
 - D a pure substance.
7. A property that can be used to separate the components of liquid solutions is their:
 - A boiling point.
 - B melting point.
 - C particle size.
 - D reactivity.

8. A piece of magnesium metal was allowed to react with an acid to form a solution. Compared to the piece of metal, the solution:
- A has the same density.
 - B has a different boiling point.
 - C has the same physical properties.
 - D does not contain magnesium because it has turned into a different element.

Questions 9 and 10 refer to the following information.

A light beam was shone through three different clear liquids. The results from the test are shown in the following table.

Liquid	Observation
X	Visible light beam
Y	Visible light beam
Z	No visible light beam

9. On the basis of these observations:
- A Z is a heterogeneous mixture.
 - B X is a heterogeneous mixture.
 - C X and Y must be the same substance.
 - D X and Y can both be considered to be homogeneous mixtures.
10. Substance Y:
- A is a pure substance.
 - B is made up of pure substances.
 - C has consistent chemical properties.
 - D has particles that are unevenly distributed.

SHORT RESPONSE

11. When melted, sodium chloride is considered a pure substance whereas when dissolved in water, the resultant solution is considered a mixture. **Explain** how this is possible.
12. Copper has a melting point of approximately 1084°C , whereas yellow brass, which contains copper, zinc, tin and lead, has a melting point of approximately 930°C . **Explain** why the two substances could have different boiling points even though both contain copper.

CROSS-CHAPTER QUESTION

13. Calcium chloride can be used to de-ice roads.
- a **Identify** whether calcium chloride is a metallic, an ionic or a covalent compound
 - b Draw a Lewis diagram to show the formation of the bond between calcium and chlorine.
 - c Calcium chloride can dissolve in water to form a homogeneous mixture. **Describe** what this means in terms of how the solute particles are arranged in the mixture.

DATA ANALYSIS

14. Analyse data

The following table shows the different densities of ethanol/water mixtures.

Mixture	Density (g mL^{-1})
52% ethanol/water	0.9110
38% ethanol/water	0.9408
24% ethanol/water	0.9549

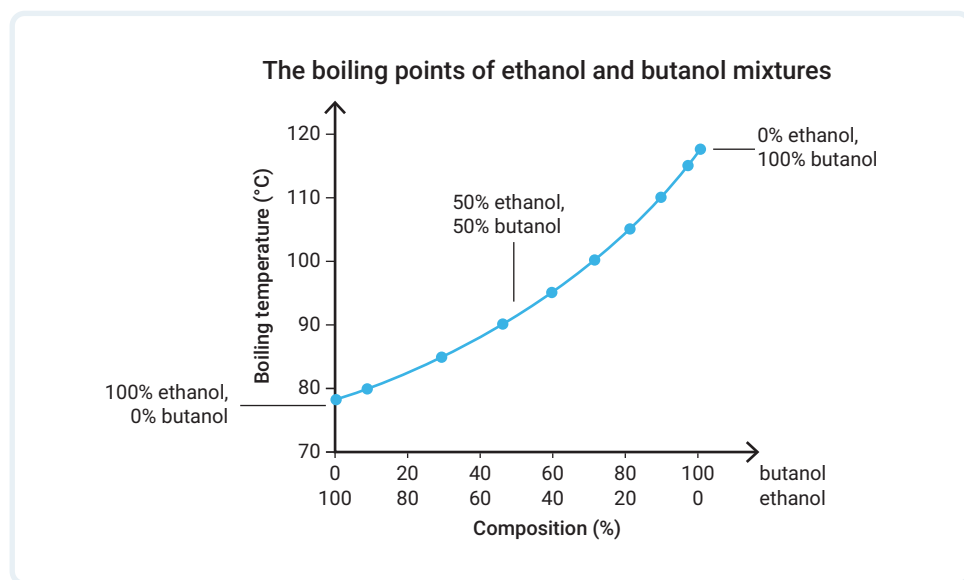
- Identify** the relationship between the amount of ethanol in the mixture and the density.
- On the basis of this information, what would you expect the density of pure ethanol to be relative to pure water?

15. Analyse data

A student conducted an experiment to measure the boiling points of the pure substances butanol and ethanol and a mixture of the two.

A graph of the results is shown below.

- Determine** the boiling points of ethanol and butanol.
- Comment on the boiling point of a mixture of 50% ethanol and 50% butanol. What does this suggest about the boiling point of mixtures?



Bonding and properties



Torychemistry/Shutterstock.com

SYLLABUS DOT POINTS

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 INQUIRY

Investigate:

- the properties of ionic, metallic and covalent compounds
- tests to distinguish alkanes and alkenes.*

*Note: Simulations may be used.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

All chemicals and substances can be classified and described according to their chemical and physical properties. These properties arise from how the atoms in a substance are bonded together. Chapter 5 outlined the different types of bonding that occur between different atoms. This chapter considers and explains the properties of substances containing each type of bonding.

Practicals

- Comparing different types of substances
- Comparing the properties of an ionic compound with those of its component elements (online-only resource)

Worksheets

- Different chemical structures
- Hydrocarbons
- Comparing properties of different substances

 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ The organisation of elements on the periodic table gives information about the element itself.
- ✓ Bonds occur as a result of electrostatic forces of attraction between atoms and other substances.
- ✓ Different bonds form in different ways.
- ✓ Metallic, ionic and covalent substances all differ in their bonding and properties.
- ✓ Chemical substances have a specific formula.
- ✓ The naming of ionic and covalent compounds follows different rules.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ describe common properties of metallic, ionic and covalent compounds
- ✓ distinguish between the types of bonding of substances by explaining the properties they exhibit
- ✓ explain the properties of ionic and covalent compounds by modelling their bonding arrangements
- ✓ identify how the strong electrostatic force of attraction between oppositely charged particles is related to the type of bonding
- ✓ discriminate between ionic and metallic bonding
- ✓ discriminate between covalent molecules, giant covalent networks and allotropes of carbon
- ✓ identify hydrocarbons based on their bonding between carbon and hydrogen
- ✓ name and draw simple unbranched hydrocarbons
- ✓ summarise the chemical properties of alkanes, alkenes and benzene
- ✓ relate the differences in chemical properties of alkanes, alkenes and benzene to their differences in bonding
- ✓ analyse data to determine the type of bonding in a compound
- ✓ distinguish between alkanes and alkenes by using chemical tests.

7.1 Ionic compounds

ionic compound

a compound consisting of atoms that have lost and gained electrons to form charged ions that are electrostatically attracted to one another

brittle a substance that will shatter, rather than bend, when a force is applied



Syllabus link

Chapter 5 discusses the model of ionic bonds.

The physical properties of **ionic compounds**, such as sodium chloride, are different from the properties of the elements from which they are composed.

Most ionic compounds:

- are hard and **brittle**, so they do not scratch easily, but shatter on impact
- do not conduct electricity in a solid state, but are good conductors in a molten state and in solution
- have high melting and boiling points.

From these physical properties, scientists can infer that there are strong bonds between the particles in ionic compounds and they are made of charged particles, as shown by their ability to conduct electricity in the liquid state and when dissolved in water.

Explaining properties of ionic compounds

In an ionic compound, each ion is strongly held in place by the attraction of adjacent, oppositely charged ions. This bonding is described as non-directional because ions may be attracted to each other in any direction. The ions become arranged in a regularly repeating pattern, known as a **lattice**. **Table 7.1.1** summarises the properties of ionic substances.

lattice a regular repeating three-dimensional arrangement of ions, which results in the formation of crystals

TABLE 7.1.1 Properties of ionic compounds

Property	Explanation
High melting point	There is strong attraction between ions, so a lot of energy is needed to move the ions out of their fixed position in the lattice.
Conducts electricity in the molten state and in solution	Charged particles are free to move in the molten state and when dissolved, but are held in a fixed position in the solid state.
Low thermal conductivity	Ions are fixed in their positions within the lattice and cannot move freely to transfer heat. Thermal conductivity can vary depending on crystal structure, bonding strength and temperature.
Hard but brittle	Ions are held rigidly in place in the lattice, but applying stress brings ions of the same charge closer together, so the crystal can shatter.
Mostly soluble in water	Ions are attracted to the water molecules and move out of position.
Giant lattice of repeating ions	Electrostatic attraction between positive and negative charges causes ions of opposite charge to surround each other.



Weblink
Properties of
ionic compounds

The non-directional bonding holds the ions in place, preventing ionic crystals from being malleable like metals. Ionic crystals are brittle, and shatter when hit with a strong force. When the crystal is distorted by force, like-charged ions move closer to each other and the strong electrostatic repulsions shatter the crystal. Ionic bonds are strong, as shown by the hardness and high melting and boiling points of ionic substances. The strength of the bond directly depends on the quantity of the charges and inversely depends on the distance between the charged particles. A cation with a 3+ charge forms a much stronger ionic bond than a cation with a 1+ charge. A larger ion forms a weaker ionic bond because of the greater distance between its electrons and the nucleus of the oppositely charged ion. Ionic compounds tend to have very low thermal conductivity owing to their lattice structure and the lack of freely moving electrons for heat transfer.

Figure 7.1.1 shows the information from Table 7.1.1 schematically, using sodium chloride as an example.

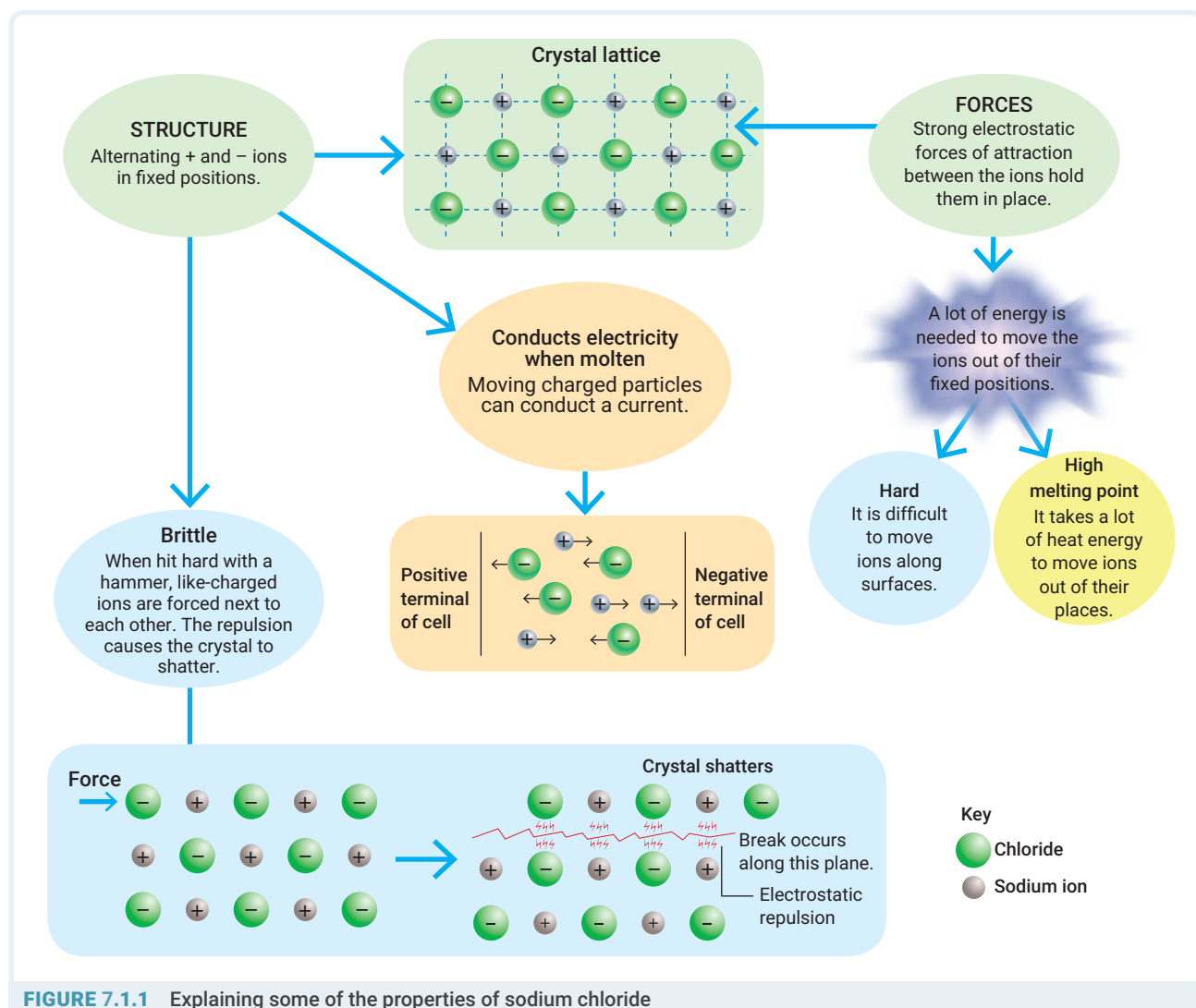


FIGURE 7.1.1 Explaining some of the properties of sodium chloride

LEARNING CHECK 7.1

DESCRIBING

- 1 **Define** 'ionic bonding'.
- 2 **Recall** three properties of ionic solids.

APPLYING

- 3 **Explain** the three properties from Question 2, in terms of ionic-bonding.
- 4 **Explain** why the bonding in an ionic solid is described as 'non-directional'.
- 5 **Explain** why ionic compounds conduct electricity when dissolved in water, but not in the solid state.

ANALYSING

- 6 Two ionic solids have the following melting points:

- sodium chloride 801°C
- lead(II) iodide 402°C.

Use your knowledge of ionic bonding and the relative positions of the elements involved on the periodic table to **explain** why the melting points of these compounds are significantly different.

7.2 Metals

Elements are grouped on the periodic table on the basis of common properties. Metals make up 75% of the elements, while the remainder are non-metals or metalloids. Each of the three groupings has distinct physical properties by which it can be classified. The common properties of metals are listed in [Table 7.2.1](#).

TABLE 7.2.1 Common properties of metals

Property	Description
Metallic lustre	Mirror-like shininess
Good conductor of heat	Allows heat to travel from one end to the other
Good conductor of electricity	Allows an electric current to easily pass through in the solid state and the molten state
Malleable	Can be beaten into another shape or flattened into a thin sheet with a hammer without breaking
Ductile	Can be drawn out into a wire



Syllabus link
Chapter 5 discusses the model of bonding within a metal.

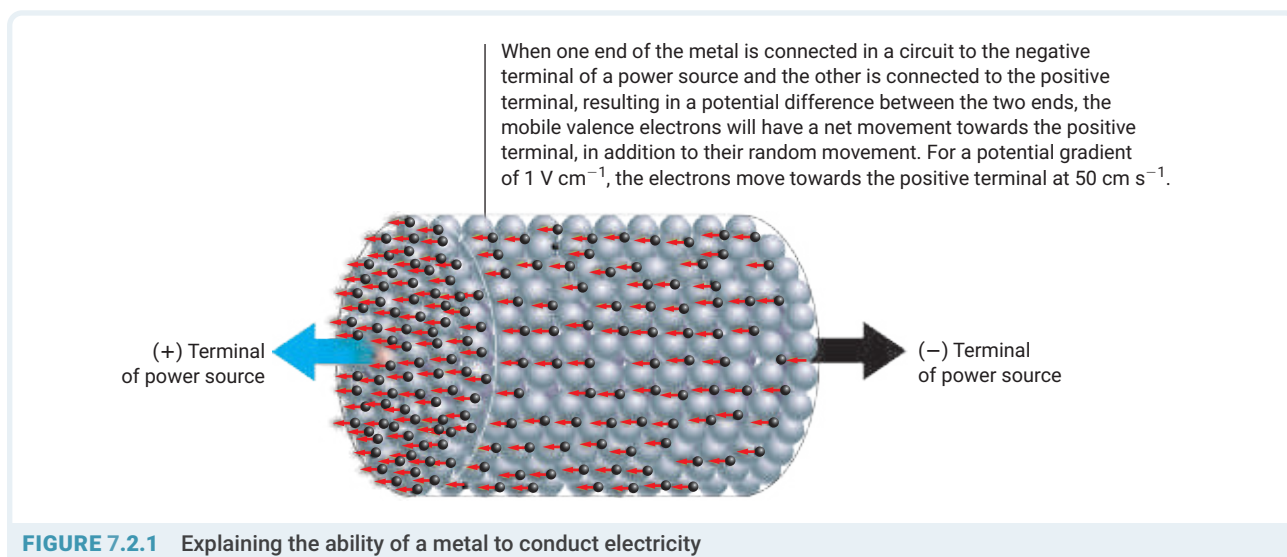


Weblink
[Properties of metals](#)

Conductivity

Metals are good conductors of electricity because of the **delocalised electrons** that can move within the lattice. When metals are connected into an electrical circuit, the delocalised electrons have a net movement towards the positive terminal ([Figure 7.2.1](#)). The movement of the electrons results in an electric current. Metals are also good conductors of heat because the delocalised electrons acquire energy from the heat source and rapidly transfer it to cooler parts of the lattice ([Figure 7.2.2](#)).

delocalised electrons
electrons that are not held in a fixed orbit around one nucleus, but are shared by many atoms



When one end of the metal is dipped into something hot, the moving sea of electrons in contact with this will absorb some of the heat energy, moving with far greater speed. These electrons collide with other electrons and their energy is transferred. In this way, the heat energy will be conveyed (very rapidly) to the other end of the metal.

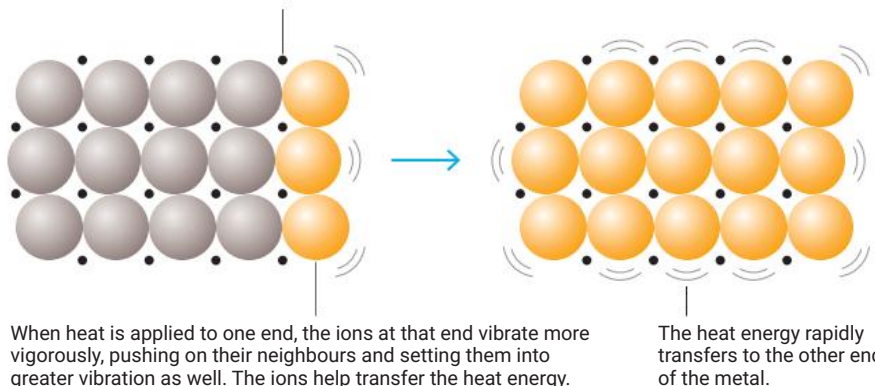


FIGURE 7.2.2 Explaining the ability of a metal to conduct heat

Lustre

A metal's lustre is explained by light rays being reflected off delocalised electrons. The close packing of the metal cations prevents light from passing through, which makes the metal opaque (**Figure 7.2.3**).

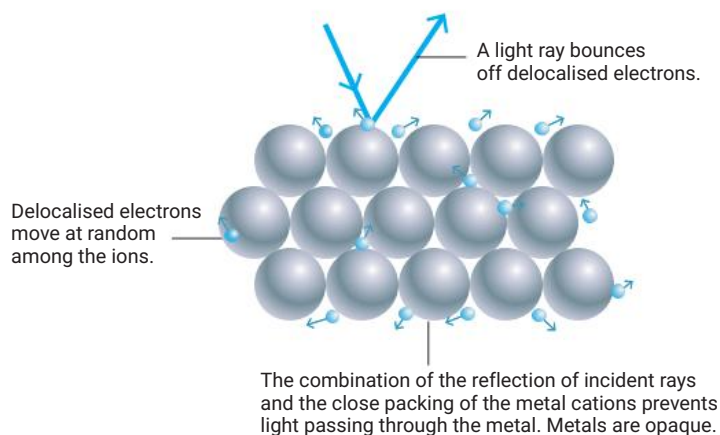


FIGURE 7.2.3 Explaining the lustre of a metal

Malleability and ductility

malleable can be repeatedly deformed, or hammered, without shattering

ductile can be drawn out to form a thin wire

Metals are **malleable** and **ductile** because, when the orderly array of cations is deformed, the delocalised electrons adjust to the new arrangement. This maintains the electrostatic attraction between the electrons and metal cations, thus allowing one layer of cations to slide over another without disrupting the metallic bonding (**Figure 7.2.4**).

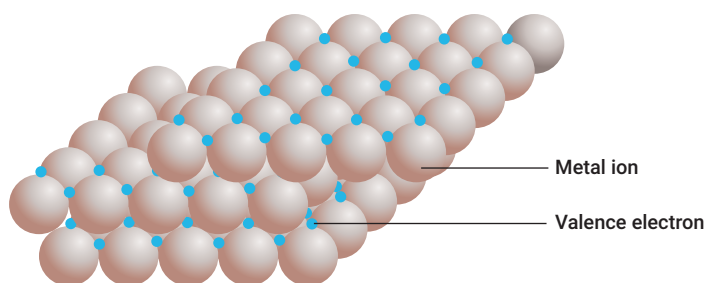


FIGURE 7.2.4 Explaining why metals are malleable and ductile

Properties that vary from metal to metal

Table 7.2.2 summarises the properties that vary from metal to metal.

TABLE 7.2.2 Some of the properties that vary from metal to metal

Property	Definition	Description
Density	The number of grams in each cubic centimetre of a material. If a material is denser than water, then it will usually sink in water.	Density varies from metal to metal. For example, 1 cm ³ of gold weighs 19.3 g, but the same volume of aluminium weighs only 2.70 g. Gold is described as a very dense metal, whereas aluminium is considered a 'light' (not very dense) metal. Density depends on how closely the ions are packed, the volume of each ion and the mass of each ion.
Hardness	Resistance to being scratched. The hardest substance cannot be scratched by any other material.	Most metals are difficult to scratch. When a substance is scratched, the particles in the surface layers are shifted out of position, disrupting the bonding forces. The more strongly the particles are held, the harder it is to scratch. Gold, lead and calcium are soft metals, i.e. they are easily scratched.
Melting point	The temperature at which the material melts, i.e. changes from a solid to a liquid.	Melting point varies considerably from metal to metal. Mercury (melting point -39°C) is already a liquid at room temperature, but many metals require temperatures much higher than 1000°C to melt. This means a lot of energy is required to overcome forces holding ions together. A strong metallic bond is the result of more delocalised electrons.
Boiling point	The temperature at which the material boils, i.e. changes from a liquid to a gas.	Boiling point varies considerably from metal to metal. This is due to the different strengths of metallic bonding in different metals. The metal with the highest boiling point is tungsten (5555°C) and the metal with the lowest boiling point is mercury (357°C).
Tensile strength	A material's ability to withstand a stretching force.	Tensile strength varies considerably from metal to metal, e.g. steel (which is mostly iron) has greater tensile strength than aluminium. The materials in cables that are suspended between poles must have high tensile strength.
Reactivity	Indicates the rate that a reaction occurs with a metal.	Reactivity varies considerably from metal to metal, e.g. group 1 alkali metals are more reactive than group 2 alkaline earth metals because they can easily lose electrons due to their low ionisation energies. Gold and platinum have low reactivity because they have stable electron configurations.

Recall from Chapter 5 that an alloy is a mixture of two or more elements, at least one of which is a metal. Alloys are engineered to have specific properties that improve on those of the constituent elements. They are typically designed to have enhanced strength, durability, corrosion resistance, electrical conductivity and other physical and chemical properties.

An example of an alloy is steel. Steel is composed of iron and carbon (and sometimes other elements such as chromium, nickel and manganese). Its properties include high tensile strength, durability and malleability, and it is used for construction of buildings and bridges, manufacturing tools and machinery, and in the automotive industry.

LEARNING CHECK 7.2

DESCRIBING

- 1 **Recall** three properties that are characteristic of metals.
- 2 **Define**, in reference to metals:
 - a hardness
 - b ductile
 - c malleable.
- 3 **Describe** three properties that vary between different metals.

APPLYING

- 4 **Explain** the following observations in terms of metallic bonding.
 - a You can draw out metals into a wire.
 - b You pick up a metal spoon that has been sitting in a cup of hot water and burn your hand.
 - c You should not hold a metallic umbrella during a lightning storm.
- 5 For each of the following, **identify** an important metallic property that makes it suitable for that purpose.
 - a Copper is used in electrical wiring.
 - b Aluminium is used in saucepans.
 - c Electric light filaments are made of tungsten.
 - d Iron is used in construction of bridges.
 - e Jewellery is commonly silver and gold.

ANALYSING

- 6 Use the information in the table to answer the questions that follow.

Element	Atomic mass (g mol ⁻¹)	Melting point (°C)	Thermal conductivity at 25°C (J s ⁻¹ m ⁻¹ K ⁻¹)	Electrical conductivity at 25°C (mS m ⁻¹)
Sodium	22.99	98	141	21
Magnesium	24.31	650	156	22
Aluminium	26.98	660	237	37
Copper	63.55	1085	401	58.4

- a **Compare** the thermal and electrical conductivities of the metals. Is there a trend in the data?
- b **Explain** any trend identified in part a in terms of delocalised electrons.
- c **Determine** if there is a relationship apparent between positions on the periodic table and melting point. Suggest why this may be so.

7.3 Covalent compounds

As discussed in section 5.4, non-metallic atoms are held together as elements and compounds by covalent bonding. In covalent compounds, atoms share electron pairs to form covalent bonds. The strong electrostatic force of attraction between the shared electron pair and the positively charged nuclei of the bonded atoms holds the atoms together. This results in the formation of strong covalent bonds, which contribute to the stability and structural integrity of the compound.

Covalent bonds are directional, meaning that the electron density between the nuclei is concentrated along the axis connecting the bonded atoms. This directional bonding leads to specific geometric arrangements of atoms, depending on the number and arrangement of electron pairs shared between atoms. Covalent compounds can exhibit different types of covalent bonding depending on the number of electron pairs shared between atoms. Single bonds involve the sharing of one electron pair (two electrons), double bonds involve the sharing of two electron pairs (four electrons) and triple bonds involve the sharing of three electron pairs (six electrons). The strength and properties of covalent compounds can vary depending on the type and number of covalent bonds present in the molecule. Covalent substances can be divided into two distinct groups on the basis of their properties: **covalent molecular substances** and **giant covalent network substances**. In each of these substances, although the bonding remains covalent, the arrangement of the atoms and ultimately the properties of the resultant substances differ, as shown in **Table 7.3.1**.

covalent molecular substance a substance composed of discrete molecules whose atoms are held together by intramolecular bonds

giant covalent network substance a compound made up of carbon or silicon whose structure is a three-dimensional network of covalently bonded atoms

TABLE 7.3.1 Contrasting properties of covalent substances

Property	Covalent molecular substances	Covalent network substances
Melting and boiling points	Low melting and boiling points (many are liquids and gases at room temperature)	Very high melting points (usually solids at room temperature)
Electrical conductivity	Non-conductors in solid and liquid states	Non-conductors in solid and liquid states
Thermal conductivity	Molecules are held together by weak forces; there is limited interaction between neighbouring molecules, leading to poor thermal energy transfer.	Atoms are bonded together by strong covalent bonds in a rigid and interconnected lattice, leading to higher thermal conductivity than molecular substances.
Solid	Generally soft	Hard and brittle
Solubility	Variable	Insoluble in water and most other solvents
Reactivity	Variable	Chemically unreactive
Examples	Oxygen, water and propane (C_3H_8)	Diamond and quartz (silicon dioxide)



Weblink
Simple covalent molecular substances

Covalent molecular substances

Covalent molecular substances, whether elements or compounds, are composed of discrete molecules. Covalent bonding holds the atoms in the molecule together and is referred to as

intramolecular bonding
bonds that occur
between atoms within a
molecule



Syllabus link
Bond energy is
discussed in further
detail in Chapter 10

intermolecular bonding
bonds that exist between
molecules; much weaker
than intramolecular
bonds but important
in determining the
physical properties of
the molecular substance
such as melting and
boiling points and
solubility



Syllabus link
Intermolecular
bonds are discussed
in greater detail in
Chapter 11.

intramolecular bonding because it operates within the molecules. The strength of the covalent bond is measured by the energy required to break it; that is, the energy necessary to separate the bonded atoms. Separating any pair of bonded atoms requires energy; the stronger the bond, the greater the energy required to break it. Bond energy, or bond dissociation energy, will be discussed in Chapter 10. The strength also depends on the number of electron pairs in the bond. Generally, as bond strength increases, the bonded atoms are drawn closer together so bond length decreases. Triple bonds are stronger and shorter than double bonds between the same two atoms, and double bonds are stronger and shorter than single bonds between the same two atoms.

The forces between separate molecules are quite weak compared to those within the molecule. These weak forces act to bond molecules together and are called **intermolecular bonding**. The nature of intermolecular bonds will be discussed in Chapter 11. **Table 7.3.2** shows the properties of covalent molecular substances in terms of their bonding.

TABLE 7.3.2 Explaining properties of covalent molecular substances

Property	Covalent molecular substances	Explanation
Melting and boiling points	Low (many are liquids and gases at room temperature)	The forces that hold the molecules together in the solid and liquid state are weak.
Conductivity	Non-conductors of electricity and heat in solid and liquid states	There are no charged particles that can move through the substance.
Solid	Generally soft	The forces that hold the molecules together in the solid state are weak.
Malleability	Tend to be malleable rather than shatter	The forces between molecules are weak, so molecules are easily moved relative to each other.
Solubility	Variable	Solubility depends on the intermolecular forces, which vary between substances.

Examples of covalent molecular substances are methane (CH_4), water and carbon dioxide (**Figure 7.3.1**). Although some variation occurs among the properties of different covalent molecular substances, the intramolecular forces within them are significantly stronger than the intermolecular forces between them.

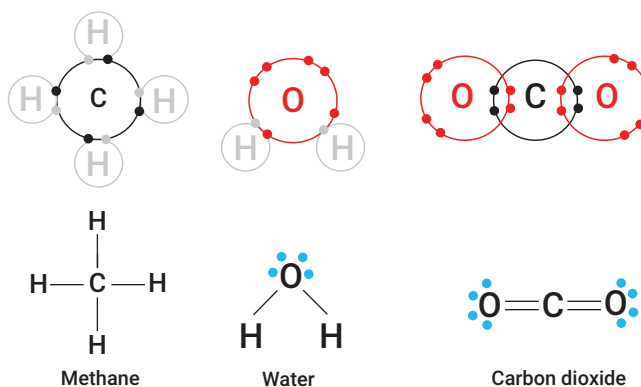


FIGURE 7.3.1 Examples of covalent molecular substances showing Lewis structure and valence structure

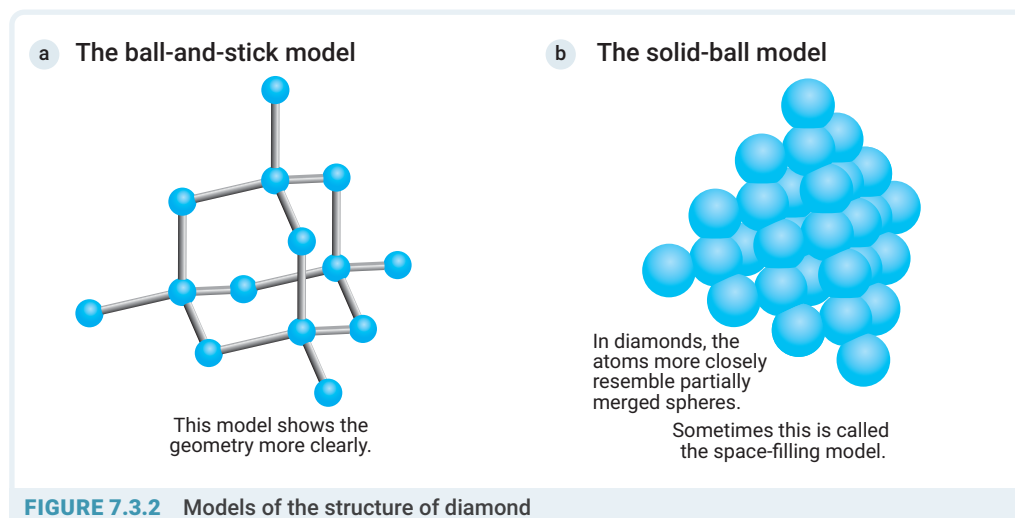
Giant covalent network substances

Carbon and silicon both have a valency of four. This means they can form four covalent bonds. As a result, they can form three-dimensional networks of covalently bonded atoms. In these networks, the covalent bonding extends indefinitely throughout the whole crystal. The three-dimensional nature of the network means the atoms are held rigidly in position by very strong bonds.

Elemental carbon exists as a range of different forms or **allotropes**, including graphite, diamond and fullerenes, with significantly different structures and physical properties.

Carbon in the form of diamond is an example of a giant covalent network. Each carbon atom is covalently bonded to four other carbon atoms in an orderly pattern that continues throughout the whole structure to make a giant crystalline lattice (**Figure 7.3.2**). The bonding electrons are tightly bound and highly localised.

allotrope one of two or more forms of the same element, with different physical properties, e.g. diamond and graphite are both different forms of carbon



Carbon's properties in diamond form are consistent with those listed in Table 7.3.1. Silicon has the same valency as carbon and exists in a structure similar to diamond.

There are also compounds that exist as giant covalent network substances. Silicon and carbon covalently bond to form the crystalline compound silicon carbide (SiC), commonly known as carborundum. This compound has a rigid diamond-like structure with alternating silicon and carbon atoms.

Silicon dioxide (SiO₂), also known as quartz, is another, more common giant covalent network compound. Each silicon atom is bonded to four oxygen atoms, and each oxygen atom is bonded to two silicon atoms (**Figure 7.3.3**). This pattern extends throughout the entire crystalline lattice. Quartz is the main component of sand. The properties of silicon dioxide and silicon carbide are also consistent with covalent network substances in Table 7.3.1.

In a giant covalent network substance, the forces are strong and extend through the whole network, which explains why their properties differ from those of covalent molecular substances. **Table 7.3.3** shows the properties of covalent network substances.



Worksheet
Different chemical
structures

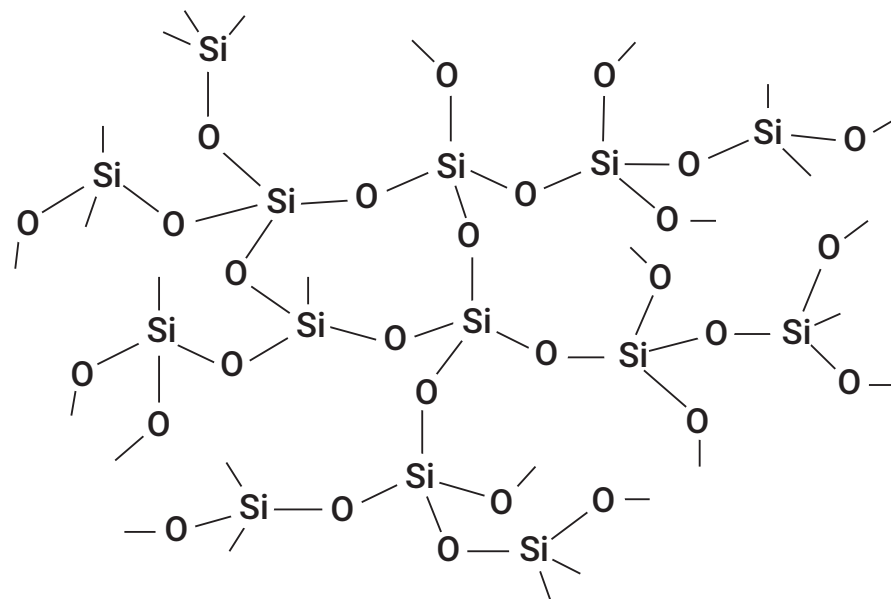


FIGURE 7.3.3 The covalent network structure of silicon dioxide

TABLE 7.3.3 Explaining properties of covalent network substances

Property	Covalent network substances	Explanation
Melting point	Very high (usually solids at room temperature)	The covalent bonds that hold the atoms together are strong and extend throughout the network.
Conductivity	Non-conductors of electricity and heat in solid and liquid states	There are no charged particles that are free to move throughout the structure.
Hardness	Extremely hard and brittle solids	The bonds are very strong between atoms so it is difficult to scratch, but an impact force disrupts the positions of the atoms and causes the network to shatter.
Solubility	Insoluble in water and most other solvents	There is no attraction between the atoms in the network and water molecules.

Other allotropes of carbon

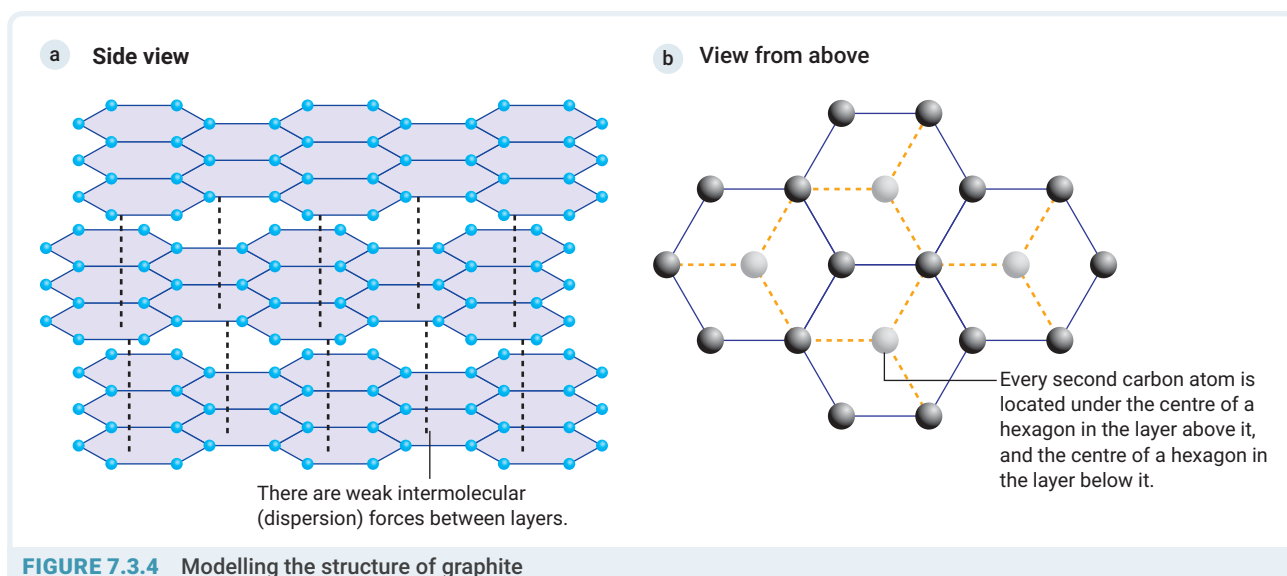
Graphite



Weblink
Graphite

Although graphite is a covalent compound and an allotrope of carbon, it has different properties from those of diamond. Like diamond, graphite has a high melting and boiling point, but graphite is soft and a good conductor of electricity.

In graphite, the carbon atoms are arranged in flat parallel layers. Within each layer, each carbon atom is covalently bonded to three other carbon atoms, forming a hexagonal arrangement. This means that each layer is a two-dimensional network of carbon atoms. Weak bonding holds the layers together (**Figure 7.3.4**).



In graphite, only three of carbon's four electrons are involved in covalent bonding. The fourth electron is delocalised, which means it is free to move along the layer occupied by its 'parent' atom. The attraction between the delocalised electrons and the layers above and below it hold the layers together.

Graphite's high melting and boiling points can be explained by the strong covalent bonds within the layers, while its softness and slippery feel can be explained by the weak bonds between layers that allow them to easily slide over each other. Graphite is an electrical conductor because of the presence of the delocalised electrons that can move and conduct electricity along the layers.

Fullerenes

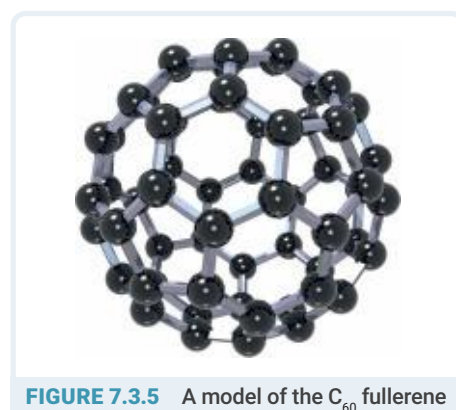
Although fullerenes are not covalent network substances, they are an allotrope of carbon and are held together by covalent bonds, so it is appropriate to include them here.

The first fullerene molecule was discovered in 1985 by a group of researchers at Rice University, Texas, United States. It has the formula C_{60} and was named after Buckminster Fuller because it resembles the geodesic domes he designed. The molecules are more commonly referred to as 'buckyballs'.

Fullerenes occur in nature and in space. The discovery of fullerenes expanded the number of known carbon allotropes. They have become the subject of intense research, especially in materials science, electronics and nanotechnology.

Fullerenes are molecules composed entirely of carbon. They are similar in structure and bonding to graphite. Each carbon is covalently bonded to three other carbons and the fourth electron is delocalised and free to move throughout the structure. Fullerenes differ from graphite because they contain pentagonal and hexagonal rings that prevent them from being planar. The structure resembles a soccer ball and is often referred to as an open cage structure (**Figure 7.3.5**).

Like graphite, fullerenes are soft and slippery because of the weak forces between molecules. They do not conduct electricity because there is no movement of electrons from one molecule to the next. Like diamond and graphite, they are insoluble in water because there is no attraction between the carbon atoms and water molecules. They typically have low melting points due to weak bonding between molecules.



LEARNING CHECK 7.3

DESCRIBING

- 1 **Recall** three differences between the properties of covalent molecular substances and covalent network solids.
- 2 **Identify** the elements most likely to form giant covalent network solids.

APPLYING

- 3 **Explain** why covalent molecular substances and giant covalent network substances have different properties even though they both have covalent bonding.
- 4 **Explain** why graphite can conduct electricity whereas diamond cannot.
- 5 **Explain** why graphite and diamond both have high melting points, yet graphite is soft and diamond is extremely hard.
- 6 **Compare** the properties of graphite and fullerenes.

ANALYSING

- 7 Substances A–D all have covalent bonds between atoms.

Substance	Melting point (°C)	Boiling point (°C)
A	–125	–90
B	5	170
C	2200	3900
D	55	325

- a **Identify** which substance has a giant covalent network structure.
- b **Identify** the type of structure the remaining substances have.
- c **Identify** which substance is a gas at room temperature.
- d **Identify** which substance is a liquid at room temperature.

7.4 Hydrocarbons

organic chemistry the study of carbon-based compounds

The vast majority of compounds known to chemists are carbon compounds, because carbon is found in more places, in more forms, than any other element. A branch of chemistry, **organic chemistry**, is devoted to carbon-based compounds. A major reason for this is that carbon readily bonds to other carbon atoms (as seen in its allotropes) and each atom can form a total of four bonds from single to double to triple.

Chemists divide organic compounds into ‘families’ based on their structural features. This makes managing the huge number of carbon compounds easier. Families are organised according to which atoms or groups of atoms are bonded to carbon, as well as whether multiple bonds are present.

Hydrocarbons are compounds made up of only hydrogen and carbon. The simplest group of hydrocarbons is the family of alkanes.

Alkanes

The alkane family is very important because it includes the fuels that supply us with energy every day. Most Australians use alkanes to run our cars, supply us with hot water, heat our homes and cook our foods. Alkanes also supply us with the raw materials to make many other substances, including plastics.

hydrocarbon a compound containing only hydrogen and carbon

The simplest alkane is methane (CH_4). Ethane (C_2H_6), propane (C_3H_8) and butane (C_4H_{10}) are next in the series. **Figure 7.4.1** shows models and structural formulas for propane and butane.



Weblink
Alkanes

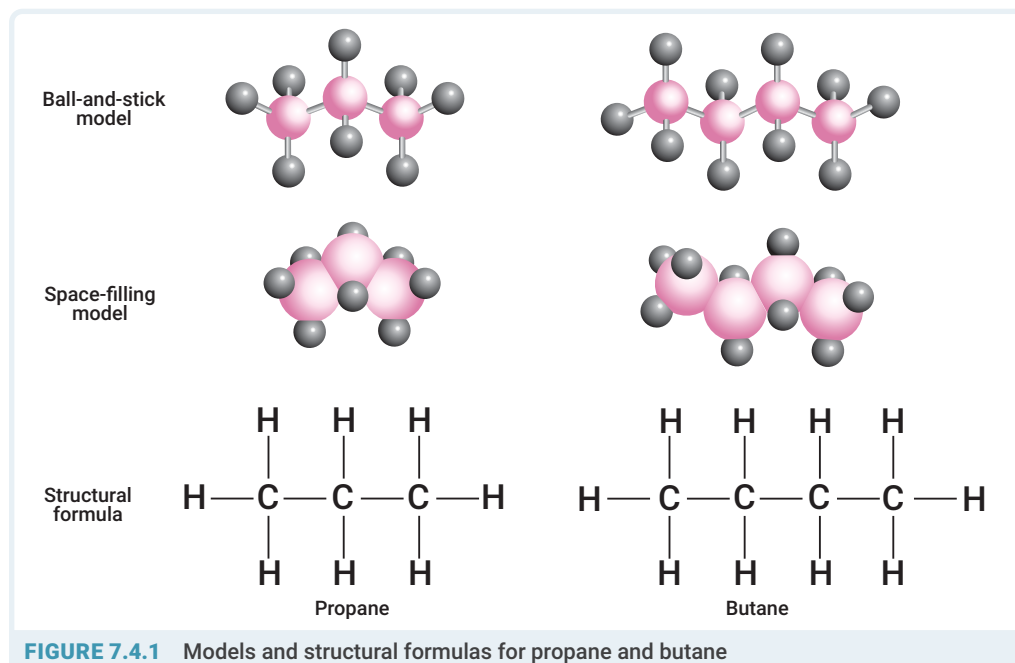


FIGURE 7.4.1 Models and structural formulas for propane and butane

TABLE 7.4.1 The first 10 members of the alkane family

Name	Formula
Methane	CH_4
Ethane	C_2H_6
Propane	C_3H_8
Butane	C_4H_{10}
Pentane	C_5H_{12}
Hexane	C_6H_{14}
Heptane	C_7H_{16}
Octane	C_8H_{18}
Nonane	C_9H_{20}
Decane	$\text{C}_{10}\text{H}_{22}$

Table 7.4.1 lists the first 10 alkanes. They are called ‘straight-chain alkanes’ because the carbon atoms are joined in one continuous string. The chain is not straight but more like a zigzag because of the way carbon atoms are arranged.

Alkanes are also classified as **saturated compounds**, which means they contain only single bonds between carbon atoms.

Each successive member of the alkane series is formed by replacing one H with a C and then adding three more hydrogens. This pattern follows for all straight-chain alkanes, so the general formula for alkanes can be expressed as $\text{C}_n\text{H}_{2n+2}$, where n is a whole number.

saturated compound
an organic compound with only single bonds between carbon atoms

Naming alkanes

As with ionic and covalent compounds, there is a systematic naming system for alkanes that allows us to work out the formula and structure of a compound. Although the naming of alkanes (and alkenes in the next section) is not required in this syllabus, it is an important skill for Unit 4.

The name of a carbon compound consists of a stem, which gives the length of the carbon chain, and an ending, which denotes the family of hydrocarbons the compound belongs to. The names of all compounds in the alkane family end in ‘-ane’. The stems of the first 10 members of the series vary according to the number of carbon atoms in the chain, as outlined in **Table 7.4.2**.

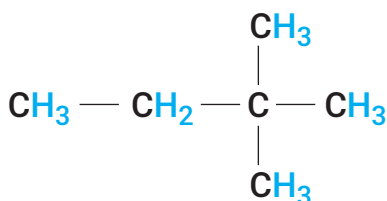


Syllabus link
Naming alkanes is an important skill for Unit 4.

TABLE 7.4.2 The stems of organic compounds, according to the number of carbon atoms in the chain

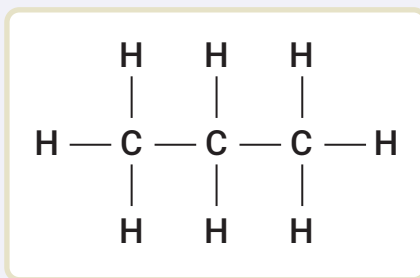
Number of C atoms	1	2	3	4	5	6	7	8	9	10
Stem	Meth-	Eth-	Prop-	But-	Pent-	Hex-	Hept-	Oct-	Non-	Dec-

The alkanes considered so far have been straight-chain alkanes; however, many carbon compounds contain branched chains. In branched-chain compounds, at least one carbon atom is attached to more than two other carbon atoms (**Figure 7.4.2**). These will be discussed further in Unit 4 Chemistry.

**FIGURE 7.4.2** An example of a branched chain alkane. The third carbon from the left has two other carbon atoms attached to it, causing two branches from the parent alkane.

WORKED EXAMPLE 7.4.1

Name the following alkane.



ANSWER

- 1 Identify the number of carbons in the carbon chain of the molecule.**
Three
- 2 Identify the prefix to signal the number of carbons in the carbon chain.**
Prop-
- 3 Write the name of the alkane.**
Propane
Since it is an alkane, the ending of the name of the compound is 'ane'.

Properties of alkanes

Alkanes are covalent molecular substances, so their properties are similar to those previously identified for covalent molecular substances. **Table 7.4.3** summarises some physical properties of alkanes. These properties can be explained by the covalent molecular nature of the molecules.

TABLE 7.4.3 Some physical properties of alkanes

Property	Alkanes
Appearance	Colourless and odourless
Melting point and boiling point	Low melting point and boiling point, increasing with molecular mass; the first four alkanes are gases at room temperature
Electrical conductivity	Non-conductors
Hardness	Soft in solid form
Malleability	Malleable in solid form
Solubility	Insoluble in water but soluble in some other solvents

Alkanes are known for their relatively inert (non-reactive) nature due to the strong carbon-carbon and carbon-hydrogen bonds. Some key chemical properties of alkanes are:

1. **Combustion:** Alkanes are highly flammable and undergo combustion reactions in a plentiful supply of oxygen to produce carbon dioxide and water. The combustion of alkanes releases heat energy, making them valuable as fuels. For example, the combustion of methane: $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$.
2. **Halogenation:** In the presence of halogen gases such as chlorine and ultraviolet (UV) light or heat, one or more hydrogen atoms are replaced by halogen atoms, resulting in the formation of alkyl halides. This reaction is a type of **substitution reaction**. For example, methane reacts with chlorine gas to produce chloromethane and hydrogen chloride: $\text{CH}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{CH}_3\text{Cl}(\text{g}) + \text{HCl}(\text{g})$.

substitution reaction

a reaction in which an atom or group of atoms takes the place of hydrogen on a parent carbon atom in a hydrocarbon molecule

Alkenes

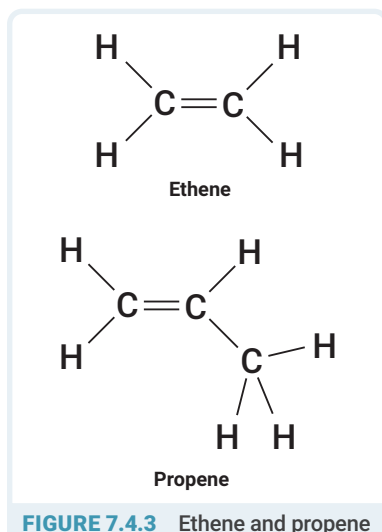
The alkene family contains hydrocarbons in which one pair of carbon atoms is joined by a double bond and all the other carbon atoms are joined by single bonds. The names of all compounds in the alkene family end in ‘-ene’. The prefix of the name is the same convention as that used for alkanes. The simplest alkene is ethene (C_2H_4) and the next member is propene (C_3H_6). The structures of these two alkenes are shown in **Figure 7.4.3**.

Compounds that contain multiple bonds between carbon atoms, such as alkenes, are called **unsaturated compounds**.

Table 7.4.4 shows the names and formulas of the first nine alkenes.

unsaturated compound

an organic compound that has at least one double bond or triple bond between carbon atoms

**FIGURE 7.4.3** Ethene and propene**TABLE 7.4.4** The first nine members of the alkene family

Name	Formula
Ethene	C_2H_4
Propene	C_3H_6
Butene	C_4H_8
Pentene	C_5H_{10}
Hexene	C_6H_{12}
Heptene	C_7H_{14}
Octene	C_8H_{16}
Nonene	C_9H_{18}
Decene	$\text{C}_{10}\text{H}_{20}$



Weblink
Alkenes



Syllabus link
Naming alkenes is
an important skill for
Unit 4.

Alkenes have the general formula C_nH_{2n} . As with alkanes, each successive member of the alkene series is formed by adding CH_2 to the formula of the previous member.

Naming alkenes

Naming alkenes is not as straightforward as naming alkanes. This is because, in alkenes with more than three carbon atoms, the double bond can occupy a number of different positions. Although this does not change the molecular formula of the compound, it does change the structural formula, which shows the position of the double bond. The example in **Figure 7.4.4** demonstrates this.

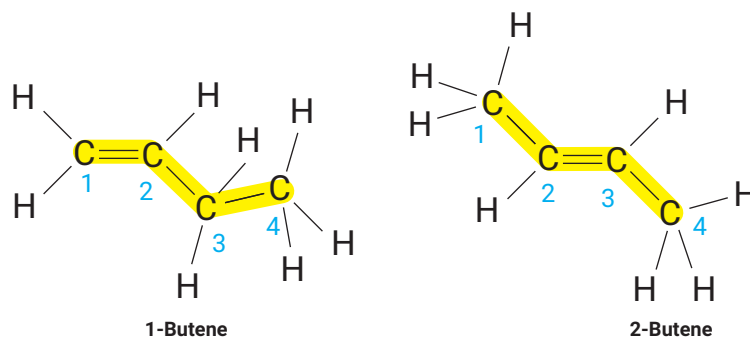


FIGURE 7.4.4 The different structures of butene

When naming alkenes, you need to indicate the location of the double bond. Alkenes are named according to the following rules.

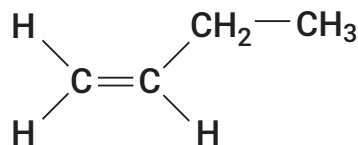
1. Take the usual stem name (such as 'eth-' and 'but-'), which indicates the number of carbon atoms in the chain, and add the ending '-ene'.
2. Determine the location of the double bond by numbering the carbon atoms from the end of the chain closest to the double bond (Figure 7.4.4).
3. Show the double bond's location by inserting, as a prefix, the number of the carbon at which the double bond starts, followed by a hyphen (-).

Branched alkenes will also be covered in Unit 4.

As seen in Figure 7.4.4, the number at the front of the name of the alkene denotes the location of the double bond. For 1-butene, this signals that the carbon-to-carbon double bond occurs between carbons 1 and 2 while the double bond in 2-butene occurs between carbons 2 and 3.

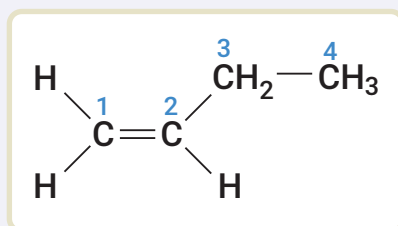
WORKED EXAMPLE 7.4.2

Name the following alkene.



ANSWER

- 1 Identify the number of carbons in the carbon chain of the molecule.
Four
- 2 Identify the prefix to signal the number of carbons in the carbon chain.
But-
- 3 Determine the location of the double bond.
It occurs between the first two carbons is on the left of the diagram.
- 4 Number the carbons from the end of the chain that is closest to the double bond.



- 5 Write the name of the alkene.

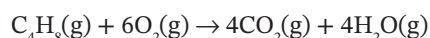
1-Butene

Remember to use the number at the front of the name followed by a hyphen to indicate the location of the double bond. IUPAC also accepts but-1-ene, i.e. putting the number near the part of the name that indicates the double (or triple) bond.

Properties of alkenes

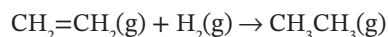
As might be expected, the physical properties of alkenes are similar to those of alkanes. As a family, alkenes show similar trends in melting point, boiling point and state under normal conditions. At room temperature, the first three alkenes (C_2 – C_4) are gases, C_5 – C_{15} are liquids, and the remainder are solids. Similar to alkanes, all alkenes are insoluble in water but dissolve in some other solvents.

The presence of the double bond introduces unique chemical properties to alkenes, allowing them to undergo various reactions. Like alkanes, alkenes undergo combustion in a plentiful supply of oxygen to produce carbon dioxide and water; for example, the combustion of butene:

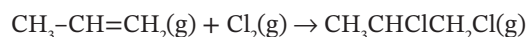


Alkenes also readily undergo **addition reactions**, where atoms or groups of atoms are added across the carbon–carbon double bond, leading to the formation of single bonds. Some key addition reactions:

1. Addition of hydrogen (H_2) in the presence of a catalyst such as Pt, Pd or Ni forms alkanes. For example, ethene is converted to ethane:



2. Addition of halogens across the double bond forms dihaloalkanes. For example, 1-propene reacts with chlorine to produce 1,2-dichloropropane:



3. Addition of water across the double bond in the presence of an acid catalyst forms alcohols. For example, ethene is converted to ethanol:



Alkenes can undergo polymerisation reactions, where multiple alkene (monomer) units are joined to form a long polymer chain. This occurs through the repeated addition of monomer



Worksheet
Hydrocarbons

addition reaction a reaction in which atoms are added to a molecule by breaking a double or triple bond

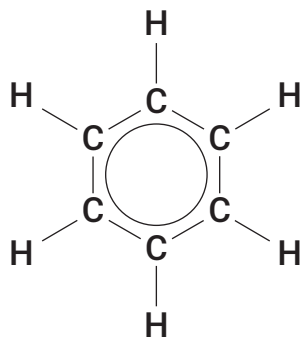


FIGURE 7.4.5 The structure of a benzene ring

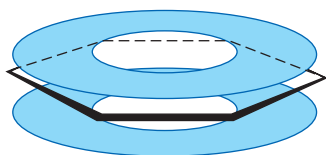


FIGURE 7.4.6 Bonding in benzene, shown schematically

aromatic hydrocarbon a hydrocarbon containing one or more benzene rings

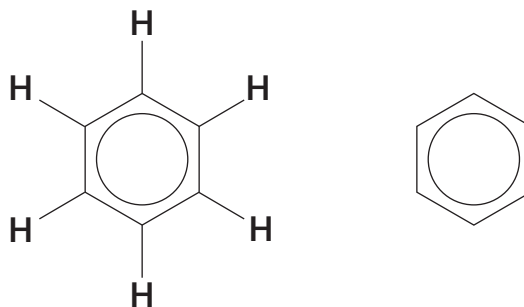


FIGURE 7.4.7 Different ways of drawing benzene

units across the double bond, resulting in polymers. This is an industrially significant process to produce various synthetic polymers, such as polyethylene, polypropylene and polyvinyl chloride (PVC).

Benzene

Benzene is the simplest member of the family of **aromatic hydrocarbons**. This group got their name because many of them have a strong aroma (smell).

Benzene is a colourless liquid with the molecular formula C_6H_6 , so it is an unsaturated molecule. The structure is a six-sided flat ring with all bond angles of 120° . The carbon atoms are joined in a regular hexagon with one hydrogen atom attached to each carbon (**Figure 7.4.5**).

On the basis of this representation, it was thought that benzene contained alternating single and double bonds, so they would undergo reactions similar to the alkenes. However, this rarely happens. The reactions of benzene are also completely different from those of other double-bonded hydrocarbons.

Further investigation of the benzene ring found that all the carbon-carbon bonds had the same bond length, somewhere between a single bond and a double bond. This can be explained if the electrons of the double bond are spread out around the ring of carbon atoms, making a cloud of delocalised electrons above and below the ring, shown schematically in **Figure 7.4.6**. This is similar to the delocalisation of electrons found in the structure of metals and graphite.

In chemical structure representations, the benzene ring is usually represented in either of the two forms shown in **Figure 7.4.7**.

Properties of benzene

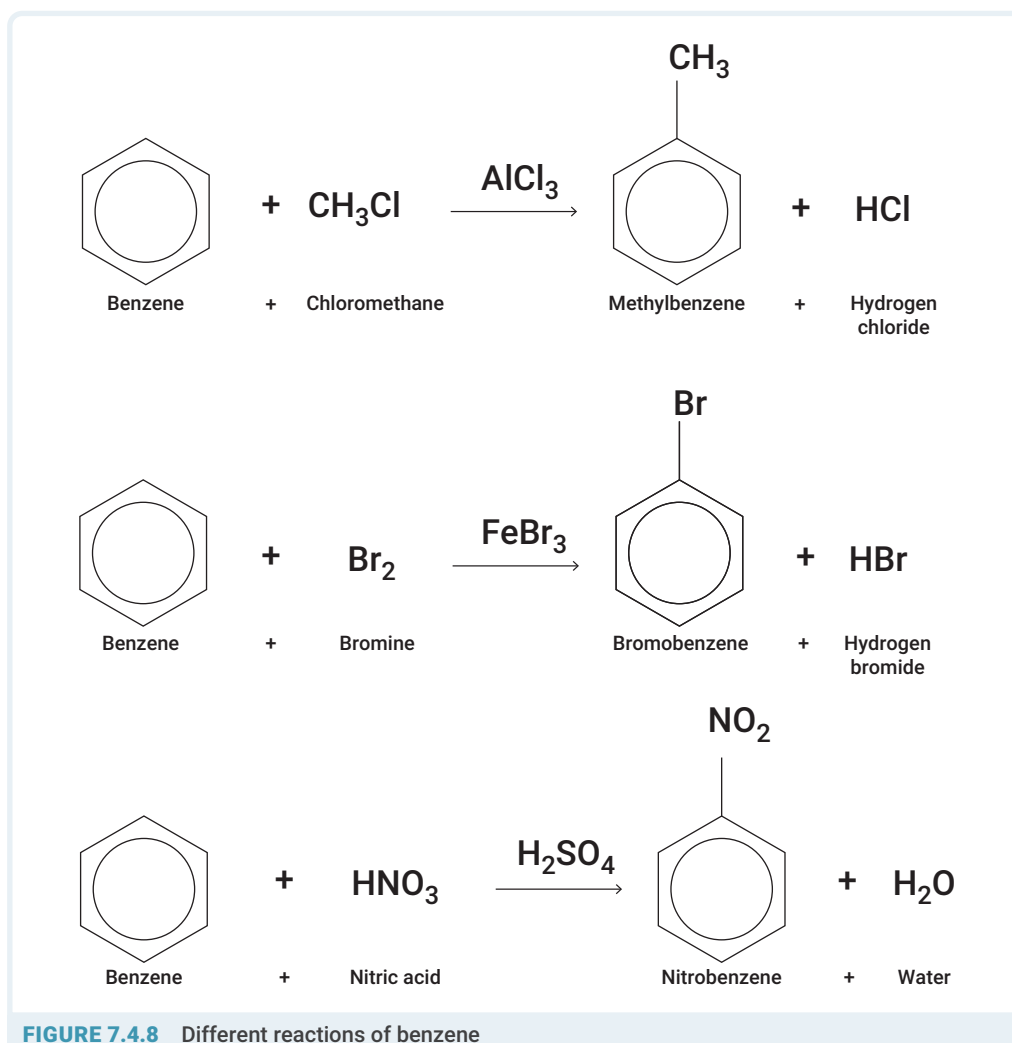
As with alkanes and alkenes, benzene has a low melting point, boiling point and density. It is also insoluble in water.

The cyclic delocalisation of electrons in the benzene molecule makes it extremely stable. Benzene tends to undergo reactions in which the stable ring is preserved because reactions that disrupt the bonding due to the delocalised electrons are less favourable and will only occur at higher temperatures and under more vigorous conditions.

Most reactions of benzene are substitution reactions (like alkanes) where one or more of the hydrogens are replaced by a different atom or group of atoms.

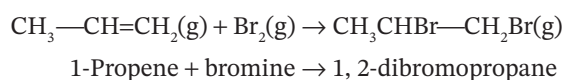
Some examples of substitution reactions are shown in **Figure 7.4.8**:

- alkylation: substitution with an alkyl group
- halogenation: substitution with a halogen
- nitration: substitution with a nitro group.



Tests to distinguish alkanes and alkenes

Because alkanes and alkenes are usually colourless gases or liquids, it is difficult to visually distinguish between the two families of compounds. The most common test is to react the alkane and alkene compounds with bromine water. Bromine water is a brown coloured mixture containing bromine gas (Br_2) dissolved in water. In the absence of UV light, bromine water reacts rapidly with alkenes in an addition reaction. The bromine atoms add across the double bond. This is a test for unsaturation. For example, 1-propene reacts with bromine to produce 1,2-dibromopropane:



The reaction is complete when the brown colour of the bromine water disappears and the solution becomes colourless. Bromine water does not react with saturated alkanes. The mixture stays the brown colour because there are no double bonds for Br_2 to add across.

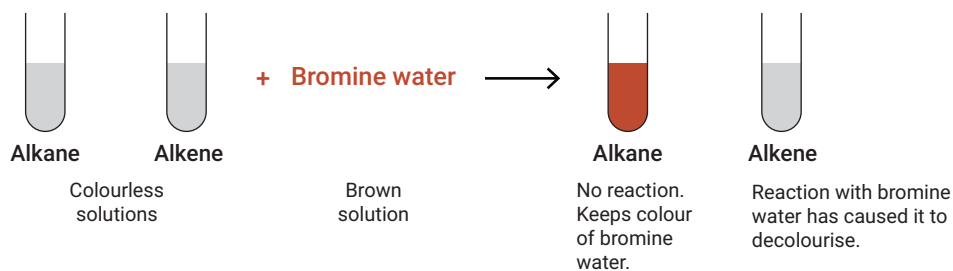


FIGURE 7.4.9 The bromine water test for unsaturation



Weblink
Reactions involving alkenes
and bromine

LEARNING CHECK 7.4

DESCRIBING

- State** the general formula for:
 - alkanes
 - alkenes
 - benzene.
- Distinguish** between saturated and unsaturated compounds.
- Compare** the structure and properties of alkanes, alkenes and benzene.
- Explain** why benzene is extremely stable.

APPLYING

- Identify** each of the following as an alkane or an alkene. **Justify** your choice in each instance.
 - C_6H_{12}
 - C_6H_{14}
 - C_9H_{18}
- Draw** a structural formula for:
 - ethane
 - butane
 - hexane
 - octane
 - 1-propene
 - 1-pentene
 - 2-hexene
 - 4-octene.
- Identify** each of the following compounds as an alkane or an alkene. Provide a name for each of the compounds.
 - C_4H_{10}
 - C_9H_{20}
 - $CH_3CH_2CH_2CH_3$
 - $CH_3-CH_2-CH_2-CH_2-CH_2-CH_3$
 - $CH_2=CH-CH_2-CH_2-CH_3$
 - $CH_3-CH=CH-CH_3$
 - $CH_3-CH_2-CH=CH-CH_3$

8 **Identify** what is wrong with the following names. Give the correct names.

- a 5-Hexene
- b 4-Heptene

ANALYSING

9 In the laboratory, the labels have come off four bottles containing colourless and transparent liquids. The labels are found on the floor and have the names: hexene (C_6H_{12}), pentane (C_5H_{12}), water (H_2O) and ethanol (C_2H_5OH).

A group of chemistry students was asked to devise a series of tests to identify the contents of each of the four bottles. The students labelled the bottles A–D and conducted their tests. Their results are shown in the following table.

Bottle	Soluble in water	Flammable	Decolourises bromine in the absence of UV light
A	Yes	No	No
B	No	Yes	Yes
C	No	Yes	No
D	Yes	Yes	No

Using the information from the tests, **identify** which bottle contained which chemical, giving reasons for your decision.

7.5 Comparing different substances

The structure and bonding of substances affects their properties. For example, the lattice structure and strength of ionic bonds give ionic compounds their relatively high melting and boiling points. In contrast, the intermolecular bonds of covalent molecular substances lead to melting and boiling points that are lower than those of ionic compounds.

Understanding the differences in structure and bonding of substances can help us explain the differences in these properties.



Worksheet
Comparing properties of
different substances

PRACTICAL ACTIVITY 7.5.1

COMPARING DIFFERENT TYPES OF SUBSTANCES

Introduction

Solids can be divided into four categories on the basis of their properties. Solids can be metallic, ionic, covalent molecular and covalent network. The property of electrical conductivity effectively distinguishes between metallic, ionic and covalent solids. To distinguish between covalent molecular and covalent network types, we need to consider melting point.

In this experiment, you will classify substances on the basis of their ability to conduct electricity in the solid, molten and/or aqueous states.

Note: Because of the amount of equipment and time involved, this experiment is best done by rotating through stations, each of which tests the properties of one substance.

Research question

How does the type of bonding of a substance affect its ability to conduct electricity in different forms?

Aim

To classify substances according to their physical properties

Materials

All stations

- power pack, kit for measuring conductivity (Stations 1–5 need 2 kits.)
- paper towel

Stations 1 and 2

- 4 crucibles with lids, Bunsen burner, tripod, pipe clay triangle
- station 1: sodium hydroxide pellets (enough to half-fill 2 crucibles)
- station 2: silver nitrate crystals (enough to half-fill 2 crucibles). Platinum electrodes work best at this station because of contamination of the electrodes in silver nitrate.

Stations 3–5

- tongs
- 6 crucibles with lids, hot plate
- station 3: candle wax (enough to half-fill 2 crucibles)
- station 4: sulfur (enough to half-fill 2 crucibles). Place this station in a fume cupboard.
- station 5: camphor or naphthalene (enough to half-fill 2 crucibles). Place this station in a fume cupboard.

Stations 6–12

- distilled water
- 100 mL beaker
- station 6: 50 mL distilled water
- station 7: 50 mL kerosene
- station 8: 50 mL ethanol
- station 9: 50 mL of 0.1 mol L⁻¹ sucrose solution
- station 10: 50 mL of 0.1 mol L⁻¹ NaCl solution
- station 11: 50 mL of 0.1 mol L⁻¹ NaOH solution
- station 12: 50 mL of 0.1 mol L⁻¹ HCl solution

Station 13

- 1 piece each of copper sheet, tin foil, aluminium foil and quartz



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Hydrochloric acid and sodium hydroxide are corrosive.	Wear safety glasses, gloves and protective clothing. Take care when pouring. Clean up spills immediately; if spilled on skin, wash affected area with plenty of water and notify your teacher immediately.
Silver nitrate is toxic.	Wear appropriate safety gear. Dispose of the chemical waste in the jar provided.
Substances may splash out of heated crucible.	Wear appropriate safety gear; use tongs.
Sulfur and naphthalene are irritants to eyes, nose and throat.	Work in a fume cupboard. Do not breathe in the fumes.
Kerosene is flammable and can be an irritant.	Place kerosene well away from Bunsen burners and hotplates. Work in a well-ventilated area; do not breathe the fumes.



Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

- 1 At all stations, start the power pack at 2 V. If you get a reading, record it. If not, turn the voltage up one step at a time until you get a reading, or get to a maximum of 6 V. Record the result, either a current reading or no current.
- 2 At stations 1 and 2, place one of the crucibles on the tripod and heat it with a blue flame until the contents melt. Turn the Bunsen burner off or to a very low flame. Use one of the conductivity kits to test the molten substance. Use the other kit to test the solid in the other crucible. Clean the electrodes between each test.
- 3 At stations 3–5, there is a chance of the substance catching fire. If it does, place the lid on the crucible to put it out. Use one of the conductivity kits to test the molten substance. Use the other kit to test the solid in the other crucible. Clean the electrodes between each test.
- 4 At stations 6–12, dip the conductivity kit into the solution and record the result. Rinse the electrodes with distilled water between each test.
- 5 At station 13, place the electrodes against each sample.

Results

Record your results in a table similar to the following.

Chemical	State	Conductivity	Classification

Analysis of results

- 1 What pattern, if any, is there to the conductivity results?

Interpretation

- 2 Use your knowledge of theory to identify the types of bonding in each of the substances tested. Compare your results with the theoretical classifications and suggest any reasons for differences.
- 3 What generalisations can you make about particles in substances and conductivity?



Practical activity

Comparing the properties of an ionic compound with those of its component elements

LEARNING CHECK 7.5

APPLYING

- 1 **Discriminate** between ionic bonding and metallic bonding.
- 2 **Discriminate** between covalent molecules, giant covalent networks and allotropes of carbon.

ANALYSING

- 3 An element X reacts with another element Y to form a compound with the formula XY_2 . Some properties of X, Y_2 and XY_2 are listed below.

Substance	Melting point ($^{\circ}\text{C}$)	Electrical conductivity	
		Solid	Molten material
X	1018	High	High
Y_2	88	None	None
XY_2	935	None	High

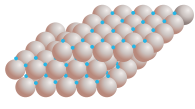
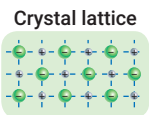
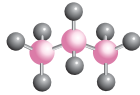
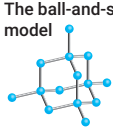
- a **Deduce** the type of particles that move when a voltage is applied to:
 - i solid X
 - ii molten XY_2 .
 - b **Justify** the relatively high melting point of XY_2 .
 - c **Infer** which bonds or forces are breaking when solid Y_2 melts.
- 4 Five solids A–E have the properties listed below. The relevant properties of sodium chloride and copper are also given.

Solid	Melting point ($^{\circ}\text{C}$)	Relative conductivity		Soluble in		'Hammer' test
		Solid	Liquid	water?	hexane?	
A	327	5	2	No	No	Flattens
B	2030	0	0	No	No	Shatters
C	91	0	0	No	Yes	Flattens
D	734	0	0.2	Yes	No	Forms powder
E	2870	0	0	No	No	Shatters
NaCl	801	0	0.2	Yes	No	Forms powder
Cu	1083	60	4	No	No	Flattens

Note: The hammer test describes what occurs when the material is continually hit with a hammer.

- a **Classify** each of the solids A–E as ionic, covalent molecular, covalent network or metallic.
- b **Explain** why sodium chloride and copper have the conductivity properties listed in the table.
- c For either covalent molecular compounds or covalent network compounds, **explain** why they have the melting points, conductivities and 'Hammer' test results shown in the table.

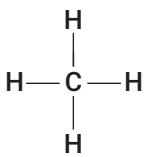
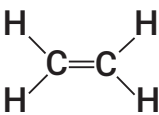
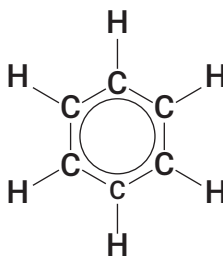
Ionic, covalent and metallic compounds

Property/ characteristic	Metallic	Ionic	Covalent molecular	Giant covalent network
Representative unit	Atoms  Key ● Metal ion • Valence electron	Formula unit  Crystal lattice	Molecule 	Molecule  The ball-and-stick model This model shows the geometry more clearly.
Type of elements	Metallic	Metallic combined with non-metallic	Non-metallic	Non-metallic
Bond	Involves electrons becoming delocalised around a lattice of cations	Electron transfer from metal atom to non-metal atom to form an ionic bond	Atoms share localised electrons to form covalent bonds	Covalent bonds continue throughout whole structure
What the strength of bond depends on	Number of electrons that become delocalised, charge of cation, size of cation	Number of electrons to lose, size of ion, distance to nucleus of other ion	Bond energy between atoms, number of electron pairs in the bond	Number of network bonds around each atom
Physical state	Solid, liquid	Solid	Solid, liquid or gas	Solid
Melting and boiling points	Variable but most commonly high	High	Low	High
Electrical conductivity	Good conductor in solid and liquid states	Non-conductor in solid state but good conductor in molten state and in solution	Does not conduct	Does not conduct (except graphite)
Thermal conductivity	High	Low	Low	Good
Hardness and malleability	Hard and malleable	Hard and brittle	Mostly soft	Hard and brittle
Solubility	Generally insoluble	Generally soluble in water	Variable – depends on solvent	Insoluble



Property/ characteristic	Metallic	Ionic	Covalent molecular	Giant covalent network
Examples	Sodium (Na), magnesium (Mg), aluminium (Al)	Sodium chloride (NaCl), magnesium nitrate ($\text{Mg}(\text{NO}_3)_2$), aluminium oxide (Al_2O_3)	Methane (CH_4), water (H_2O), carbon dioxide (CO_2)	Carbon in the form of diamond (C), silicon carbide (SiC), silicon dioxide (SiO_2)

Hydrocarbons

Property/ characteristic	Alkane	Alkene	Benzene
Bond type	Single bonds  Methane	At least one double bond  Ethene	Cyclic double bonds 
General formula	$\text{C}_n\text{H}_{2n+2}$	C_nH_{2n}	C_6H_6
Appearance	Colourless and odourless	Colourless with a faint petrol odour	Colourless with a strong odour
Melting and boiling points	All have low melting and boiling points.		
Conductivity	All are non-conductors of electricity and heat.		
Solubility	All are insoluble in water but soluble in some other solvents		
Chemical properties	Relatively inert Undergo combustion Undergo substitution reactions with halogens Will stay brown when reacted with bromine water	Undergo combustion Undergo addition reactions with hydrogen, halogens and water Undergo polymerisation Will go colourless when reacted with bromine water	Undergo substitution reactions with alkyl groups, halogens and nitro groups
Examples	Methane – gas Ethane – gas Propane – gas Butane – gas Pentane – liquid	Ethene – gas Propene – gas Butene – gas Pentene – liquid	Benzene – liquid

MULTIPLE CHOICE

- A particular element has two electrons in its valence shell. This element will form:
 - a metallic lattice when pure, and covalently bonded molecules when combining with chlorine.
 - an ionic lattice when pure, and a covalent lattice when bonding to chlorine.
 - an ionic lattice when pure, and covalently bonded molecules when combining with chlorine.
 - a metallic lattice when pure, and an ionic lattice when combining with chlorine.
- Solid sodium fluoride (NaF) melts at 992°C , whereas magnesium oxide (MgO) melts at 2800°C .
Which of the following is *not* a reasonable statement about the compounds MgO and NaF?
 - There is ionic bonding between the particles in the lattice of each solid.
 - There are stronger forces holding the particles together in MgO than in NaF.
 - Electrons were shared between atoms before lattices of the solids were formed.
 - The electron configuration of the ions accounts for the different values of melting temperature of the two solids.
- At room temperature, carbon disulfide is likely to:
 - consist of triatomic molecules with weak bonds between the molecules.
 - be a liquid that has high electrical conductivity.
 - have ionic bonds and be hard but brittle.
 - be a covalent network solid with a high melting temperature.
- Which of the following substances conduct electricity as solids?
 - Cu and Al_2O_3
 - Cu, Al and graphite
 - Cu, CuCl_2 , Al and Hg
 - Cu, CuCl_2 , H_2O , Al, Hg and CH_4
- Which of the following substances conduct electricity as molten liquids?
 - Cu and CH_4
 - Cu, H_2O and CH_4
 - Cu, CuCl_2 and Al
 - Cu, CuCl_2 , H_2O , Al, Hg and CH_4
- How many hydrogen atoms are there in 2-pentene?
 - 5
 - 9
 - 10
 - 12
- Which of the substances H_2O , AlCl_3 , SO_2 , NH_3 , Cu, NaBr, Li_2O and Cl_2 are covalent molecular substances?
 - AlCl_3 , NaBr and Li_2O
 - H_2O , SO_2 , NH_3 and AlCl_3
 - H_2O , SO_2 , NH_3 and Cl_2
 - H_2O , AlCl_3 , SO_2 , NH_3 , Cu, NaBr, Li_2O and Cl_2

8. A sample of substance Z was analysed and found to have a low density and a low melting point. Z was insoluble in water and did not conduct electricity in either the liquid or solid form. The type of bonding in substance Z is probably:
- A ionic.
 - B metallic.
 - C covalent network.
 - D covalent molecular.
9. What does a line in a structural formula of a hydrocarbon represent?
- A One electron
 - B Two electrons
 - C Four electrons
 - D Unpaired electrons
10. Which of the following hydrocarbons can be classified as saturated?
- | | |
|------------|------------------|
| A C_2H_6 | C C_6H_6 |
| B C_4H_8 | D $C_{14}H_{30}$ |

SHORT RESPONSE

11. Phosphorus trichloride is a liquid with a boiling point of 76°C ; it does not conduct electricity. Calcium chloride is a solid with a melting point of 772°C ; when molten, it conducts electricity. **Explain**, in terms of bonding, why these compounds have such different properties.
12. Compound A is a liquid at room temperature. It does not conduct electricity. When heated, it turns to a gas at 78°C . It can be easily condensed back to the original liquid.
- a What evidence is there that the liquid is not mercury metal?
 - b Give a reason why the liquid cannot be an ionic material.
 - c What is the bonding in the substance? **Explain** your answer.

CROSS-CHAPTER QUESTION

13. Third ionisation energies of 10 consecutive elements in the periodic table are given in the following table. The symbols do not necessarily represent any of the symbols used in the periodic table.

Element	Third ionisation energy (kJ mol^{-1})
A	3659
B	4619
C	4577
D	5301
E	6045
F	6276
G	6912
H	7732
I	2744
J	3229

- a Identify** the element that is most likely to exist as a diatomic molecule with a double bond.
- b i Determine** the formula for the compound formed between the elements E and G.
- ii Identify** the type of bonding in the resulting compound.

DATA ANALYSIS

- 14.** The physical properties of six solid substances (U–Z) were investigated by experimentation. Examine the data below and then **classify** each of the six substances into one of the following groups:

- metallic substance
- ionic substance
- covalent molecular substance
- covalent network substance.

Briefly **explain** how you classified the substances using your understanding of bonding and evidence from the table.

Solid	Appearance	Melting point (°C)	Solubility in water	Electrical conductivity		
				Solid	Molten	Aqueous solution
U	Lustrous silvery	419	Insoluble	Good	Good	N/A
V	Red crystals	727	Soluble	Nil	Good	Good
W	White waxy crystals	51	Insoluble	Nil	Nil	N/A
X	Lustrous silvery white	232	Insoluble	Good	Good	N/A
Y	White crystal	Decomposes	Soluble	Nil	N/A	Good
Z	Grey-black	2700	Insoluble	Nil	N/A	N/A

N/A, not applicable.

SCIENCE AS A HUMAN ENDEAVOUR

Syllabus dot point

- Consider the benefits and potential risks associated with the use of nanomaterials in consumer products, health care, transportation, energy and agriculture.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Nanotechnology in cosmetics

The nanoscale

Nanoparticles are particles with diameters in the range 1–100 nanometres (nm). To put this in perspective, the smallest objects that can be seen with the naked eye are about a millimetre in size (10^6 nm, Figure 1). Optical microscopes can distinguish an individual bacterium at the micrometre scale (1000 nm, Figure 1). The resolution of electron microscopes allows the study of viruses, antibodies and molecular structures.

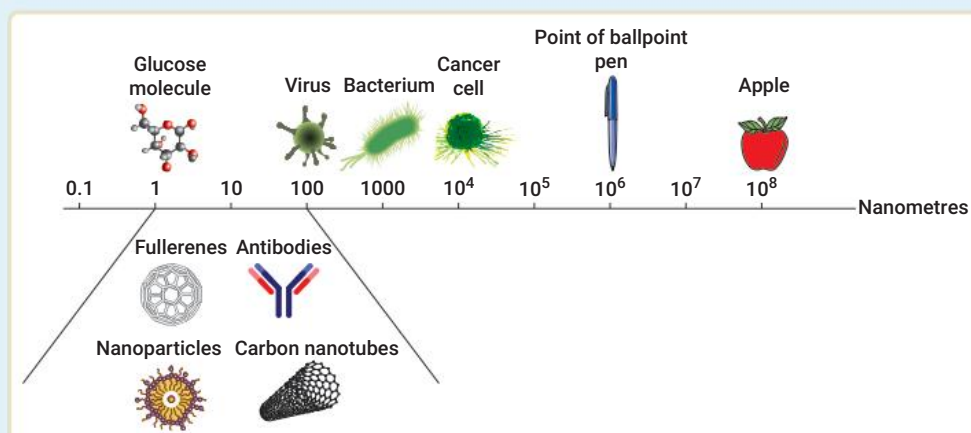


FIGURE 1 A log scale of objects in nanometres; the nanomaterials are the fullerenes, antibodies, nanoparticles and carbon nanotubes.

Their small size means that nanoparticles have a large surface area-to-volume ratio, which is useful in many applications.

Nanomaterials

Fullerenes are a nanomaterial (Figure 2) that were discovered by Richard Smalley and co-workers, using laboratory synthesis and mass spectrometry, in 1985. Fullerenes were also found to occur naturally in 1992 by Peter Buseck and co-workers in unusual carbonaceous rock in Russia. The first fullerene discovered was made up of 60 covalently bound carbon atoms in a nearly spherical arrangement of 32 faces – 12 pentagons and 20 hexagons – like a soccer ball. Since this initial discovery, scientists have identified fullerenes that contain between 20 and 84 covalently bound carbon atoms.

Because of their unique structure and properties, fullerenes have been used in a range of applications, including drug delivery, electronics and sensors.

An application of nanomaterials – cosmetics

In 1986, cosmetics companies Lancôme and Dior launched the first nanotechnology-based products under the names of Niosomes® and Capture®, respectively.



FIGURE 2 A model of the C_{60} fullerene showing the hexagon and pentagon faces

Science Photo Library/Russell Kightley

Nanomaterials are used in the cosmetic industry to produce nanocosmetics with many benefits, including:

- prolonged action of the cosmetic by controlling the delivery of active ingredients
- site specificity
- improvements in biocompatibility
- the ability to control the amount of ingredient delivered.

However, as with anything, there are limitations of using nanomaterials, including stability, scalability, toxicity, cost and many more. The safety and toxicity of nanomaterials are debatable. The major concern in the advancement of nanocosmetics is that they are small enough to allow the active ingredients to reach the blood, which could lead to increased toxicity within the body. **Figure 3** shows some advantages and disadvantages of using nanocosmetics.

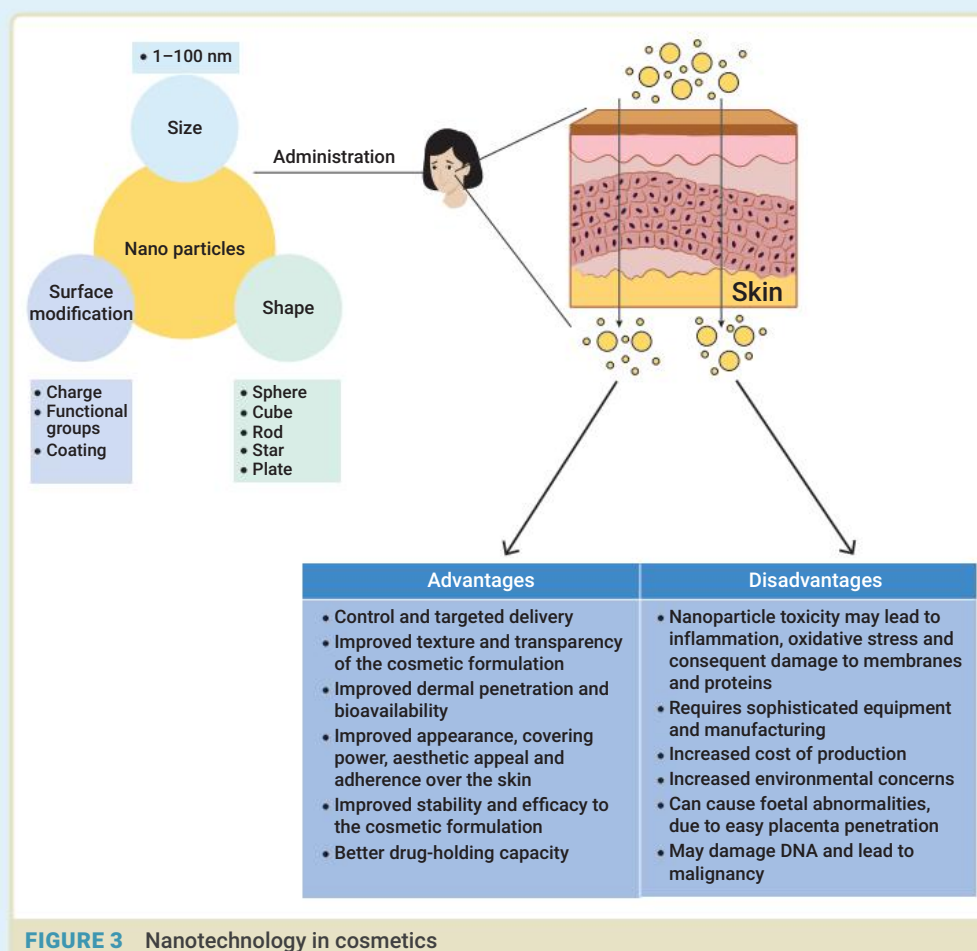


FIGURE 3 Nanotechnology in cosmetics

Currently, nanotechnology is regarded as a promising and revolutionising field and is being used in many industries. However, further research needs to be conducted into the toxicity of these particles and related health issues.

Reference

Gupta, V., Mohapatra, S., Mishra, H., Farooq, U., Kumar, K., Ansari, M.J., Aldawsari, M.F., Alalaiwe, A.S., Mirza, M.A., Iqbal, Z. (2022). 'Nanotechnology in cosmetics and cosmeceuticals – a review of latest advancements. *Gels*. 8(3), 173. (<https://www.mdpi.com/2310-2861/8/3/173>)

Chemical reactions



Albert Russ/Shutterstock.com

SYLLABUS DOT POINTS

SCIENCE UNDERSTANDING

- 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 simple redox reactions.

SCIENCE INQUIRY

Investigate:

- types of chemical reactions.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Chemical reactions involve the breaking of bonds in reactants as well as interactions between atoms, ions and molecules to form new atomic, ionic and molecular substances as products. In this topic, you will learn how to describe chemical reactions using balanced chemical equations. A chemical reaction can either involve the production of energy (heat) or require heat for it to occur.

Substances can also change from solid to liquid, liquid to gas, or vice versa. These are not chemical reactions as such but are called phase changes. Phase changes can also involve temperature changes; for example, when ice melts, it absorbs energy. The phase is whether the substance is in solid, liquid or gaseous form.

The five basic types of chemical reactions are combination, decomposition, single-replacement, double-replacement and combustion. Analysing the reactants and products of a given reaction will allow a reaction to be placed into one of these categories.

Practicals

- Displacement reactions involving metals
- Reaction of metal with oxygen
- Classifying chemical reactions (online-only resource)

Worksheets

- Chemical or physical reactions
- Balancing types of chemical reactions



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ The states of matter include solids, liquids and gases.
- ✓ Compounds can be represented by chemical formulas.
- ✓ The type of compound can be identified from its chemical formula.
- ✓ Chemical reactions may occur when substances are combined.
- ✓ Chemical reactions start with reactants and end with products.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ identify that physical and chemical changes involve energy changes
- ✓ identify the main indicators that a chemical reaction has occurred
- ✓ identify the reactants and products in a chemical equation
- ✓ write word equations for described reactions
- ✓ determine balanced chemical equations
- ✓ use the correct symbols to identify the states of matter of substances in a chemical reaction
- ✓ discriminate between a spontaneous and non-spontaneous reaction
- ✓ identify types of chemical reactions.

8.1 Physical and chemical change

Chemical change

chemical reaction
the formation of new substances and the associated energy transformations

Chemical reactions are important in our lives, in our homes and in industry. Chemical reactions occur every time we light a gas stove or a barbecue. Steel rusting in a car body is an example of a chemical reaction. Cooking food causes the ingredients to change chemically and form new substances. In industry, new materials are always being developed, and they require chemical reactions to change the starting materials into new materials. WikiCell is an edible form of packaging developed at Harvard University in the US. Based on gelatin and sugarcane fibre, WikiCell may substantially decrease our need for plastic food packaging.

Some materials, such as liquefied natural gas (which is mostly methane), require very little chemical processing to make them useful. Yet most materials are produced by chemical reactions, such as paints, fuels, textiles, hard finishes for cupboards and benchtops, paper and cardboard, electronic devices, components of household appliances, cleaning agents, and pharmaceuticals.

Photosynthesis, respiration, rusting and combustion all involve chemical reactions. Photosynthesis occurs in plants in the presence of light (**Figure 8.1.1**). Respiration occurs in both plants and animals all the time. Iron rusts when it is exposed to oxygen and moisture. Petrol burns in the combustion chamber of car engines. In each case, chemical changes occur.

All chemical reactions involve the formation of new substances and associated energy transformations, which scientists commonly observe as changes in the temperature of the surroundings, the emission of light, or both.

123RF/firina



FIGURE 8.1.1 Green leaves and stems – the site of photosynthesis, a reaction that builds molecules that store energy. Photosynthesis is more important to life on Earth than any other chemical process. It is the largest chemical ‘factory’ on the planet!

The main signs that can indicate that a chemical process has occurred are:

- solid precipitates out from a solution or a substance is consumed
- a gas or vapour is produced (evolved)
- a colour change takes place
- there is a change in smell
- the temperature changes or light is given off
- it is difficult to reverse what has happened.

Spontaneity of chemical reactions

Some chemical reactions require the addition of energy whereas others don't. **Spontaneous chemical reactions** occur without the ongoing input of energy. Spontaneous reactions that take place at room temperature include the ripening of fruit, salt dissolving in water and the rusting of iron. Substances that are more reactive are more likely to undergo reactions spontaneously at room temperature. Some spontaneous reactions occur quickly and others occur more slowly, but all occur without the input of additional energy.

In contrast, **non-spontaneous chemical reactions** require external agents such as heat, light or electricity. Photosynthesis is a good example of a non-spontaneous reaction. A plant absorbs water and carbon dioxide from its surroundings. But for these reactants to be converted into glucose and oxygen, sunlight is required.

spontaneous chemical reaction a reaction that occurs without the input of additional energy such as heat, light or electricity

non-spontaneous chemical reaction a reaction that can only occur with the input of additional energy such as heat, light or electricity

Chemical equations

A chemical reaction can be represented by a chemical equation, which can use words or formulas. In a chemical equation, the reactants (the substances that exist before a chemical change (or reaction) takes place) are written to the left of the arrow. The products (the new substances formed) are written to the right of the arrow. The single arrow always points in the direction of the reaction; i.e. from the reactants to the products.

For example, the word equation for photosynthesis is:



Carbon dioxide and water are the reactants. Glucose and oxygen are the products. The arrow points towards the products. The '+' is used when there is more than one reactant or more than one product. Energy changes associated with the chemical reaction can also be included in the chemical equation. This is discussed in Chapter 10.



Syllabus link
Chapter 10 discusses energy changes associated with a chemical reaction.

Phase change

A **phase change** is also referred to as a 'change of state' or a physical change. Melting, freezing, vaporisation (evaporation and boiling), condensation, sublimation and deposition are all changes of state, involving energy but not involving chemical reactions.

Familiar examples of a phase change are water freezing (liquid to solid) and ice melting (solid to liquid), water boiling (liquid to vapour), and steam condensing and forming water droplets (vapour to liquid) (**Figure 8.1.2**). On hot days, rain water lying on a road surface evaporates as the rain water undergoes a phase change from liquid to vapour. If you leave moth balls (made from the organic chemical naphthalene) lying in a cupboard, they slowly disappear because they change state directly from solid to vapour, a process called 'sublimation'. On very cold mornings, we often see grass and leaves becoming frosty. Frost is ice crystals that form when warm moist air comes into contact with cold surfaces. The formation of frost is an example of a gas changing state directly to a solid, bypassing the liquid phase; this process is known as 'deposition'.

phase change the transformation of a substance between solid, liquid or gaseous forms



Worksheet
Chemical or physical reactions

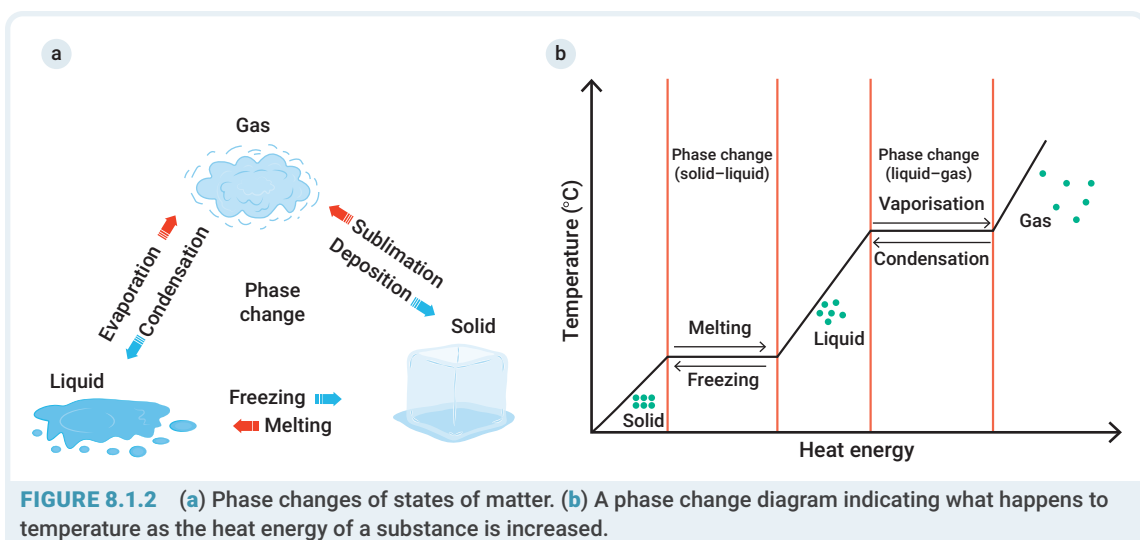
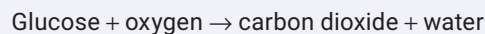


FIGURE 8.1.2 (a) Phase changes of states of matter. (b) A phase change diagram indicating what happens to temperature as the heat energy of a substance is increased.

WORKED EXAMPLE 8.1.1

Respiration is the combustion reaction of glucose and oxygen in body cells to form carbon dioxide and water. The word equation is:



Identify the:

- a reactants
- b products.

ANSWERS

- a **Identify the substances to the left of the arrow.**
Glucose and oxygen are the reactants because they are to the left of the arrow.
- b **Identify the substances to the right of the arrow.**
Carbon dioxide and water are the products because they are to the right of the arrow.

LEARNING CHECK 8.1

DESCRIBING

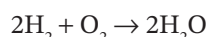
- 1 **Define:**
 - a chemical reaction
 - b phase change
 - c spontaneous reactions.

APPLYING

- 2 Write word equations and **identify** the reactants and products for the following reactions.
 - a Rust (iron(III) oxide) forms when iron reacts with oxygen in the air.
 - b Petrol is composed of many hydrocarbon compounds, including octane. Octane reacts with oxygen in the combustion chamber of cars to form carbon dioxide and water.
- 3 When water is added to baking soda, the mixture starts to bubble. **Identify** whether the reaction that is occurring is a chemical reaction.

8.2 Balanced chemical equations

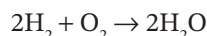
Chemists normally represent chemical reactions using chemical symbols and formulas for the reactants and products. Numbers (or coefficients) are written in front of chemical formulas to balance the equation so that the same number of atoms of each element is shown on each side. We do not write the number '1' because it is assumed when no other number appears before the formula. You cannot change the subscripts in chemical formulas in an equation because it will change the chemical you are using. For example, in this simple chemical reaction:



the '2' in front of hydrogen and water is a coefficient. This tells us that two molecules of hydrogen are required to react with one molecule of oxygen to produce two molecules of water.

Generally, equations are written as the lowest whole number ratio between reactants and products. This ensures that the law of conservation of mass is maintained.

The law of conservation of mass refers to Antoine Lavoisier's 1789 discovery that mass is neither created nor destroyed in chemical reactions, although it may be rearranged in space. Simply put, the number of atoms that you start with on the reactant side equals the number of atoms that are produced, and the products will be different substances from the reactants. For example, using the simple equation again:



there are four atoms of hydrogen on the reactant side and four on the product side. The coefficient in front of the formula acts as a multiplier. There are two atoms of oxygen on the reactant side and two atoms on the product side. The number of atoms is equal on both sides of the reaction and no new atoms have been introduced to the product side.

In chemical reactions, the states for each of the reactants and products are often included:

- (s) for solid
- (l) for liquid
- (g) for gas
- (aq) for aqueous solution.

In some sources, (aq) is not included; if a reactant or product does not have a state included, it is assumed that it is an aqueous solution (dissolved in water).

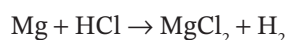
Chemists use balanced chemical equations to summarise information about chemical reactions. These equations communicate information concisely about reactants, products and numbers of atoms of each element present.

Steps to balance chemical equations:

1. Write the unbalanced equation.
2. Count the number of atoms on each side of the reaction.
3. Add coefficients to make the numbers of atoms equal (coefficient $\times$ subscript = number of atoms).
4. If necessary, reduce the coefficients to the lowest possible ratio (never change the subscripts).
5. Update and double-check the atom balance until complete (remember the law of conservation of mass).

For example, as seen in **Figure 8.2.1**, magnesium (Mg) reacts with hydrochloric acid (HCl) to form magnesium chloride (MgCl_2), releasing bubbles of hydrogen gas.

The gas can be identified as hydrogen (H_2) by using a flame to cause a 'pop'. The reaction can be described by the equation:



Syllabus link
Chapter 9
discusses the law
of conservation
of mass.



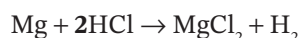
FIGURE 8.2.1 Magnesium reacting with hydrochloric acid

Science Photo Library/Martyn F. Chillmaid

The equation then needs to be balanced by altering the coefficients. Equations must be balanced to achieve the same number of atoms of each element on each side of the equation. A summary of the number of atoms is shown in the following table.

Element	Reactant side	Product side
Mg	1	1
H	1	2
Cl	1	2

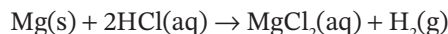
There is an unequal number of hydrogen and chlorine on the two sides of the equation. To increase the number of hydrogen and chlorine on the reactant side, we need to add an additional HCl, resulting in the equation:



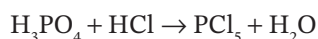
This is shown in the following table:

Element	Reactants side	Products side
Mg	1	1
H	2	2
Cl	2	2

With the equation balanced, we now need to add the states for each of the reactants and products to complete the equation for the reaction. Magnesium is a solid (s); hydrogen is a gas (g). Both hydrochloric acid and magnesium chloride are dissolved in water (aq).



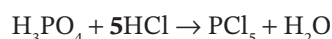
The method for balancing (by adjusting the coefficients) is the same for complex reactions. For example, phosphoric acid (H_3PO_4) reacts with hydrochloric acid (HCl) to form a white solid of phosphorus pentachloride (PCl_5) and water (H_2O). The reaction can be described by the equation:



An initial count of the atoms shows:

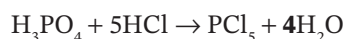
Element	Reactant side	Product side
H	4	2
P	1	1
O	4	1
Cl	1	5

The equation needs to be balanced by changing coefficients. This time it will not be completed in one step. Start by balancing the chloride atoms by adding the coefficient 5 to hydrochloric acid:



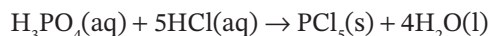
Element	Reactant side	Product side
H	8	2
P	1	1
O	4	1
Cl	5	5

There are now five chloride atoms on each side of the equation, but the number of hydrogen atoms has also changed. Note that there is hydrogen in both phosphoric and hydrochloric acid, giving a total of eight hydrogen atoms on the reactant side. Balance these by adding the coefficient 4 to the water on the product side:



Element	Reactant side	Product side
H	8	8
P	1	1
O	4	4
Cl	5	5

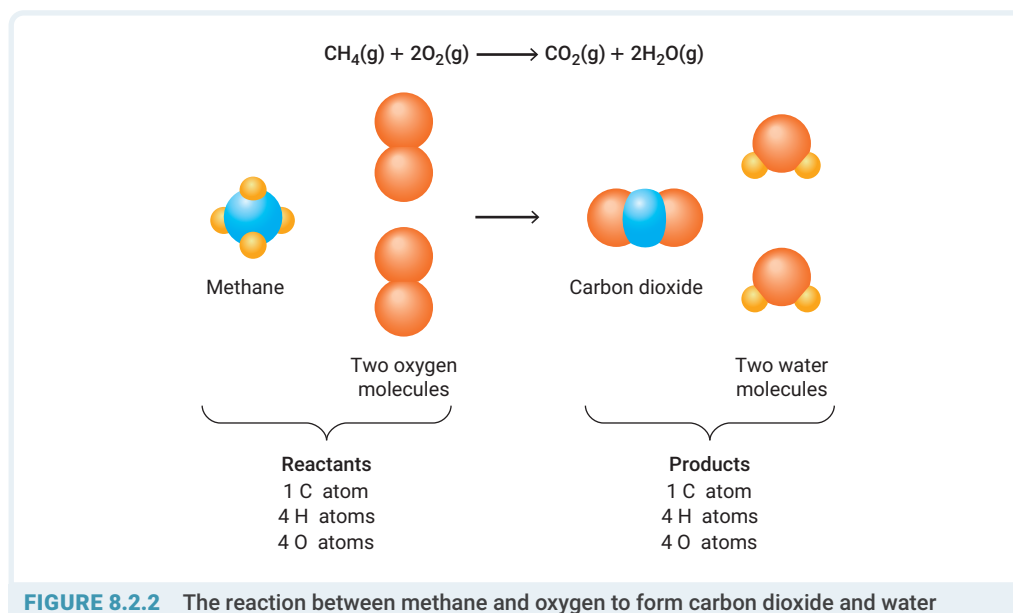
The chloride and hydrogen atoms are now equal on both sides of the equation. Next check the number of phosphorus and oxygen atoms. There is one phosphorus and four oxygen atoms on each side of the reaction. The equation is completely balanced. Now add the states for each of the reactants and products to complete the equation:



Balancing more complex equations can be confusing, especially considering there is no 'rule' to follow. However, keeping a constant check on the number of atoms on the reactant side and product side can be very helpful.

The combustion of methane

Methane is the main constituent of liquefied natural gas and has the molecular formula CH_4 . The combustion of methane in air involves the reaction with oxygen (O_2) to produce carbon dioxide (CO_2) and water vapour (H_2O). The balanced reaction is shown in **Figure 8.2.2**.



The reaction between methane and oxygen is shown as a chemical equation, but also as pictures showing the constituent atoms and shapes of the molecules. The reactants are on the left side of the equation, and the products are on the right. The equation is balanced – the numbers of C, H and O atoms are the same on both sides of the equation.



Weblink
Balancing chemical equations

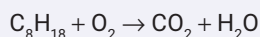
Worksheet
Balancing types of chemical reactions

WORKED EXAMPLE 8.2.1

Write a balanced chemical equation, including states, for the reaction of octane (C_8H_{18}) with oxygen in a car's combustion chamber to produce carbon dioxide and water. Note that the combustion chamber is at a high temperature so, in this case, the water formed will be gaseous until it condenses after cooling through the exhaust pipe.

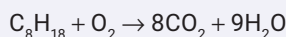
ANSWER

1 Write reactants and products using their chemical formulas.

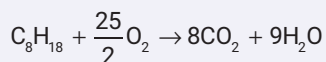


2 Balance the equation.

- a** Start with the elements that appear in one substance (choose C_8H_{18}) and balance atom numbers on each side of the equation for that substance and the atoms it contains. That is, balance only the C and H atoms to start because C_8H_{18} contains only C and H ($8 \times \text{C}$ and $18 \times \text{H}$), so $8 \times \text{CO}_2$ and $9 \times \text{H}_2\text{O}$ are needed on the right to balance the atom numbers:



- b** Consider the O atom numbers. There are now $(8 \times 2) + (9 \times 1) = 25$ atoms of oxygen on the right and two atoms of oxygen on the left. O_2 is a molecule with two O atoms, so the fraction $\frac{25}{2}$ must go in front of the O_2 , on the left, to balance the 25 O atoms on the right.

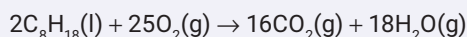


Note that, when balancing equations, whole numbers are used, so all quantities in the equation are doubled in this example, so 50 O atoms end up on both left and right sides, as needed.



Each side of the equation now has the same number of 'atoms' for each element.

3 Add the states for each of the reactants and products.



Octane is a liquid; oxygen and carbon dioxide are gases. This is a combustion reaction, so water will be produced at a high temperature and therefore will be gaseous.

LEARNING CHECK 8.2

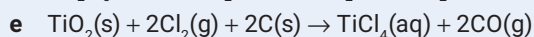
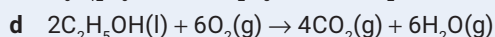
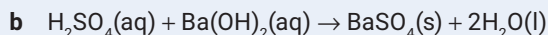
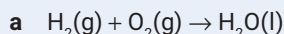
DESCRIBING

- 1 Identify** the purpose of a chemical equation.
- 2 Identify** the steps required to write balanced chemical equations.

APPLYING

- 3** Balance the following chemical equations.
- $\text{Fe}(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2(\text{g})$
 - $\text{NaHCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - $\text{CuFeS}_2(\text{s}) + \text{SiO}_2(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{Cu}_2\text{S}(\text{s}) + \text{FeSiO}_3(\text{s}) + \text{SO}_2(\text{g})$
 - $\text{PbS}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{PbO}(\text{s}) + \text{SO}_2(\text{g})$
 - $\text{C}_3\text{H}_8(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$

4 **Identify** whether the following equations are balanced. Correctly balance the equations that are not balanced and **justify** your answer.



5 Write balanced chemical equations for the:

a combustion of cyclohexane (C_6H_{12}) to yield carbon dioxide and water vapour

b rusting of iron when it reacts with oxygen to form solid iron(III) oxide

c respiration of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) with oxygen to produce carbon dioxide and water

d Haber process – an important chemical reaction in which hydrogen gas reacts with nitrogen gas (N_2) to produce gaseous ammonia (NH_3)

e reaction of magnesium hydroxide ($\text{Mg}(\text{OH})_2$) (an ingredient in common antacids) with dilute hydrochloric acid (HCl) in the stomach to produce a solution of magnesium chloride (MgCl_2) and water

f reaction of zinc metal with dilute sulfuric acid (H_2SO_4) to form dilute zinc sulfate (ZnSO_4) solution and hydrogen gas.

8.3 Types of chemical reactions

Chemicals react to achieve stability, i.e. the products are more stable than the reactants. During chemical reactions, bonds in the reactants are broken, and new bonds are formed in the products. Energy is needed to break bonds, which are the forces of attraction that occur between:

- metal ions and the delocalised electrons in metallic lattices
- cations and anions in ionic lattices
- non-metallic elements in covalent molecules and covalent lattices.

Energy is released when bonds are formed. Whether the reaction results in an increase or decrease in temperature depends on the difference between the energy input to break the bonds and the energy released when bonds are formed.

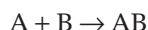
Types of chemical reactions

We will investigate five types of chemical reactions in the following section:

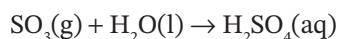
- combination reactions
- decomposition reactions
- combustion reactions
- single displacement reactions
- double-displacement reactions.

Combination reactions

Combination reactions are reactions in which two or more substances combine to form a single product. The reactants may be elements or compounds, but the product is always a single compound. The general equation for a combination reaction is:



An example of a combination reaction is the reaction of sulfur trioxide and water to form sulfuric acid:



Syllabus link

Chapter 10 goes into further detail about energy changes in chemical reactions.



Practical activity

Classifying chemical reactions

combination reaction a reaction in which two or more reactants combine to form a single product

decomposition reaction

a reaction in which a single substance breaks down into smaller, simpler substances

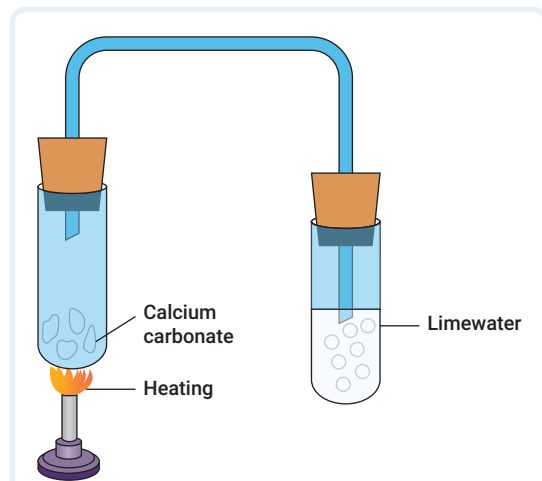


FIGURE 8.3.1 When calcium carbonate is heated, it decomposes into solid calcium oxide and carbon dioxide gas. Bubbling carbon dioxide through limewater produces a solid. This is a positive test for carbon dioxide.

combustion reaction

an exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that produces energy

single-displacement reaction

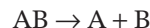
a reaction in which an element within a compound is displaced by another element



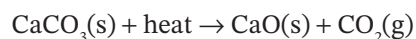
FIGURE 8.3.2
Zinc reacting with hydrochloric acid

Decomposition reactions

Decomposition reactions are reactions in which a single substance breaks down into two or more simpler substances. There is always just a single reactant in a decomposition reaction. Products of decomposition reactions may be elements such as O_2 and H_2 or simple compounds such as H_2O and CO_2 . Most require energy in the form of heat, light or electricity to proceed. The general equation for decomposition reactions is:



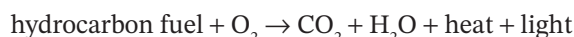
An example of a decomposition reaction is the breakdown of calcium carbonate on heating (**Figure 8.3.1**):



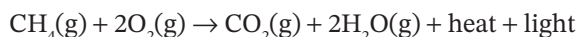
Combustion reactions

Combustion reactions are reactions in which an element or a compound reacts rapidly with oxygen gas to liberate heat and light energy. Commonly, the compounds combining with oxygen in these reactions are hydrocarbons, which are wholly made up of hydrogen and carbon. The well-known combustible fuels kerosene and petrol, for instance, are hydrocarbon mixtures. The complete combustion of a hydrocarbon yields carbon dioxide and water as the reaction products. If insufficient oxygen is available, combustion will not be complete and carbon monoxide and elemental carbon

may be obtained as additional products of the reaction. The general unbalanced equation for a combustion reaction involving hydrocarbon fuels is:

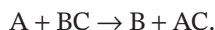


An example of a combustion reaction is the burning of methane gas to give water (in the form of steam), carbon dioxide, heat and light:

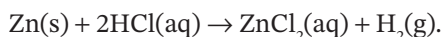


Single-displacement reactions

Single-displacement reactions are reactions in which an element within a compound is displaced by another element. In simpler terms, an element in a reactant compound is 'swapped' for a different element from another reactant. This type of reaction always has two reactants, one of which is always an element and will always result in two products, one of which is always an element. The general equation of a single-displacement reaction is:



An example of a single-replacement reaction is the reaction of zinc metal with hydrochloric acid (**Figure 8.3.2**):



The hydrogen from the compound HCl is replaced by zinc, resulting in the formation of hydrogen gas. Single-displacement reactions often involve metals displacing another element from a compound.

PRACTICAL ACTIVITY 8.3.1

DISPLACEMENT REACTIONS INVOLVING METALS

Introduction

Metals have the ability to react with different substances.

Research question

Do different metals react differently with hydrochloric acid and water?

Aim

To identify metals that would take part in displacement reactions with water or HCl

Equipment per group

- 3 × 5 cm × 0.5 cm pieces of copper
- 3 approx. 5 cm long ungalvanised iron nails
- 3 × 5 cm pieces of magnesium
- 3 × 5 cm × 0.5 cm pieces of zinc
- rice-sized piece of calcium metal
- 50 mL of 1 mol L⁻¹ hydrochloric acid (HCl)
- 5 test tubes and a test-tube rack
- spatula
- distilled water
- forceps
- paper towel
- PPE



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Acid could splash or spill onto your hands.	Wash your hands immediately if they come in contact with acid and after using acids.
Vigorous bubbling may splash into your eyes.	Wear safety glasses.
When calcium comes in contact with water, a flammable gas is released.	Handle calcium in a well-ventilated area to avoid formation of explosive air/hydrogen mixtures.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

Part A: Reacting metals with water

- 1 Place a rice-sized piece of calcium into a test tube.
- 2 Place a piece of each of the other metals into separate test tubes.
- 3 Add approximately 3 cm depth of water to each test tube.
- 4 Record your observations.
- 5 Using forceps, remove any unreacted metal and place it on a piece of paper towel.
- 6 Dispose of any remaining water and solutions and rinse the test tubes.

Part B: Reacting metals with 1 mol L⁻¹ hydrochloric acid

- 1 For any metals that did not violently react with water, place a small quantity or piece of each into separate clean test tubes.
- 2 Add approximately 3 cm depth of 1 mol L⁻¹ HCl to each test tube.
- 3 Record your observations.
- 4 Using forceps, remove any unreacted metal and place it on a piece of paper towel.
- 5 Dispose of any remaining water and solutions. Do not pour any solids down the drain.

Results

Draw a table to show your results.

Analysis of results

- 1 In which of the experiments did a chemical reaction occur?
- 2 Of the metals that reacted, which resulted in a displacement reaction with HCl?
- 3 Write a balanced equation for any chemical reaction that occurred in parts A and B.

Interpretation

- 4 What evidence indicated that a reaction had taken place?

double-displacement reaction a reaction in which a positive ion from one ionic compound exchanges with the positive ion of another ionic compound



Web link

Double-displacement reactions using matchmaker models

precipitation reaction a reaction between two solutions that results in the formation of a solid

neutralisation the reaction between an acid and a base to form a salt and water



FORMULA AND
DATA BOOK



Syllabus link

Chapter 15 discusses precipitation reactions in further detail.

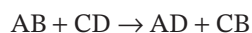


Syllabus link

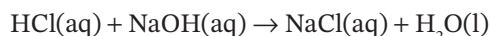
Chapter 17 discusses acid-base (neutralisation) reactions in further detail.

Double-displacement reactions

Double-displacement reactions are reactions in which a positive ion from one ionic compound exchanges with the positive ion of another ionic compound. These reactions typically occur in aqueous solution and result in the formation of a precipitate (a solid), the production of a gas, or the formation of a molecular compound such as water. The general equation for a double-displacement reaction is:



An example of a double-replacement reaction is the reaction that occurs between aqueous hydrochloric acid and sodium hydroxide:



Science photo Library/GipfotoSTOCK



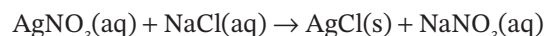
FIGURE 8.3.3 Both silver nitrate and sodium chloride are clear aqueous solutions. When mixed, a white precipitate of silver chloride is formed, which eventually settles on the bottom of the beaker.

Special types of double-replacement reactions

Precipitation reactions

In a precipitation reaction, one of the products is insoluble and forms a precipitate and the other is soluble and forms an aqueous solution. To tell if a compound is soluble, use the solubility rules. A precipitate is a solid that forms and separates from a liquid solution as the result of a chemical reaction.

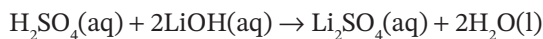
An example of a precipitation reaction is the reaction that occurs between aqueous silver nitrate and aqueous sodium chloride (**Figure 8.3.3**). A precipitate of solid silver chloride is formed in this reaction:



Acid-base (neutralisation) reactions

When an acid (consisting of H^+ and X^- aqueous ions) and a base (consisting of M^+ and OH^- ions) react, the cations and anions switch partners, resulting in the formation of water and a new ionic compound (or salt), which is usually soluble. Neutralisation reactions are generally accompanied by a noticeable release of heat.

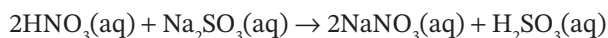
For example, in the reaction between sulfuric acid and aqueous lithium hydroxide, the predicted products are lithium sulfate (salt) and water:



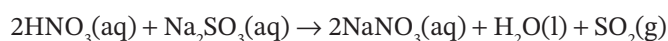
Gas-forming reactions

In these reactions, one of the products after the double displacement is in the gaseous state, such as hydrogen sulfide (H_2S) or ammonia (NH_3). One of the products could also be carbonic acid (H_2CO_3) or sulfurous acid (H_2SO_3).

For example, in the reaction of nitric acid with aqueous sodium sulfite, the predicted products are sodium nitrate and sulfurous acid:



However, sulfurous acid decomposes to sulfur dioxide and water, resulting in a final equation of:



Simple redox reactions

A reduction–oxidation or **redox reaction** involves the transfer of electrons between chemical species (the atoms, ions or molecules involved in the reaction). Redox reactions will be studied in detail in Unit 3.

Redox reactions are all around us: the combustion of fuels, the corrosion of metals (**Figure 8.3.4**), and even the processes of photosynthesis and cellular respiration involve oxidation and reduction. During a redox reaction, some species undergo oxidation or the loss of electrons, while others undergo reduction, or the gain of electrons. Some specific examples of redox reactions are:

- $2\text{Cu}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{CuO}(\text{s})$ (oxidation of copper)
- $4\text{Fe}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$ (corrosion of iron)
- $\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightarrow 2\text{HF}(\text{g})$

redox reaction short for reduction–oxidation reaction, a type of reaction that involves transfer of electrons between two species; one reactant is oxidised (loses electrons) and the other reactant is reduced (gains electrons)

123RF/ Jose Angel Astor



FIGURE 8.3.4 Metals such as iron in cars react with oxygen in the air to corrode.



Syllabus link
Redox reactions are covered in more detail in Unit 3.

PRACTICAL ACTIVITY 8.3.2

REACTION OF METAL WITH OXYGEN

Introduction

Corrosion is the spontaneous reaction of a metal with oxygen in a moist environment.

Research question

To what degree do different metals react with oxygen in a moist environment?

Aim

To compare the reactivity of metals by comparing their reactions in a moist environment

Equipment per group

- 5 cm × 0.5 cm piece of copper
- 5 cm × 0.5 cm piece of zinc
- 5 cm × 0.5 cm piece of aluminium foil
- 5 cm piece of magnesium
- 5 cm iron nail (ungalvanised)
- 5 medium test tubes and a test-tube rack
- 5 cm × 5 cm piece of sandpaper



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Metal residue may remain on your hands.	Wash your hands after handling metals.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

- 1 Clean each of the metals with the sandpaper.
- 2 Place a piece of each metal into separate test tubes.
- 3 Place the test tubes into the test-tube rack.
- 4 Add water to each of the test tubes until the piece of metal is covered.
- 5 Record observations daily for one week.

Results

Record your observations in an appropriate format. For example, you could take a series of photos of each test tube.

Analysis of results

- 1 Write a balanced chemical equation for any reaction that has occurred.

Interpretation

- 2 Which metal was the most corroded? Justify your answer.
- 3 Which metal was the least corroded? Justify your answer.
- 4 Rank the metals from most corroded to least corroded. Justify your answer.

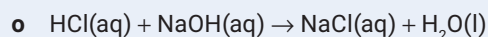
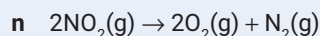
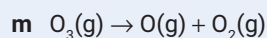
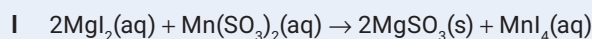
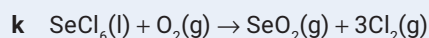
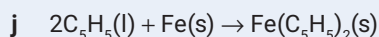
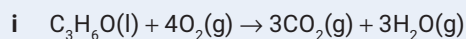
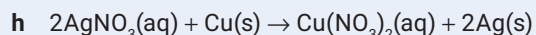
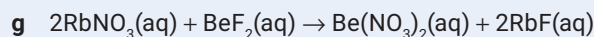
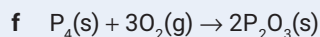
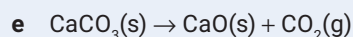
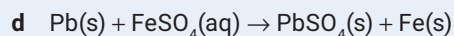
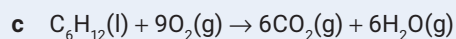
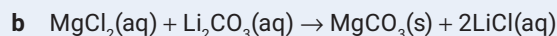
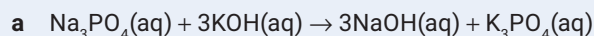
LEARNING CHECK 8.3

DESCRIBING

- 1 **State** two tests to identify gases produced during a chemical reaction.

APPLYING

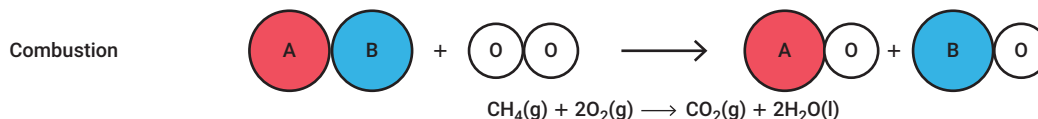
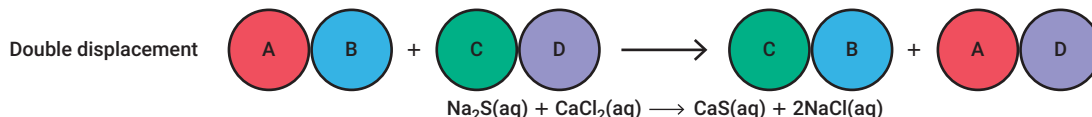
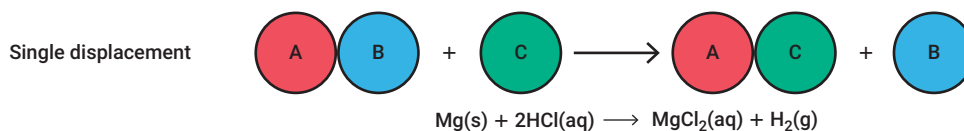
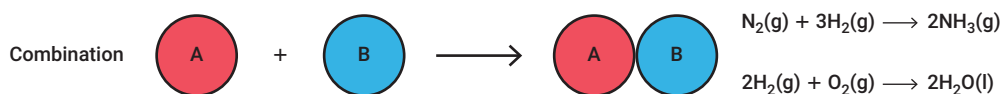
- 2 **Classify** the following chemical reactions (combination, decomposition, single-displacement, double-displacement or combustion).



CHAPTER SUMMARY

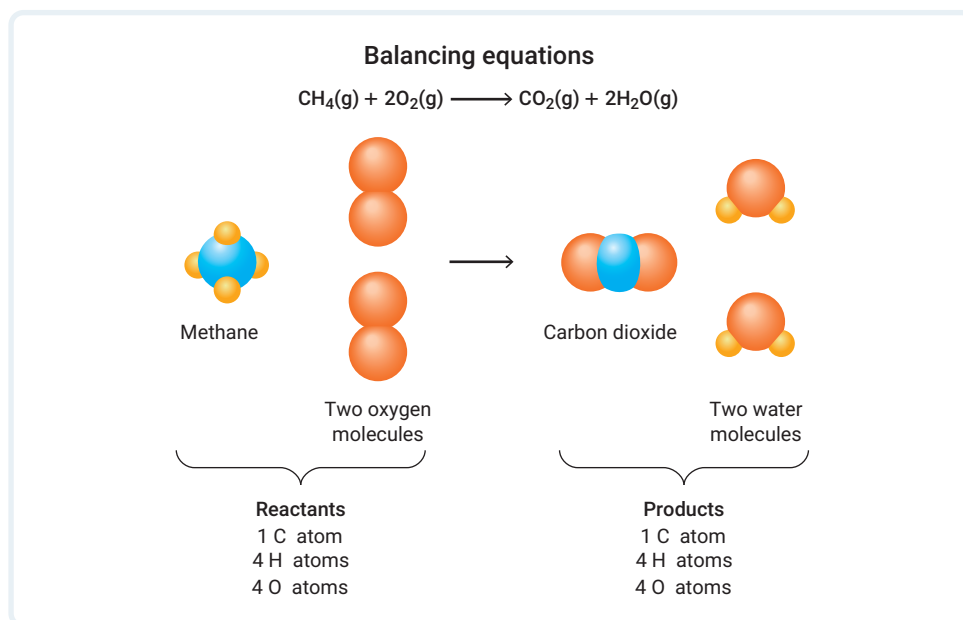
Chemical reactions and phase changes

- The main signs that can indicate that a chemical reaction has occurred are:
 - a solid precipitates from a solution or a substance is consumed
 - a gas or vapour is produced (evolved)
 - a colour change takes place
 - there is a change in smell
 - there is an input or release of energy, indicated by temperature changes or light being given off
 - it is difficult to reverse what has happened.
- Phase change is a physical process – change of state: melting, freezing, vaporisation, condensation, sublimation and deposition.
- Chemical equations show reactants on the left-hand side of the arrow and products on the right-hand side of the arrow: reactants → products
- The states for each of the reactants and products are often included in the chemical equation:
 - (s) for solid
 - (l) for liquid
 - (g) for gas
 - (aq) for aqueous solution.
- Spontaneous chemical reactions occur naturally without the need for continuous external energy input once they have started, whereas non-spontaneous chemical reactions do not occur naturally without the input of external energy.
- The main types of chemical reactions are combination, decomposition, combustion, single-displacement and double-displacement reactions. There are some other special types of reactions, which include acid–base neutralisation reactions, precipitation reactions and redox reactions.



Writing balanced chemical equations

- Steps to balance chemical equations:
 - 1 Write the unbalanced equation.
 - 2 Count the number of atoms on each side of the reaction.
 - 3 Add coefficients to make the numbers of atoms equal (coefficient $\times$ subscript = number of atoms)
 - 4 If necessary, reduce coefficients to the lowest possible ratio (never change subscripts).
 - 5 Update and double-check atom balance until complete (remember the law of conservation of mass).

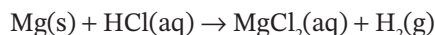


CHAPTER EXAM

MULTIPLE CHOICE

- Which of the following is not an example of a chemical reaction?
A Respiration
B Lighting a match
C An iron nail rusting
D Sugar crystals dissolving in water
- A student performed the pop test by holding a lit taper over the top of a closed test tube containing hydrogen gas. When the gas was released, a loud 'pop' was heard and small clear, colourless droplets of liquid were formed at the mouth of the test tube. These were found to be water. What is the correct word equation to represent this reaction?
A Hydrogen $\rightarrow$ water
B Hydrogen $\rightarrow$ water + sound
C Hydrogen + oxygen $\rightarrow$ water
D Hydrogen + heat $\rightarrow$ water + sound
- What does (aq) mean when written in a chemical equation?
A Water is a reactant in the chemical reaction.
B Water is a product in the chemical reaction.
C The substance is dissolved in water.
D The substance directly before it is in solution.

- Consider the following unbalanced chemical equation:



Which of the following shows the correct coefficients for the balanced chemical equation?

- 1 : 1 : 1 : 1
 - 1 : 2 : 1 : 2
 - 1 : 2 : 1 : 1
 - 2 : 4 : 2 : 2
- In a balanced chemical equation, what do the coefficients represent?
A The energy change during the reaction
B The mass of each reactant and product
C The number of molecules of each reactant and product
D The volume of gases involved at standard temperature and pressure
 - The reaction between sodium and water can be written as:



Which of the following equations is the correct balanced chemical equation for the reaction?

- $\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{NaOH(aq)} + 2\text{H}_2\text{(g)}$
 - $2\text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$
 - $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$
 - $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + 2\text{H}_2\text{(g)}$
- Iron(III) oxide reacts with aluminium to produce aluminium oxide (Al_2O_3) and metallic iron. Which of the following is the correct balanced chemical equation for the reaction?
A $\text{Fe}_2\text{O}_3\text{(s)} + \text{Al(s)} \rightarrow \text{Fe(s)} + \text{Al}_2\text{O}_3\text{(s)}$
B $\text{Fe}_2\text{O}_3\text{(s)} + \text{Al}_2\text{(s)} \rightarrow \text{Fe}_2\text{(s)} + \text{Al}_2\text{O}_3\text{(s)}$
C $\text{Fe}_2\text{O}_3\text{(s)} + 3\text{Al(s)} \rightarrow 2\text{Fe(s)} + 3\text{AlO(s)}$
D $\text{Fe}_2\text{O}_3\text{(s)} + 2\text{Al(s)} \rightarrow 2\text{Fe(s)} + \text{Al}_2\text{O}_3\text{(s)}$

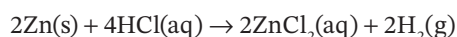
Questions 8 and 9 refer to the following table where the tick means a reaction has occurred.

Metal	Reacts with water	Reacts with dilute acid	Reacts with oxygen
Ag			✓
Ca	✓	✓	✓
Al		✓	✓
Pb		✓	✓
Cu			✓
Mg		✓	✓
Au			
K	✓	✓	✓

8. Aluminium can take part in a single-displacement reaction with dilute acid. A single-displacement reaction can be described as a reaction that:
- A releases oxygen.
 - B involves reactants exchanging elements.
 - C occurs between a metal and a compound.
 - D involves reactants joining together to produce one product.
9. According to the results, which metal is unable to take part in a combustion reaction?
- A Au
 - B Ca
 - C Cu
 - D Mg
10. If energy does not need to be added to a reaction to make it occur at room temperature, then it can be referred to as a:
- A reversible reaction.
 - B combustion reaction.
 - C spontaneous reaction.
 - D non-spontaneous reaction.

SHORT RESPONSE

11. **Explain** how this equation should be changed in order for it to be correctly balanced.



12. Write balanced chemical equations for each of the following reactions. Include the correct state for each of the reactants and products.
- a The reaction of nitrogen with hydrogen to form ammonia (gas)
 - b The decomposition of iron(II) sulfate (FeSO_4) to yield Fe_2O_3 , SO_2 and SO_3
 - c The combustion of propane (C_3H_8) to produce carbon dioxide and water vapour
 - d The combustion of ethanol ($\text{C}_2\text{H}_5\text{OH}$) to produce carbon dioxide and water vapour
 - e Photosynthesis: the reaction of carbon dioxide and water to produce $\text{C}_6\text{H}_{12}\text{O}_6$ (glucose) and oxygen
 - f Copper(II) nitrate ($\text{Cu}(\text{NO}_3)_2$) is heated in a crucible over a Bunsen burner, forming solid copper(II) oxide (CuO), nitrogen dioxide (NO_2) and oxygen gas
 - g A sodium hydroxide solution is added to a magnesium chloride solution. Solid magnesium hydroxide is formed with another clear solution.

CROSS-CHAPTER QUESTION

13. An aqueous solution of calcium chloride is mixed with an aqueous solution of sodium bicarbonate (NaHCO_3) to produce solid calcium carbonate, an aqueous solution of sodium chloride, water and carbon dioxide.
- Determine** the type of compound each of the reactants and products are, i.e. ionic, covalent or metallic compound.
 - Determine** the general properties of each type of compound, including melting and boiling point, electrical conductivity and hardness.
 - Write a balanced chemical equation for this reaction.
 - When this reaction occurs, a precipitate is formed. **Explain** how the precipitate is considered evidence that a chemical reaction has occurred.
 - Identify** the type(s) of chemical reaction(s) that have occurred.

DATA ANALYSIS

14. **Apply understanding**

A student uses two substances in a series of tests. The observations from one test are shown below.

	Substance A	Substance B
Observations	No colour change Changed from a solid to a liquid Liquid was colder than the original solid.	A funny smell was apparent after the test. Some bubbling occurred, but no colour change.

Based on the observations above, **identify** if either substance A or B took part in a chemical reaction. Give reasons for your answer.

15. **Analyse evidence**

A student investigates the reaction of different metals with hydrochloric acid. The acid and metal are placed into a beaker for the reaction to occur. The student conducts the experiment three times with different metals. They measure the temperature change in each experiment. The following table summarises the results.

Metal	Temperature change ($^{\circ}\text{C}$)		
	Trial 1	Trial 2	Trial 3
Magnesium	10.3	10.5	10.2
Zinc	8.6	8.7	5.6
Iron	5.2	4.9	5.1

- Identify** an anomalous result in the results table.
 - Suggest a reason why the result could be anomalous.
- Calculate** the average temperature change for each of the metals used.
- The student concludes that the more reactive a metal is, the greater the temperature change. **Describe** further tests that could be conducted to confirm their conclusion.

The mole concept and the law of conservation of mass



Sergii Molchenko/Shutterstock.com

SYLLABUS DOT POINTS

SCIENCE UNDERSTANDING

- 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. $\left(\text{Formula: mole } (n) = \frac{\text{mass } (m)}{\text{molar mass } (M)} \right)$
- 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.

$$\left(\text{Formula: Percentage yield } (\%) = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1} \right)$$



SCIENCE INQUIRY

Investigate:

- limiting reactant/s and percentage yield.
- the empirical formula of a compound from reactions involving mass change.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Calculating chemical quantities requires the use of measurements that are related to atoms and molecules, their masses, how they are distributed in space (or in a volume), and how that distribution can be quantified. Chemical changes that occur as a result of reactions also need to be measured and identified.

Practicals

- Determining the empirical formula of magnesium oxide
- Investigating limiting reactants
- Measuring molar quantities (online-only resource)

Worksheets

- Mass, moles and molar mass
- Formula and molecular masses
- Percentage composition
- Limiting reactants

 **Nelson MindTap**

To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ The relative atomic mass of an element is an average of the mass of an atom (relative to the mass of a carbon-12 atom)
- ✓ The relative atomic mass of an element can be identified from the periodic table.
- ✓ Chemical formulas of ionic and covalent compounds can be written.
- ✓ The empirical formula illustrates the lowest whole number ratio of the formula for a compound.
- ✓ In balanced equations, the total number of particles of each element is the same on both the reactant and product sides of the equation.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ state the Avogadro constant
- ✓ identify the relationship between the Avogadro constant and the mole
- ✓ convert between moles and number of particles
- ✓ identify the relationship between moles and chemical formulas
- ✓ state the law of conservation of mass
- ✓ identify the relationship between moles/mass and chemical reactions
- ✓ determine the relative mass and molar mass of a substance
- ✓ convert between moles and mass
- ✓ calculate the percentage composition of a substance
- ✓ distinguish between empirical and molecular formulas
- ✓ determine the empirical formula of a substance and deduce the molecular formula
- ✓ calculate mole–mole and mass–mass using stoichiometry
- ✓ identify the limiting reactant and excess reactant in a chemical reaction.
- ✓ calculate the percentage yield of a chemical reaction.

9.1 The mole and Avogadro's number of particles

Individual atoms and molecules are too small to see even with the most specialised equipment. For this reason, chemists devised a unit that is practical for counting and handling a standard number of particles.

The Avogadro constant

Chemists deal with amounts of substances in terms of their masses, so they use a relationship between mass and number of particles.

As we discovered in Chapter 1, the mass of an atom is its **relative atomic mass** (represented by the symbol A_r) and is calculated relative to the mass of a carbon-12 atom. One **atomic mass unit (amu)** is equal to one-twelfth the mass of one atom of carbon-12. You may recall from Chapter 2 that the atomic mass values given in the periodic table are the weighted averages of all of the isotopes of an element. Chemists were able to work out that one carbon-12 atom, with a relative atomic mass of 12 amu, has an actual mass of $1.992\,65 \times 10^{-23}$ grams. Using this

relative atomic mass (A_r)
the average mass of one atom of an element, relative to the mass of a carbon-12 atom

atomic mass unit (amu)
one-twelfth of the mass of an atom of carbon-12

information, it was determined that there are approximately 6.022×10^{23} atoms of carbon in 12 g of carbon-12. This value is called the **Avogadro constant (N_A)** in honour of Amedeo Avogadro, a 19th-century Italian scientist.

Figure 9.1.1 shows a sample of carbon powder weighing exactly 12.00 g. Therefore, we are looking at a sample that must contain 6.022×10^{23} atoms of carbon! The Avogadro constant is a scaling factor between macroscopic and microscopic (atomic scale) observations of nature.

The Avogadro constant is a number so large, it is difficult to comprehend. If you had the same number of oranges as the Avogadro constant, they would form a sphere the size of Earth. Alternatively, this number of grains of rice would cover Australia to a depth of approximately 1 kilometre.

The relationship of the Avogadro constant to the mole

Just as it is convenient to group sheets of paper into reams (500 sheets) and eggs into cartons of a dozen, chemists measure the amount of substances in a batch size called the mole (mol) and represent the number of moles by the symbol n .

Chemists have chosen 1 mol as a standard unit for large numbers of atoms, ions or molecules. A mole is the amount of substance containing N_A particles of that substance. For most calculations, it is sufficiently precise to take $N_A = 6.02 \times 10^{23}$. The **mole** is the SI base unit representing the chemical quantity of a substance.

Converting between moles and number of particles

The relationship between the number of moles (n) of a substance and the number of particles (atoms, ions or molecules) is given by:

$$\text{Number of moles } (n) = \frac{\text{number of particles}}{\text{number of particles per mole}}$$

where the number of particles per mole = 6.02×10^{23} .

Substituting in the formula, we can see that:

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

Conversion between number of moles and number of particles, or vice versa, can be carried out using this relationship, as shown in **Figure 9.1.2**.

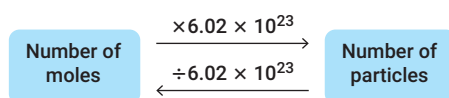


FIGURE 9.1.2 Converting between the number of moles and number of particles



FIGURE 9.1.1 A mass of 12 g of powdered carbon – a couple of teaspoons in volume

Avogadro constant (N_A) the number of units of one mole of substance, approximated as 6.022×10^{23}

mole (n) the amount of a substance (g) containing $N_A = 6.022 \times 10^{23}$ particles (atoms, ions or molecules) of that substance



Weblink
How big is a mole?

WORKED EXAMPLE 9.1.1

Calculate the number of moles of magnesium in 1.45×10^{23} atoms of magnesium.

ANSWER

To convert from particles to moles, divide the number of particles given by N_A :

$$\begin{aligned}n(\text{Mg}) &= \frac{1.45 \times 10^{23}}{6.02 \times 10^{23}} \\&= 0.241 \text{ mol}\end{aligned}$$

The number of moles of magnesium is 0.241 mol.

WORKED EXAMPLE 9.1.2

Calculate the number of moles of methane (CH_4) molecules represented by 4.50×10^{22} methane molecules.

ANSWER

To convert from particles to moles, divide the number of particles given by N_A :

$$\begin{aligned}n(\text{CH}_4) &= \frac{4.50 \times 10^{22}}{6.02 \times 10^{23}} \\&= 0.0748 \text{ mol}\end{aligned}$$

The number of moles of methane is 0.0748 mol.

WORKED EXAMPLE 9.1.3

Calculate the number of atoms in 0.25 mol of neon atoms.

ANSWER

To convert from moles to particles, multiply the number of moles given by N_A :

$$\begin{aligned}\text{Number of atoms of Ne} &= 0.25 \times 6.02 \times 10^{23} \\&= 1.5 \times 10^{23} \text{ atoms}\end{aligned}$$

There are 1.5×10^{23} neon atoms.

WORKED EXAMPLE 9.1.4

Calculate the number of molecules in 5.6 mol of water.

ANSWER

To convert from moles to particles, multiply the number of moles given by N_A :

$$\begin{aligned}\text{Number of molecules} &= 5.6 \times 6.02 \times 10^{23} \\&= 3.4 \times 10^{24} \text{ molecules}\end{aligned}$$

There are 3.4×10^{24} molecules of water.

As you can see in the Worked examples, the number of particles can refer to both single atoms and molecules.

Moles and chemical formulas

With an understanding of the mole concept, we can start to appreciate the importance of being able to write chemical formulas. As we know, the formulas of elements and compounds contain subscripts that represent the numbers of atoms of each element in a molecule or formula unit of the substance. For example, in one molecule of water (H_2O), there are two atoms of hydrogen and one atom of oxygen. In two molecules of water, there are double the number of hydrogen (4) and oxygen (2) atoms, and so on.

The subscripts also represent the number of moles of each atom in one mole of molecules of the substance. So, if there were 1 mol of water molecules, there would be 2 mol of hydrogen atoms and 1 mol of oxygen atoms. Similarly, if there were 2 mol of water molecules, then there would be 4 mol of hydrogen atoms and 2 mol of oxygen atoms.

This relationship also applies to ionic compounds. One mole of sodium sulfate (Na_2SO_4) formula units contains 2 mol of sodium ions (Na^+) and 1 mol of sulfate ions (SO_4^{2-}). One mole of sulfate ions is made up of 1 mol of sulfur atoms and 4 mol of oxygen atoms.

WORKED EXAMPLE 9.1.5

- a Calculate the number of moles of carbon atoms and oxygen atoms in 5 mol of carbon dioxide (CO_2).
- b Determine the total number of atoms present in 5 mol of carbon dioxide.

ANSWERS

- a **1 Use the formula to determine the number of moles of carbon and oxygen atoms present.**

From the formula, you can see that 1 mol of CO_2 contains 1 mol C and 2 mol O.

- 2 Use the formula to calculate the number of mol of each element.**

For 5 mol CO_2 :

$$n(\text{C}) = 5 \times 1 \text{ mol C} = 5 \text{ mol of C}$$

$$n(\text{O}) = 5 \times 2 \text{ mol O} = 10 \text{ mol of O}$$

Therefore, there are 5 mol of C and 10 mol of O.

- b **1 Determine the total number of moles of atoms in 5 mol CO_2 .**

From part a, there are 5 mol of C and 10 mol of O, so the total number of moles of atoms = 15 mol

- 2 Calculate the total number of atoms.**

$$\text{Using: } n = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

$$\begin{aligned} \text{Total number of individual atoms} &= 15 \times 6.02 \times 10^{23} \\ &= 9.03 \times 10^{24} \end{aligned}$$

Therefore, there are 9.03×10^{24} individual atoms present.

LEARNING CHECK 9.1

DESCRIBING

- 1 Define:**

a the Avogadro constant

b mole

c atomic mass unit.

APPLYING

- 2 Explain** the relationship between the Avogadro constant and the mole.
- 3 Calculate** the number of moles of each of the following substances, given the number of particles.

- a 4.00×10^{24} atoms of helium
- b 1.03×10^{22} molecules of chlorine gas
- c 1.204×10^{23} atoms of copper
- d 8.22×10^{21} formula units of aluminium oxide
- e 3.3×10^{25} carbon atoms
- f 1.20×10^{22} molecules of water
- g 6.6×10^{21} sodium ions
- h 3.01×10^{24} formula units of calcium carbonate
- i 9.05×10^{23} molecules of hydrogen gas

4 **Calculate** the number of particles of each of the following substances, given the number of moles.

- a 2.5 mol of oxygen molecules
- b 0.330 mol of gold atoms
- c 1.6 mol of iron(II) ions
- d 0.035 mol of silver nitrate (AgNO_3)
- e 0.100 mol of methane (CH_4) molecules
- f 6.0 mol of hydrogen atoms
- g 0.5 mol of aluminium ions
- h 0.082 mol of argon gas
- i 2.10 mol of ammonia (NH_3) molecules

5 Copper(II) carbonate has the formula CuCO_3 . In 3 mol of CuCO_3 **identify** how many of each of the following there are.

- a Moles of Cu^{2+} ions
- b Moles of CO_3^{2-} ions
- c Moles of oxygen atoms
- d Individual atoms of oxygen

6 Octane (C_8H_{18}) is an important constituent of petrol. In 3.00×10^{26} molecules of octane, **identify** how many moles of each of the following are present.

- a C_8H_{18} molecules
- b C atoms
- c H atoms

7 **Identify** how many moles of oxygen atoms are in each of the following.

- a 2.5 mol of H_2SO_4 molecules
- b 0.2 mol of CH_3COOH molecules
- c 3.02×10^{24} formula units of $\text{Al}_2(\text{CO}_3)_3$

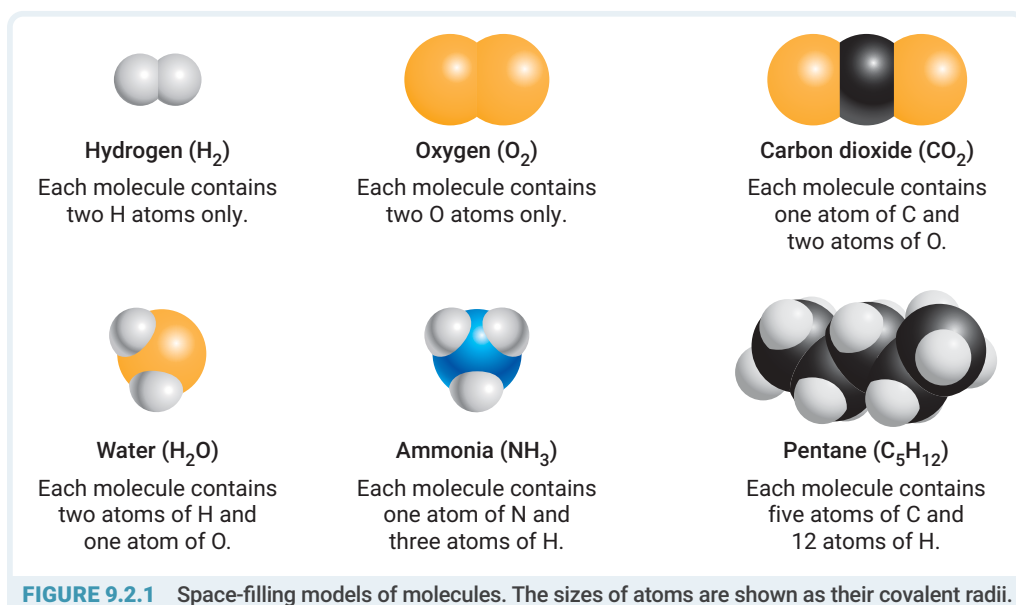
9.2 The mole concept

Relative mass of substances

Chemical symbols and formulas, such as He, Na, Cl_2 and H_2O , are short-hand representations for elements and compounds. Symbols and formulas may also represent a group of atoms or formula units. Just as relative atomic mass is used to describe the mass of an atom, **relative molecular mass (M_r)** is used to describe the mass of one molecule of a molecular substance. The relative molecular mass (M_r) of a substance is the mass of one molecule of the substance on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12 (units are atomic mass units).

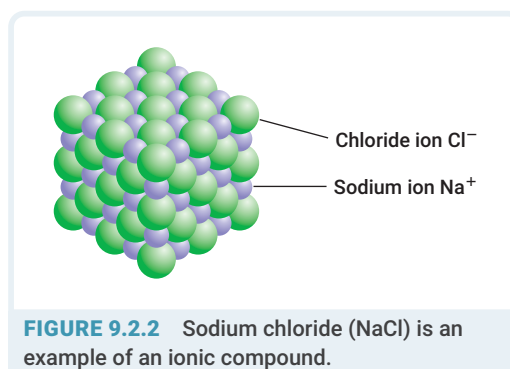
The relative molecular mass of a molecule, expressed in atomic mass units, is calculated by adding the relative atomic masses of all the component atoms of the molecule.

relative molecular mass (M_r) the mass of a molecular substance relative to the mass of a carbon-12 atom



Worksheet
Formula and molecular masses

Not all pure substances exist as molecules. Ionic compounds consist of a crystalline lattice of cations and anions combined in a definite fixed ratio (**Figure 9.2.2**), as indicated by the formula of the compound, such as Al_2O_3 . The term ‘relative molecular mass’ is incorrectly applied to ionic substances and **relative formula mass** should be used instead. This is also given the symbol M_r . The relative formula mass is the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12. The relative formula mass of a compound (M_r) is the sum of the relative atomic masses of the atomic species as given by the formula of the compound.



WORKED EXAMPLE 9.2.1

Calculate the relative molecular mass of ethane (C_2H_6) in amu.

ANSWER

1 Identify the number of each type of atom.

Number of C atoms = 2

Number of H atoms = 6

2 Multiply the number of each type of atom by their relative atomic mass (from the periodic table).

Mass of C = $2 \times A_r(\text{C}) = 2 \times 12.01 = 24.02$

Mass of H = $6 \times A_r(\text{H}) = 6 \times 1.01 = 6.06$

3 Add the masses of each element to obtain the relative mass of the compound.

$M_r(\text{C}_2\text{H}_6) = 24.02 + 6.06 = 30.08$

This can be simplified to:

$M_r(\text{C}_2\text{H}_6) = 2 \times A_r(\text{C}) + 6 \times A_r(\text{H}) = 2 \times 12.01 + 6 \times 1.01 = 30.08$

Therefore, the relative molecular mass is 30.08 amu.

Note: The relative atomic mass used when showing working in calculations should be the same as on the periodic table of elements in the *Formula and data book*.

relative formula mass (M_r)
the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 atom is exactly 12



FORMULA AND DATA BOOK

WORKED EXAMPLE 9.2.2

Calculate the relative formula mass of calcium hydroxide (Ca(OH)_2) in amu.

ANSWER

1 Identify the number of each type of atom.

Number of Ca atoms = 1

Number of H atoms = 2

Number of O atoms = 2

2 Multiply the number of each type of atom by its relative atomic mass (from the periodic table).

Mass of Ca = $1 \times A_r(\text{Ca}) = 1 \times 40.08 = 40.08$

Mass of H = $2 \times A_r(\text{H}) = 2 \times 1.01 = 2.02$

Mass of O = $2 \times A_r(\text{O}) = 2 \times 16.00 = 32.00$

3 Add the masses of each element to obtain the mass of the compound.

$$M_r(\text{Ca(OH)}_2) = 40.08 + 2.02 + 32.00 \\ = 74.10$$

This can be simplified to:

$$M_r(\text{Ca(OH)}_2) = 1 \times A_r(\text{Ca}) + 2 \times A_r(\text{H}) + 2 \times A_r(\text{O}) \\ = 1 \times 40.08 + 2 \times 1.01 + 2 \times 16.00 \\ = 74.10$$

The relative formula mass is 74.10 amu.

Note: The relative atomic mass used when showing working in calculations should be the same as what is on the periodic table of elements in the *Formula and data book*.

FORMULA AND
DATA BOOK

molar mass (M) mass of 1 mol of a substance (equals mass of $N_A = 6.022 \times 10^{23}$ atoms, molecules or formula units of a substance)

Moles and mass

Chemists often need to calculate the number of moles present in a given mass of a substance. The mass (in grams) of 1 mol of a particular element or compound is numerically equal to its relative atomic, molecular or formula mass. The relative formula mass, relating to a single atom, molecule or formula unit, is expressed in atomic mass units. The mass of 1 mol of a substance is called the **molar mass (M)**. The units of molar mass are grams per mole (g/mol or g mol^{-1}).

For example, the molar mass of magnesium (Mg) is 24.31 g mol^{-1} , the molar mass of oxygen gas (O_2) is $(2 \times 16.00) = 32.00 \text{ g mol}^{-1}$ and the molar mass of CO_2 is $(1 \times 12.01 + 2 \times 16.00) = 44.01 \text{ g mol}^{-1}$.

KEY CONCEPT

One mole of any substance has a mass equal to the relative atomic, molecular or formula mass of that substance, but expressed in grams. The molar mass (M) is an actual mass that can be measured in grams.



FIGURE 9.2.3 One mole each of five different substances. The mass of 1 mol varies depending on the substance.

Converting between moles and mass

If you know the amount of a given substance in grams, then you can calculate the number of moles and vice versa. For example, if you have 18.02 g of water, then you must have 1 mol of water because the molar mass is 18.02 g mol⁻¹. If you want to measure out 2 mol of water, then you would measure (2 × 18.02) = 36.04 g of water.

The relationship between the number of moles (n), mass (m) and the molar mass (M) of a substance is given by:

$$\text{Number of moles } (n) = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

$$n = \frac{m}{M}$$

This formula can be rearranged if you want to calculate the mass of a given number of moles:

$$m = n \times M$$

Alternatively, the conversion between mass and moles can be represented schematically as shown in **Figure 9.2.4**.

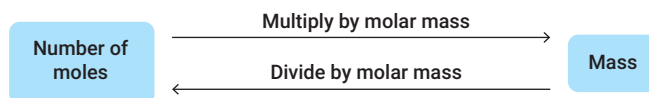


FIGURE 9.2.4 Converting between the number of moles and mass



Worksheet
Mass, moles and
molar mass

Practical activity
Measuring molar quantities

KEY FORMULA

$$\text{Number of moles } (n) = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$
$$n = \frac{m}{M}$$

WORKED EXAMPLE 9.2.3

Calculate the molar mass of 1 mol of ethane (C₂H₆).

ANSWER

Add the masses of each element in the compound.

$$\begin{aligned} M &= 2 \times A_r(\text{C}) + 6 \times A_r(\text{H}) \\ &= 2 \times 12.01 + 6 \times 1.01 = 30.08 \end{aligned}$$

$$M(\text{C}_2\text{H}_6) = 30.08 \text{ g mol}^{-1}$$

The molar mass of ethane is the same numerical value as its relative molecular mass.

WORKED EXAMPLE 9.2.4

Calculate the number of moles of copper(II) sulfate (CuSO₄) in 12.2 g of copper(II) sulfate.

ANSWER

1 Write the relationship of copper(II) sulfate.

$$n = \frac{m}{M}$$

$n = ?$, $m = 12.2 \text{ g}$, M can be calculated.

**2 Calculate the molar mass of copper(II) sulfate.**

$$M(\text{CuSO}_4) = M(\text{Cu}) + M(\text{S}) + 4 \times M(\text{O}) = 63.55 + 32.06 + 4 \times 16.00 \\ = 159.6 \text{ g mol}^{-1}$$

3 Substitute values into the equation.

$$n(\text{CuSO}_4) = \frac{12.2}{159.6} \\ = 0.0764 \text{ mol}$$

There are 0.0764 mol in 12.2 g of copper(II) sulfate.

WORKED EXAMPLE 9.2.5

Calculate the mass of 0.75 mol of sodium hydroxide (NaOH).

ANSWER**1 Write the relationship.**

$$n = \frac{m}{M}$$

$$n(\text{NaOH}) = 0.75 \text{ mol}, m = ?, M \text{ can be calculated.}$$

2 Calculate the molar mass of sodium hydroxide.

$$M(\text{NaOH}) = 22.99 + 16.00 + 1.01 = 40.00 \text{ g mol}^{-1}$$

3 Rearrange to find m .

$$m = n \times M$$

4 Substitute and solve.

$$m(\text{NaOH}) = 0.75 \times 40.00 \\ = 30 \text{ g}$$

There are 30 g of sodium hydroxide in 0.75 mol.

**Worksheet**

Percentage composition

percentage composition

the percentage by mass of each of the different elements in the compound

Percentage composition

The chemical formula of a compound gives information about the elements present in a compound and the ratio in which the atoms of those elements are present. For example, in phosphoric acid (H_3PO_4), H, P and O are present in the ratio 3:1:4.

It is important to remember that the ratio of the atoms of the elements differs from the ratio of the masses of the element's atoms because different elements have different relative atomic masses. Sometimes, scientists need to know the ratio of elements by mass or the percentage composition.

The **percentage composition** of a compound is the percentage by mass of each of the different elements in the compound. For example, **Figure 9.2.5** shows aluminium foil, which



FIGURE 9.2.5 Aluminium, bauxite and aluminium oxide (alumina) have different percentage compositions of aluminium.

has approximately 98.5% of the element Al compared to bauxite, which is the mineral from which aluminium is derived, and aluminium oxide (alumina). Aluminium oxide (Al_2O_3) contains approximately 53% Al (as shown in Worked example 9.2.6). The formula to use is:

$$\text{Percentage composition (\%)} = \frac{\text{mass of element in compound}}{\text{molecular or formula mass of compound}} \times 100$$

Percentage composition provides useful information in many different instances. For example, a chemist can determine which mineral ores contain the highest percentage of a particular metal to advise mining and metal extraction industries. A chemist often compares the percentage composition of an unknown compound with the percentage composition calculated from a known formula. If the percentages match, the identity of the unknown can be confirmed.

WORKED EXAMPLE 9.2.6

Calculate the percentage composition of aluminium oxide (Al_2O_3).

ANSWER

1 Calculate the relative mass of each element in the compound.

$$\text{Mass of Al} = 2 \times 26.98 = 53.96$$

$$\text{Mass of O} = 3 \times 16.00 = 48.00$$

2 Calculate the relative formula (or molecular) mass.

$$\text{Formula mass of Al}_2\text{O}_3 = \text{mass of Al} + \text{mass of O}$$

$$= 53.96 + 48.00$$

$$= 101.96$$

3 Calculate the percentage of each element.

$$\% \text{Al} = \frac{\text{mass of Al in compound}}{\text{formula mass of Al}_2\text{O}_3}$$

$$= \frac{53.96}{101.96} \times 100$$

$$= 52.92\%$$

$$\% \text{O} = \frac{\text{mass of O in compound}}{\text{formula mass of Al}_2\text{O}_3}$$

$$= \frac{48.00}{101.96} \times 100$$

$$= 47.08\%$$

The percentage composition is 52.92% Al and 47.08% O.

Empirical and molecular formulas

As explained in Chapter 5, the empirical formula of any compound is the simplest whole number ratio in which the atoms of the elements are present in the compound. For an ionic compound, this is the formula that is given because ionic compounds have an infinite lattice structure. In contrast, a molecular substance consists of individual molecules, so its molecular formula states the actual number of each type of atom present in the molecule.

For example, the compound hydrogen peroxide has the molecular formula H_2O_2 . It is composed of two atoms of hydrogen and two atoms of oxygen bonded together. The empirical formula of hydrogen peroxide is HO because this is the simplest whole number ratio. However, for water (H_2O), the empirical and molecular formulas are the same because the ratio of atoms cannot be simplified any further.

Other examples of empirical and molecular formulas are shown in **Table 9.2.1**.

TABLE 9.2.1 Molecular and empirical formulas of different compounds

Substance	Molecular formula	Empirical formula
Methane	CH ₄	CH ₄
Octane	C ₈ H ₁₈	C ₄ H ₉
Glucose	C ₆ H ₁₂ O ₆	CH ₂ O
Calcium nitride	Not applicable (ionic)	Ca ₃ N ₂
Silicon dioxide	Not applicable (giant covalent network)	SiO ₂

It is possible for different compounds to have the same empirical formula.

Determining empirical and molecular formulas

The empirical formula of a substance can be determined experimentally. The term ‘empirical’ means obtained from experiment. This can be done by experimentally determining the percentage composition and then applying the mole concept to determine the ratio of particles. (Remember: the ratio of particles and moles of particles are the same numbers.)

WORKED EXAMPLE 9.2.7

An unidentified compound was analysed in the laboratory and found to have the following percentage composition by mass: 26.09% C, 4.35% H and 69.56% O.

- Determine its empirical formula.
- If the relative molecular mass of the compound is 92, determine its molecular formula.

ANSWERS

- a 1 Determine the mass of each element in a 100 g sample.**

Since the percentage composition is ‘per 100’, we can use this to assume that in 100 g of the compound:

Element	C	H	O
% (given)	26.09	4.35	69.56
Mass (in 100 g)	26.09	4.35	69.56

- 2 Find the number of moles of each element by dividing mass by the molar mass of that element.**

Element	C	H	O
Mass (in 100 g)	26.09	4.35	69.56
Moles (<i>n</i>)	$\frac{26.09}{12.01} = 2.172 = 2.17$	$\frac{4.35}{1.01} = 4.31$	$\frac{69.56}{16.00} = 4.347 = 4.35$

- 3 Find the simplest whole number ratio by dividing the number of moles by the smallest number then rounding to a whole number.

Element	C	H	O
Mass (in 100 g)	26.09	4.35	69.56
Moles (n)	$= \frac{26.09}{12.01} = 2.17$	$= \frac{4.35}{1.01} = 4.31$	$= \frac{69.56}{16.00} = 4.35$
Simplest ratio	$= \frac{2.17}{2.17} = 1$ $= 1$	$= \frac{4.31}{2.17} = 1.99$ $= 2$	$= \frac{4.35}{2.17} = 2.00$ $= 2$

Based on the calculation, the ratio shows that there is 1 : 2 : 2 relationship between carbon, hydrogen and oxygen. Therefore, the empirical formula is CH_2O_2 .

- b 1 Determine the relative empirical formula mass.

$$\begin{aligned}\text{Empirical formula mass} &= 1 \times 12.01 + 2 \times 1.01 + 2 \times 16.00 \\ &= 12.01 + 2.02 + 32.00 \\ &= 46.03\end{aligned}$$

- 2 To calculate the molecular formula, determine how many empirical units there are in the molecule:

$$\begin{aligned}\frac{\text{Relative molecular mass}}{\text{Relative empirical mass}} &= \frac{92}{46.03} \\ &= 2\end{aligned}$$

This means that molecular formula contains two empirical formula units.

- 3 Determine the molecular formula.

$$\begin{aligned}\text{Molecular formula} &= 2 \times \text{CH}_2\text{O}_2 \\ &= \text{C}_2\text{H}_4\text{O}_4\end{aligned}$$

The molecular formula is $\text{C}_2\text{H}_4\text{O}_4$.

As shown in Worked example 9.2.7, part b, the molecular formula is a multiple of the empirical formula. If the molecular mass is known, we can calculate the molecular formula of a compound based on its empirical formula.

WORKED EXAMPLE 9.2.8

A compound contains 57.54% carbon, 3.45% hydrogen and the remaining is fluorine. Determine the empirical formula of this compound.

ANSWER

- 1 Determine the mass of each element in a 100 g sample.

Since percentage must add to 100, determine the percentage of F.

Element	C	H	F
% (given)	57.54	3.45	$(100 - (57.54 + 3.45)) = 39.01$
Mass (in 100 g)	57.54	3.45	39.01

2 Find the number of moles of each element by dividing the mass by the molar mass.

Element	C	H	F
Mass (in 100 g)	57.54	3.45	39.01
Moles (n)	$= \frac{57.54}{12.01} = 4.791 = 4.79$	$= \frac{3.45}{1.01} = 3.415 = 3.42$	$= \frac{39.01}{19.00} = 2.053 = 2.05$

3 Find the simplest whole number ratio by dividing the number of moles by the smallest number, then rounding to a whole number. Where necessary, multiply by the lowest multiple to achieve a whole number ratio.

Element	C	H	F
Mass (in 100 g)	57.54	3.45	39.01
Moles (n)	$= \frac{57.54}{12.01} = 4.97$	$= \frac{3.45}{1.01} = 3.42$	$= \frac{39.01}{19.00} = 2.05$
Simplest ratio	$= \frac{4.79}{2.05} = 2.34$ $= 2.34 \times 3 = 7$	$= \frac{3.42}{2.05} = 1.67$ $= 1.67 \times 3 = 5$	$= \frac{2.05}{2.05} = 1$ $= 1 \times 3 = 3$

Based on the calculation, the ratio shows that there is 7 : 5 : 3 relationship between carbon, hydrogen and fluorine. Therefore the empirical formula is $C_7H_5F_3$.

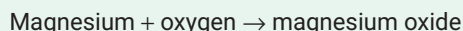
As seen in Worked example 9.2.8, sometimes dividing by the smallest number does not give a ratio close to a whole number. For example, if the ratio is 2.34 to 1, multiplying by 3 gives a ratio of 7 to 3. Other common ratios are 1.33 to 1, or 4 to 3, and 1.67 to 1, or 5 to 3. When these non-integer ratios occur, another step is needed in the calculation, i.e. multiplication by an integer that converts the ratio to a whole number (integer) ratio.

PRACTICAL ACTIVITY 9.2.1

DETERMINING THE EMPIRICAL FORMULA OF MAGNESIUM OXIDE

Introduction

The empirical formula of a substance can be calculated from experimental data. In this experiment, you will carry out the reaction:



By determining the masses of each element and using the relationship:

$$\text{Number of moles} = \frac{\text{mass}}{\text{molar mass}}$$

the number of moles of each element can be determined. The empirical formula is derived from this.

Research question

Can the empirical formula of magnesium oxide be accurately determined experimentally?

Aim

Write an aim for this experiment.

Equipment per group

- 1 strip of magnesium ribbon (about 20 cm)
- steel wool
- crucible tongs
- pipe clay triangle
- Bunsen burner
- tripod
- crucible and lid
- matches
- electronic balance



What are the risks in doing this experiment?

Burning magnesium emits a bright light.

Magnesium oxide is an irritant to eyes, nose and throat.

How can you manage these risks to stay safe?

Do not look directly at burning magnesium.

Lift only the edge of the crucible lid. Work in a well-ventilated area.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

- 1 Place the crucible with lid on a pipe clay triangle over a Bunsen burner and heat it strongly over a hot flame for 5 minutes. Allow it to cool.
- 2 Weigh the crucible and lid.
- 3 Thoroughly clean the surface of the magnesium ribbon with steel wool.
- 4 Record the appearance of the cleaned magnesium.
- 5 Coil the piece of magnesium ribbon so that it fits inside the crucible. Weigh the crucible, lid and magnesium ribbon.
- 6 Place the crucible back on the pipe clay triangle on a tripod over the Bunsen burner and carefully heat the crucible without the lid until the magnesium begins to glow. Use tongs to then put the lid on before the magnesium bursts into flames.

Warning: Do not look directly at the burning magnesium.

- 7 Once the lid is in place, heat the crucible strongly for about 10 minutes, lifting the lid occasionally to admit oxygen. Try to prevent any magnesium oxide smoke escaping.
- 8 Remove the lid and heat the crucible for a further 5 minutes to ensure complete reaction.
- 9 Replace the lid and allow the reaction mixture to cool. Reweigh the crucible, lid and magnesium oxide.

Results

Record your results in a results table similar to the one below.

Items	Mass (g)
Crucible and lid	
Crucible, lid and magnesium	
Magnesium	
Crucible, lid and magnesium oxide	
Magnesium oxide	
Oxygen	

Analysis of results

- 1 Calculate the mass of magnesium and magnesium oxide.
- 2 Calculate the mass of oxygen.
- 3 Calculate the number of moles of magnesium and oxygen.
- 4 State the ratio of moles of magnesium to moles of oxygen.
- 5 Using the mole ratio, determine the empirical formula of magnesium oxide.

Interpretation

- 6 Compare your results with those of other groups in the class and suggest possible reasons for any variations.

Evaluation

- 7 Consider the sources of error in this experiment and suggest how the accuracy of the results might be improved.

LEARNING CHECK 9.2

DESCRIBING

- 1 **Distinguish** between relative molecular mass and molar mass.
- 2 **Distinguish** between an empirical formula and a molecular formula.
- 3 **Define** 'percentage composition'.

APPLYING

- 4 **Calculate** the relative molecular mass of each of the following molecules. Ensure that you include the correct units.
 - a Phosphorus (P_4)
 - b Carbon dioxide (CO_2)
 - c Nitric acid (HNO_3)
 - d Glucose ($C_6H_{12}O_6$).
- 5 **Calculate** the relative formula mass of:
 - a aluminium nitride (AlN)
 - b barium nitrate ($Ba(NO_3)_2$)
 - c sodium carbonate (Na_2CO_3)
 - d copper(II) sulfate ($CuSO_4$).
- 6 **Calculate** the molar mass of:
 - a hydrogen peroxide (H_2O_2)
 - b magnesium bromide ($MgBr_2$)
 - c calcium carbonate ($CaCO_3$)
 - d pentane (C_5H_{12}).
- 7 **Determine** the mass of:
 - a 7.10×10^{23} atoms of neon gas
 - b 5.43 mol of lead
 - c 3.14×10^{24} molecules of phosphorus tetramer (P_4)
 - d 1.236 mol of magnesium sulfate ($MgSO_4$)
 - e 0.452 mol of potassium phosphate (K_3PO_4).
- 8 **Calculate** the number of moles in 10.00 g of:
 - a water
 - b ethanoic (acetic) acid (CH_3COOH)
 - c ethane (C_2H_6)
 - d barium chloride
 - e sodium bromide.
- 9 **Calculate** the percentage composition of each element in:
 - a methane (CH_4)
 - b sodium chloride ($NaCl$)
 - c calcium cyanide ($Ca(CN)_2$)
 - d ammonium oxalate ($(NH_4)_2C_2O_4$).

10 **Determine** the empirical formula of:

- a COCl_2
- b C_4H_{10}
- c N_2O_4
- d $\text{Na}_4\text{P}_2\text{O}_6$

11 An oxide of chlorine contains 18.4% oxygen by mass. **State** the empirical formula of the compound.

12 a **Determine** the empirical formula of the following compounds, given their percentage composition by mass:

- i 7.8% C, 92.2% Cl
- ii 64.8% C, 13.6% H, 21.6% O
- iii 32.0% C, 6.7% H, 42.6% O, 18.7% N.

b **Identify** the molecular formula of the compound in part a ii if its relative molecular mass is 74.14.

13 Nicotine was analysed to determine its composition. It was found to contain the following percentages by mass: 74.0% C, 8.7% H and 17.3% N. Its relative molar mass was found to be 162. **State** the:

- a empirical formula
- b molecular formula.

14 The artificial sweetener NutraSweet® is the compound aspartamine, which has the molecular formula $\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5$.

- a **Calculate** the mass of 1.00 mol of aspartamine.
- b **Determine** how many moles of aspartamine are in 6.22 g of the substance.
- c **Calculate** the mass of 0.245 mol of aspartamine.
- d **Calculate** the number of molecules in 4.28 mg of aspartamine.

15 a **Determine** the number of molecules of oxygen gas in 4.00 g of O_2 .

b **Determine** how many atoms of oxygen would be obtained if the molecules in part a were split in half (i.e. into O atoms).

16 a If the body of a person is 80% water, **calculate** how many moles of water would be in the body of a person who weighs 80 kg.

b **Determine** how many water molecules this is.

9.3 The law of conservation of mass in chemical equations



Weblink
The law of conservation of mass

When a chemical reaction occurs, new substances are formed. This can be represented by using chemical equations. Balanced chemical equations represent the relationship between the number of particles that react and the number of particles that are produced.

Chemical equations must be balanced because atoms are conserved. This is a consequence of the law of conservation of mass: mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or physical transformations. The law of conservation of mass was developed by Antoine Lavoisier (Figure 9.3.1).

KEY LAW

The law of conservation of mass

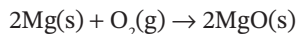
Mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or physical transformations.



FIGURE 9.3.1 Marie and Antoine Lavoisier

Moles and chemical reactions

The reaction between magnesium (Mg) and oxygen (O₂) can be represented by the balanced equation:

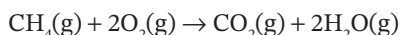


The coefficient of each species in the equation indicates the ratio in which the species react. In this case, two atoms of magnesium react with one molecule of oxygen gas to produce two formula units of magnesium oxide (MgO).

The chemical equation can also be read in terms of moles, because the number of particles is directly proportional to the number of moles through the Avogadro constant. This means the coefficients of the species in the equation also indicate the molar ratios for the species taking part in the reaction; that is, 2 mol of magnesium react with 1 mol of oxygen gas to produce 2 mol of magnesium oxide.

The study of the amounts of reactants and products in a chemical reaction is called **stoichiometry**. This word comes from the Greek words *stoicheion*, meaning 'element', and *metron*, meaning 'measurement'.

Consider the reaction for the complete combustion of methane gas (CH₄) to form carbon dioxide (CO₂) and water:



The relationships between the numbers of moles of reactants and products are represented in **Table 9.3.1**. Although the mole quantities (numbers of moles) are different, the ratios remain the same. All are in the ratio of 1:2:1:2.

TABLE 9.3.1 The ratio of particles involved when methane reacts with oxygen gas

CH ₄	+	2O ₂ (g)	→	CO ₂ (g)	+	2H ₂ O(g)
1 mol		2 mol		1 mol		2 mol
2 mol		4 mol		2 mol		4 mol
0.5 mol		1 mol		0.5 mol		1 mol
5 mol		10 mol		5 mol		10 mol

If the number of moles of any species in the reaction is known, apply the balanced equation to work out the number of moles of all the other species in that reaction. From Table 9.3.1, you can see that calculating the number of moles of a substance has to do with the molar ratio in the equation. For example, if 2 mol of methane are used in a reaction, we would need 4 mol of oxygen for this reaction to proceed. This is because the ratio of methane to oxygen is 1:2 according to the coefficients in the balanced equation. To complete the product calculations, we would look at the coefficients and determine the ratio. The ratio of carbon dioxide to water is 1:2; therefore, 2 mol of carbon dioxide and 4 mol of water will be produced in this reaction. We will cover stoichiometry and its relationship to mole in greater detail later in section 9.4.

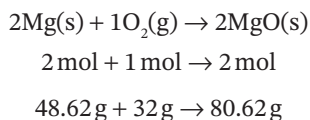
Mass and chemical equations

Chemical equations indicate a relationship between numbers of moles of reactants and products, and masses are related to moles. This means that chemical equations can also be used to determine relationships between masses of reactants and products.

Remember that chemical reactions obey the law of conservation of mass. The direct consequence of this law is that, in chemical reactions, the combined mass of the products must equal the combined mass of the reactants.

stoichiometry the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers

For example, in the reaction between magnesium and oxygen, the mass–mass relationships between reactants and products are:



Notice that, although atoms and mass are conserved, moles are not. This is because moles count the number of molecules and formula units, not necessarily individual atoms. This means that 1 mol of a substance may contain many moles of individual atoms.

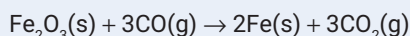
LEARNING CHECK 9.3

DESCRIBING

- 1 **Define** 'stoichiometry'.
- 2 **State** the law of conservation of mass.
- 3 **Explain** why reactants and products always need to be related by a balanced chemical equation.

APPLYING

- 4 Copy and complete the following table to **determine** the missing numbers of moles using the equation:



Fe_2O_3	CO	Fe	CO_2
1 mol			
		3 mol	
	0.6 mol		
			0.09 mol

9.4 Stoichiometry

Stoichiometry is the study of the amounts of reactants and products in a chemical reaction. As stated in section 9.3, stoichiometry is the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole numbers. This is referred to as the **stoichiometric ratio**. The rules followed when determining stoichiometric relationships are based on the law of conservation of mass. A balanced equation makes it possible to convert information about the change in one reactant or product to quantitative data about another reactant or product. Understanding this is essential to solving stoichiometric problems.

stoichiometric ratio
the exact proportion of reactants and products in a chemical reaction, based on the balanced chemical equation

Mole–mole calculations

Mole-to-mole calculations use the number of moles of a known substance and a balanced chemical equation to calculate the number of moles of all other species (reactants and products) involved in the reaction.

When solving problems involving mole–mole stoichiometry, it is important to remember that the mole relationships between reactants and products are derived from the mole ratios in a balanced chemical equation. Therefore, a useful strategy to calculate the number of mol of a required substance is:

1. Determine the number of mol of a given substance.
2. Use a balanced chemical equation to determine the ratios.
3. Determine the number of mol of required substance.

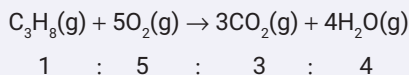
WORKED EXAMPLE 9.4.1

Propane gas (C_3H_8) burns in oxygen to produce carbon dioxide and water. If 2.5 mol of propane is burnt, calculate the number of moles of:

- a oxygen gas needed for complete combustion
- b carbon dioxide produced.

ANSWERS

Write a balanced equation and determine the reaction ratio from the coefficients:



- a 1 Establish the mole ratio between the given substance and the required substance.**

$$n(\text{C}_3\text{H}_8) : n(\text{O}_2)$$

$$1 : 5$$

- 2 Use the ratio to calculate the number of moles from the reacted amount given.**

$$n(\text{required}) = \frac{\text{ratio of number of mol of substance required}}{\text{ratio of number of mol of given substance (from balanced equation)}} \times \text{number of mol of given substance}$$

Since the ratio of propane to oxygen is 1 : 5:

$$n(\text{O}_2) = \frac{5}{1} \times n(\text{C}_3\text{H}_8)$$

$$= \frac{5}{1} \times 2.5 = 12.5 \text{ mol of O}_2$$

12.5 mol of O_2 are needed.

- b 1 Establish the mole ratio between the given substance and the required substance.**

$$n(\text{C}_3\text{H}_8) : n(\text{CO}_2)$$

$$1 : 3$$

- 2 Use the ratio to calculate the number of moles from the reacted amount given.**

$$n(\text{required}) = \frac{\text{ratio of number of mol of substance required}}{\text{ratio of number of mol of given substance (from balanced equation)}} \times \text{number of mol of given substance}$$

Since the ratio of propane to carbon dioxide is 1 : 3:

$$n(\text{CO}_2) = \frac{3}{1} \times n(\text{C}_3\text{H}_8)$$

$$= \frac{3}{1} \times 2.5 = 7.5 \text{ mol of CO}_2$$

7.5 mol of CO_2 are produced.

As shown in Worked example 9.4.1, a useful strategy that we can use to determine the number of mol of a required substance is:

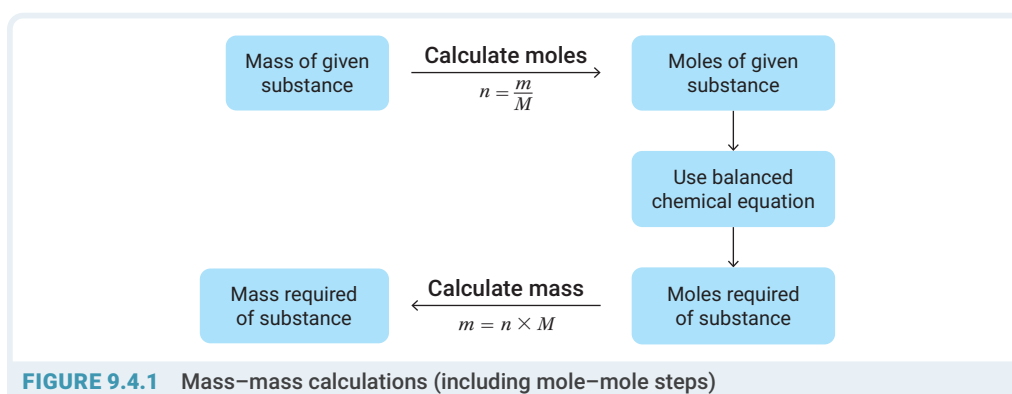
$$n(\text{required}) = \frac{\text{ratio of number of mol of substance required}}{\text{ratio of number of mol of given substance (from balanced equation)}} \times \text{number of mol of given substance}$$

In order to use this strategy, we need to know the value of at least one substance as well as a balanced equation.

Mass–mass calculations

Mass–mass calculations use the mass of a known substance and a balanced chemical equation to calculate the mass of all other species (reactants and products) involved in the reaction.

Figure 9.4.1 shows the steps involved in mass–mass calculations.

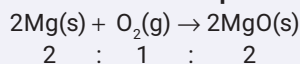


WORKED EXAMPLE 9.4.2

Solid magnesium burns in the presence of oxygen gas to form magnesium oxide. Calculate the mass of oxygen that would be needed to produce 24.0 g of magnesium oxide.

ANSWER

- 1 Write a balanced equation and determine the reaction ratio from the coefficients.



- 2 Calculate the number of moles of the known substance.

$$n(\text{MgO}) = \frac{\text{mass}}{\text{molar mass}} = \frac{24.0}{40.31} = 0.595 \text{ mol MgO}$$

- 3 Use the ratio to calculate the number of moles of the unknown substance.

1 mol of oxygen produces 2 mol magnesium oxide.

$$n(\text{required}) = \frac{\text{ratio of number of mol of substance required}}{\text{ratio of number of mol of given substance}} \times \text{number of mol of given substance}$$

$$n(\text{O}_2) = \frac{1}{2} \times 0.595 = 0.298 \text{ mol}$$

0.298 mol of oxygen are needed.

- 4 Calculate the mass of the unknown substance.

$$\begin{aligned} m(\text{O}_2) &= n \times M \\ &= 0.298 \times 32.00 = 9.54 \text{ g} \end{aligned}$$

9.54 g of oxygen are needed.

Limiting reactants

Applying the relationships between masses of products and reactants is very important in chemistry. It is vital when calculating how much of a desired product is made and how much by-product (waste) is also made. This is important in all chemical manufacturing, including the plastic, metal, pharmaceutical, fertiliser and petroleum industries.

When one or more chemicals react to give one or more product molecules, it is possible that not all of the reactants are used up. This means that there is a **limiting reactant** and an **excess reactant(s)**. The limiting reactant dictates how much product we can theoretically produce in the reaction, also known as the **theoretical yield**.

limiting reactant the reactant that is totally consumed when a chemical reaction is complete

excess reactant the reactant that is not totally consumed when a chemical reaction is complete

theoretical yield the amount (mass) of a product that is produced from a complete reaction of the limiting reactant

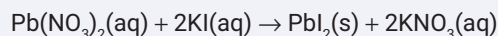
WORKED EXAMPLE 9.4.3

A solution containing 0.600 g of lead(II) nitrate ($\text{Pb}(\text{NO}_3)_2$) is added to a solution containing 2.60 g of potassium iodide (KI). From this reaction, lead iodide (PbI_2) and potassium nitrate (KNO_3) are formed.

- Determine the limiting and excess reactants.
- Calculate the mass of lead(II) iodide formed.

ANSWERS

- a 1 Write the balanced equation.**



- 2 Calculate the number of moles of each reactant.**

$$\begin{aligned} n(\text{Pb}(\text{NO}_3)_2) &= \frac{m}{M} \\ &= \frac{0.600}{331.22} = 1.81 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{KI}) &= \frac{m}{M} \\ &= \frac{2.60}{166.0} = 0.0157 \text{ mol} \end{aligned}$$

- 3 Determine the limiting reactant.**

Option 1

Use the ratio to calculate the theoretical number of moles from the reacted amounts given to find the limiting reactant.

Suppose we want to use all of the available KI. The amount of $\text{Pb}(\text{NO}_3)_2$ required would be:

$$\begin{aligned} n(\text{Pb}(\text{NO}_3)_2) &= \frac{1}{2} \times n(\text{KI}) \\ &= \frac{1}{2} \times 0.0157 = 7.85 \times 10^{-3} \text{ mol} \end{aligned}$$

7.85×10^{-3} mol of $\text{Pb}(\text{NO}_3)_2$ is required to react with 0.0157 mol of KI.

The theoretical value of 7.85×10^{-3} mol of $\text{Pb}(\text{NO}_3)_2$ is more than the number of moles of $\text{Pb}(\text{NO}_3)_2$ available in the experiment (1.81×10^{-3} mol). Therefore, the limiting reactant is lead(II) nitrate.

Option 2

Instead of using $\text{Pb}(\text{NO}_3)_2$, you could also calculate this using the number of moles of KI.

Suppose we want to use all the available $\text{Pb}(\text{NO}_3)_2$. The amount of KI required would be:

$$\begin{aligned}n(\text{KI}) &= \frac{2}{1} \times n(\text{Pb}(\text{NO}_3)_2) \\&= \frac{2}{1} \times 1.81 \times 10^{-3} = 3.62 \times 10^{-3} \text{ mol}\end{aligned}$$

3.62×10^{-3} mol of KI is required to react with 1.81×10^{-3} mol of $\text{Pb}(\text{NO}_3)_2$.

The theoretical value of 3.62×10^{-3} mol of KI is less than the number of moles of KI available in the experiment (0.0157 mol). Therefore, the excess reactant is potassium iodide.

b 1 Determine the ratio between the limiting reactant and the required substance.

$$\begin{aligned}n\text{Pb}(\text{NO}_3)_2 : n(\text{PbI}_2) \\1 : 1\end{aligned}$$

2 Use the stoichiometric ratio to calculate the theoretical number of moles from the reacted amount of limiting reactant to find the moles of the product formed.

$$\begin{aligned}n(\text{PbI}_2) &= \frac{1}{1} \times n(\text{Pb}(\text{NO}_3)_2) \\&= \frac{1}{1} \times 1.81 \times 10^{-3} = 1.81 \times 10^{-3} \text{ mol of PbI}_2 \text{ is formed.}\end{aligned}$$

3 Convert moles to mass.

$$\begin{aligned}m(\text{PbI}_2) &= n \times M \\&= 1.81 \times 10^{-3} \times 461 = 0.834 \text{ g}\end{aligned}$$

0.834 g of the precipitate, lead(II) iodide is formed.



Worksheet
Limiting reactants

As seen in Worked example 9.4.3, we can use any of the substances given to determine the limiting reactant. Either we identify the limiting reactant or we identify the excess reactant – both will give us the information we require. It is important to remember that we only use the limiting reactant to calculate the quantity of all other substances in the chemical reaction.

In instances where there is the stoichiometric ratio of reactants, then each of the reactants will be present in just sufficient quantity to be completely used up at the end of the reaction.

Experimental and theoretical yield

So far, all of our calculations have helped us to determine the theoretical or ideal yield of a reaction. However, the theoretical yield assumes perfect completion of the reaction. In an experimental situation, the actual **experimental yield** may be less than the theoretical yield. There are a number of possible reasons for this:

1. There may be competing reactions that have not been considered in the written, balanced equation.
2. The mechanism of the reaction may be different because of experimental conditions differing from what is required for the assumed reaction.
3. The conditions required for the reaction to take place, such as temperature and pressure, may not be met.
4. The reaction mixture may contain impurities that influence the reaction.

Generally, experimentally obtained yields are less than the calculated theoretical yields. The ratio of the actual (experimental) yield to the theoretical yield is a quantity known as the 'fractional yield'. This is usually quoted as a percentage of the theoretical yield and is defined as the percentage yield.

experimental yield the actual yield (mass) of a product obtained for a reaction conducted in the laboratory

KEY FORMULA

$$\text{Percentage yield (\%)} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

Calculating yields

Luckily we have spent most of this chapter practising the skills that are necessary to calculate the yield. That means we now know how to use a balanced equation to calculate the amount of any substance produced.

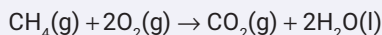
WORKED EXAMPLE 9.4.4

In a laboratory, scientists collected the water formed from the combustion of 32.0 g of methane gas (CH_4) and oxygen gas (O_2) using a condensation apparatus, and then weighed the water. A total of 60.0 g of water was collected.

Determine the percentage yield of water resulting from the combustion of 32.0 g of methane (assume the oxygen required for the reaction is in excess).

ANSWER

- 1 Write a balanced equation for this reaction.



- 2 Determine the amount of methane reacted.

$$\begin{aligned} n(\text{CH}_4) &= \frac{m}{M} \\ &= \frac{32.0 \text{ g}}{16.05 \text{ g mol}^{-1}} = 1.99 \text{ mol} \end{aligned}$$

- 3 Use the balanced equation to determine the mole ratio.

$$\begin{array}{l} n(\text{CH}_4) : n(\text{H}_2\text{O}) \\ 1 : 2 \end{array}$$

- 4 Calculate the theoretical yield of water.

$$\begin{aligned} n(\text{H}_2\text{O}) &= \frac{2}{1} \times n(\text{CH}_4) \\ &= \frac{2}{1} \times 1.99 = 3.98 \text{ mol} \end{aligned}$$

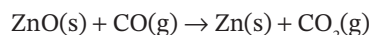
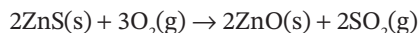
$$m(\text{H}_2\text{O}) = n(\text{H}_2\text{O}) \times M(\text{H}_2\text{O}) = 3.98 \times 18.02 \text{ g mol}^{-1} = 71.7 \text{ g}$$

- 5 Calculate the percentage yield of water for this experiment.

$$\begin{aligned} \text{Percentage yield (\%)} &= \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1} \\ &= \frac{60.0}{71.7} \times \frac{100}{1} = 83.7\% \end{aligned}$$

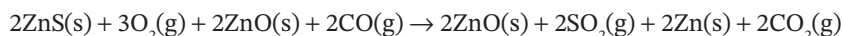
The percentage yield is 83.7%.

In some reactions, two steps may take place overall. For example, pure zinc metal can be produced in a two-step process that starts with zinc sulfide (ZnS). ZnS is the main constituent of the mineral sphalerite. The mineral contains impurities. The two-step reaction sequence, starting with ZnS , is:



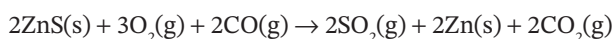
To calculate the theoretical yield, you need to calculate the number of moles of zinc produced in the two-step reaction. The strategy is to obtain an overall reaction, which ‘sums’ the individual steps and excludes any ‘intermediates’. Although zinc oxide (ZnO) is produced in the first step, it

is then consumed in the second step. Zinc oxide is an intermediate and does not appear among the final products for the overall reaction. To exclude an intermediate from the overall reaction, equations for the individual steps must be added in a specific way so that the intermediates cancel. The coefficients of intermediates need to be equalised, just as you would solve simultaneous equations in algebra. The coefficient of ZnO is '2' in the first step, but just '1' in the second step. If the second step is multiplied by 2 and added to the first step, this equation is obtained:



As in an algebraic equation, the 2ZnO appearing in the same state on both sides of the equation can be cancelled.

The overall reaction is shown in the following equation, which indicates that, in theory, 1 mol of zinc can be obtained for every mole of ZnS:



For example, beginning with 1 mol of ZnS (97.44 g), the theoretical yield would be 1 mol of Zn; that is, 65.38 g. This is the basis for a calculation of experimental yield, as shown in Worked example 9.4.5.

WORKED EXAMPLE 9.4.5

If 50.0 kg of ZnS yielded 29.5 kg of pure zinc (Zn), calculate the experimental yield for the process.

ANSWER

1 Refer to the overall reaction for the process to calculate the theoretical yield.

2 mol of ZnS yields 2 mol of Zn in the overall reaction.

Therefore, 1 mol of ZnS would produce 1 mol of Zn metal.

2 Calculate the molar mass of ZnS.

The atomic weight of Zn is 65.38. The atomic weight of S is 32.06.

The molar mass of ZnS = 65.38 + 32.06 = 97.44 g mol⁻¹

3 Convert mass of ZnS to moles of ZnS.

$$50 \text{ kg of ZnS contains } = \frac{5 \times 10^4}{97.44} = 513 \text{ mol}$$

The number of moles of ZnS used is the same as the moles of Zn produced.

$$\begin{aligned} m(\text{Zn}) &= 513 \times 65.38 \text{ g} \\ &= 33\,540 \text{ g} = 33.5 \text{ kg} \end{aligned}$$

The theoretical yield for the two-step reaction of 50.0 kg of ZnS is 33.5 kg of Zn.

4 Calculate actual (experimental) percentage yield:

$$\text{Experimental \% yield} = \frac{29.5}{33.5} \times \frac{100}{1} = 88.1\%$$

The experimental yield is 88.1%.

In industrial chemistry, including pharmaceutical chemistry, multistep reactions are often used to produce end products. The overall percentage yield is the product of the percentage yields for each step, so each step should have as high a percentage yield as possible.

For example, each step in a five-step process only has 50% yield. The fractional yield for each step is 0.5. Multiplying 0.5 by itself five times gives $(0.5)^5 = 0.031$ = overall fractional yield. This corresponds to an overall 3.1% percentage yield. Compare this with the case if each step had a 95% yield. An overall fractional yield of $(0.95)^5 = 0.77$ would be obtained, which corresponds to an overall 77% yield. The latter is obviously a preferable outcome!

Some key things to remember are:

- the mass relationships between reactants and products are derived from the stoichiometric ratios in a balanced chemical equation
- all but one of the reactants may be in excess
- always identify the limiting reactant if other reactants are in excess.
- Alternatively, if the quantities of reactants are present in exactly the right stoichiometric ratio, then each of the reactants will be available in precisely sufficient quantity for them to all be used up completely at the end of the reaction.

Worked example 9.4.6 illustrates how you would combine limiting reactants with percentage yield.

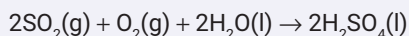
WORKED EXAMPLE 9.4.6

Sulfuric acid (H_2SO_4) is formed when sulfur dioxide (SO_2) and water react with oxygen.

If 3.20 g of SO_2 and 1.80 g of water were allowed to react in the presence of excess oxygen gas, determine the percentage yield if 4.0 g of sulfuric acid was produced from the experiment.

ANSWER

1 Write a balanced equation.



2 Determine the limiting reactant.

We first need to determine the amount of available reactants.

$$n(\text{SO}_2) = \frac{3.20}{64.06} = 0.04995 \text{ mol} = 0.0500 \text{ mol}$$

$$n(\text{H}_2\text{O}) = \frac{1.80}{18.02} = 0.09989 \text{ mol} = 0.100 \text{ mol}$$

Suppose we want to use all of the available SO_2 . The amount of water required is:

$$n(\text{SO}_2) : n(\text{H}_2\text{O})$$

$$2 : 2$$

$$n(\text{H}_2\text{O}) = \frac{2}{2} \times 0.0500 = 0.0500 \text{ mol}$$

The theoretical value of 0.0500 mol of H_2O is less than the number of moles of H_2O available in the experiment (0.100 mol). Therefore, the excess reactant is H_2O and the limiting reactant is SO_2 .

3 Use the limiting reactant to calculate the theoretical yield of sulfuric acid.

$$n(\text{SO}_2) : n(\text{H}_2\text{SO}_4)$$

$$2 : 2$$

$$n(\text{H}_2\text{SO}_4) = \frac{2}{2} \times 0.0500 = 0.0500 \text{ mol of } \text{H}_2\text{SO}_4$$

$$m = 0.0500 \times 98.08 = 4.91 \text{ g}$$

4 Determine the percentage yield.

$$\% \text{ yield} = \frac{4.0}{4.91} \times \frac{100}{1} = 81.5\%$$

The experimental percentage yield is 81.5%.

PRACTICAL ACTIVITY 9.4.1

INVESTIGATING LIMITING REACTANTS

Introduction

Limiting reactants dictate the amount of product that can be produced from a chemical reaction. In industry, it is important to know which of the reactants is limiting because it affects the quantity of the desired product that is produced.

Research question

Can the limiting reactant be determined experimentally and used to accurately determine percentage yield?

Aims

- 1 To prepare two solutions and use them to perform a precipitate reaction
- 2 To determine which reactant limits the reaction
- 3 To calculate the percentage yield of this reaction

Equipment per group

- 1.00 g calcium nitrate ($\text{Ca}(\text{NO}_3)_2$), solid
- 0.50 g sodium carbonate (Na_2CO_3), solid
- 2×100 mL beakers
- 10 mL measuring cylinder
- 100 mL conical flask
- 12.5 cm filter paper
- deionised water
- stirring rod
- filter funnel
- pencil
- electronic balance



What are the risks in doing this experiment?	How can you manage these risks to stay safe?

Copy and complete the risk assessment table in your write-up. Add any risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

- 1 Weigh and record the mass of a dry piece of filter paper.
- 2 Prepare two solutions – one of calcium nitrate and one of sodium carbonate.
 - a Measure 0.50 g Na_2CO_3 in a clean, dry 100 mL beaker. Add exactly 10.00 mL of deionised water to the beaker using a 10.00 mL measuring cylinder. Swirl to mix until all the solid has dissolved.
 - b Measure 1.00 g $\text{Ca}(\text{NO}_3)_2$ in a clean, dry 100 mL beaker. Add exactly 10.00 mL of deionised water to the beaker. Swirl to mix until all the solid has dissolved.
- 3 Pour the solution from one beaker into the other and stir the mixture. Record all observations.
- 4 Place a filter funnel in the top of a 100 mL conical flask.
- 5 Fold the 12.5 cm filter paper as instructed, and write your initials in pencil on the outside. Place the folded filter paper in the filter funnel and wet the sides with deionised water.
- 6 Filter the mixture in the beaker through the filter paper.
- 7 Continue to refilter the solution through the funnel until the liquid in the solution becomes clear. This may require you to rinse the reaction beaker with deionised water several times. Filtration may require more than 10 minutes to complete.
- 8 Remove the filter paper from the filter funnel and place it where instructed by your teacher. Leave it to dry overnight.
- 9 Pour the filtrate down the sink. Rinse all glassware used and place in rinsed equipment container.

Next day

Record the mass of the filter paper and solid. Discard the filter paper/solid as instructed.

Results

a	Mass of Na_2CO_3 (g)	
b	Mass of $\text{Ca}(\text{NO}_3)_2$ (g)	
c	Mass of CaCO_3 precipitate and filter paper (g)	
d	Mass of filter paper (g)	
e	Calculated mass of CaCO_3 precipitate (c – d) (g)	

Analysis of results

- 1 Write a balanced chemical equation for the reaction including states.
- 2 Calculate the number of moles of:
 - a sodium carbonate
 - b calcium nitrate
 - c calcium carbonate.
- 3 Determine the limiting reactant and the excess reactant.
- 4 Calculate the theoretical yield.
- 5 Determine the percentage yield of calcium carbonate.

Interpretation

- 6 Identify the type of reaction that occurred during this experiment.
- 7 If the mass of Na_2CO_3 was doubled and the mass of CaCO_3 was held constant, would you have the same limiting reactant? Show calculations to support your answer.
- 8 Determine the mole ratio of the limiting reactant to the precipitate based on your experimental data. How does it compare to the ratio given in the balanced chemical equation?

Evaluation

- 9 Calculate the percentage error.
- 10 State the value of your percentage yield and discuss possible reasons (at least three) it may be different from the expected (theoretical) yield.

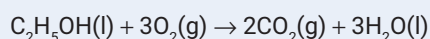
LEARNING CHECK 9.4

DESCRIBING

- 1 **Define** 'limiting reactant'.
- 2 **Describe** the theoretical yield of a reaction.
- 3 **Explain** why the experimental yield for a reaction is usually less than the theoretical yield.

APPLYING

- 4 The combustion of ethanol yields carbon dioxide and water:

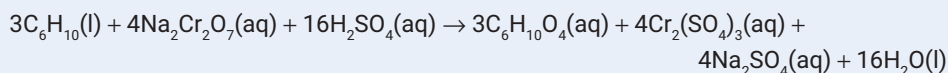


If the reaction is carried out in air:

- a **identify** which reactant is in excess
 - b **identify** the limiting reactant
 - c **determine** the stoichiometric ratio between the reactant ethanol and the product carbon dioxide.
- 5 The combustion of butane (C_4H_{10}) in air yields CO_2 and water.
- a Write a balanced chemical equation for this reaction.
 - b **Identify** the limiting reactant.
 - c **Calculate** the number of moles of O_2 needed for each mole of butane reacted.
 - d **Calculate** the number of grams of CO_2 produced when 6 g of butane undergoes complete combustion in air.
 - e **Calculate** the number of grams of H_2O produced when 6 g of butane undergoes complete combustion in air.
 - f **Determine** the percentage yield if 5.5 g of water is collected in a laboratory experiment from the combustion of 6 g of butane.

ANALYSING

- 6 Dimethylhydrazine ($(\text{CH}_3)_2\text{NNH}_2$) was used as a fuel for the Apollo Lunar Module, with N_2O_4 being used to replace oxygen during combustion. The products of the reaction are H_2O , N_2 and CO_2 .
- a Write a balanced chemical equation for the combustion reaction.
 - b If 150 kg of $(\text{CH}_3)_2\text{NNH}_2$ react with 460 kg of N_2O_4 , **determine** the theoretical yield of N_2 .
- 7 Adipic acid ($\text{C}_6\text{H}_{10}\text{O}_4$) is a raw material for making nylon and it can be prepared in the laboratory by the following reaction between cyclohexene (C_6H_{10}) and sodium dichromate ($\text{Na}_2\text{Cr}_2\text{O}_7$) in sulfuric acid:



Side reactions occur. These, plus losses of product during its purification, reduce the overall yield. A typical yield of purified adipic acid is 68.6%.

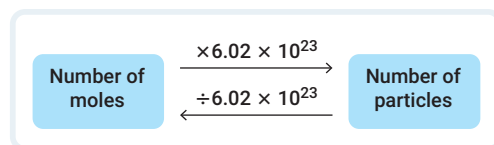
- a **Identify** how many grams of cyclohexene are required to prepare 12.5 g of adipic acid in 68.6% yield.
 - b The only available supply of sodium dichromate is its dihydrate ($\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$). (The reaction occurs in an aqueous medium, so the water in the dihydrate does not cause a problem, but it does mean that you will need a larger mass of the dihydrate form of sodium dichromate ($\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$) than if you were able to obtain just $\text{Na}_2\text{Cr}_2\text{O}_7$). **Identify** how many grams of this dihydrate are also required in the preparation of 12.5 g of adipic acid, in a yield of 68.6%.
- 8 a Calcium carbonate decomposes when heated strongly. The following equation represents the decomposition:
- $$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$$
- If 25 g of calcium carbonate decomposes, **calculate** the mass of:
- i CaO produced
 - ii CO_2 produced.
- b If an experiment produces 10.5 g of CaO , **calculate** the percentage yield.



- 9 a Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas.
If 2.6 g of zinc reacts with excess HCl, **calculate** the theoretical values for the mass of:
- zinc chloride produced
 - hydrogen gas produced.
- b If this experiment produces 5.2 g of zinc chloride, **calculate** the percentage yield.
- 10 a Octane (C_8H_{18}) is one of the many hydrocarbons in petrol. It burns in air to produce carbon dioxide and water. Write the balanced equation for the reaction.
- b If 45.6 kg of octane (this is about 65 L of fuel) is burnt in air, **calculate** the mass of:
- oxygen that reacts
 - carbon dioxide produced.
- c i If an experiment where 45.6 kg of octane burns in air produces 105.6 kg of carbon dioxide, **calculate** the percentage yield.
ii **Explain** why the yield is less than 100%.
- 11 In the wine industry, yeast ferments glucose ($C_6H_{12}O_6$) from crushed grapes to produce ethanol (C_2H_5OH) and carbon dioxide.
- a Write a balanced overall equation for the production of ethanol and CO_2 from glucose to **determine** the number of moles of ethanol produced per mole of glucose.
- b In a small-scale laboratory experiment, if 45 g of glucose are fermented, **calculate** the theoretical mass of ethanol produced.
- c If only 13.8 g of ethanol are produced, **calculate** the percentage yield and offer reasons why a 100% yield is not likely in the fermentation of crushed grapes in winemaking.
- 12 a If 4.5 mol of magnesium were burnt in oxygen gas, **calculate** the number of moles of:
- oxygen required for complete combustion of the magnesium
 - magnesium oxide produced.
- b Iron reacts with hydrochloric acid (HCl) to produce iron(II) chloride and hydrogen gas:
- $$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$
- Identify** how many moles of Fe are needed to completely react with 1.5 mol of HCl.
 - Identify** how many moles of H_2 are produced when 3 mol of HCl react.
 - If 4 mol of $FeCl_2$ are produced, **calculate** how many moles of Fe react.
- c Butane (C_4H_{10}) undergoes complete combustion in the presence of oxygen gas to produce carbon dioxide and water. If 5.5 mol of butane react, **calculate** the number of moles of:
- oxygen consumed
 - carbon dioxide produced
 - water produced.

The mole and the Avogadro constant

- A mole is the amount of substance containing N_A particles.
- The Avogadro constant (N_A) is 6.02×10^{23} .



The law of conservation of mass

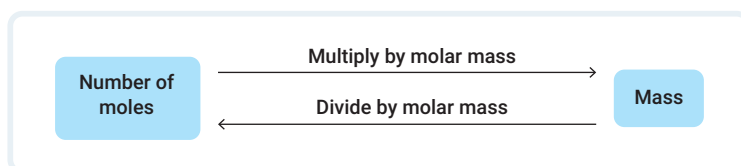
- The law of conservation of mass: mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or physical transformations.
- Balanced chemical equations follow the law of conservation of mass.
- The coefficient in front of each substance in a balanced chemical equation indicates the ratio of each substance in which they react.

The mole and molar mass

- The relative molecular and formula mass (M_r) of a substance in atomic mass units (amu) is the mass of one molecule or formula unit of the substance on a scale in which the mass of an atom of the carbon-12 atom is exactly 12 atomic mass units.
- The molar mass (M) of a substance is the mass of 1 mol of the substance and equals the mass of $N_A = 6.022 \times 10^{23}$ atoms, molecules or formula units of the substance.
- The number of moles of substance can be calculated by:

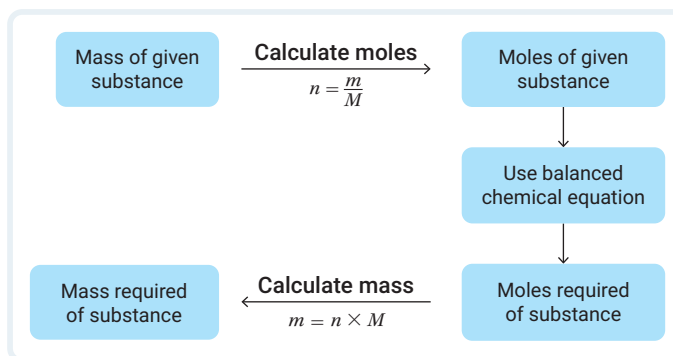
$$\text{Number of moles } (n) = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}} = \frac{m}{M}$$

- Rearranged, the mass of a substance can be calculated by:



Stoichiometry

- Stoichiometry is the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers.
- If the mass of a substance is known, the mass of a required substance can be calculated:



Types of formulas

- The empirical formula of any compound is the simplest whole number ratio in which the atoms of the elements are present in the compound.
- The molecular formula states the actual number of each type of atom present in the molecule.
- The percentage composition is the percentage by mass of each of the different elements in the compound:

$$\text{Percentage composition (\%)} = \frac{\text{mass of element in compound}}{\text{molecular or formula mass of compound}} \times 100$$

Calculating yield

- Percentage yield can be calculated by using the following formula:

$$\text{Percentage yield (\%)} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

- Experimental yield is the actual yield of a product obtained for a reaction and theoretical yield is the amount of a product that is produced from the complete reaction of the limiting reactant.
- A limiting reactant is the reactant that is totally consumed when the chemical reaction is complete. It dictates how much product can be theoretically produced in a chemical reaction.

MULTIPLE CHOICE

- Which of the following is closest to the relative molecular mass of carbon dioxide?
 - 14 amu
 - 22 amu
 - 28 amu
 - 44 amu
- The relative mass of silver chloride is more correctly called:
 - relative atomic mass.
 - relative formula mass.
 - relative molecular mass.
 - relative compound mass.
- How many feathers are there in 2 mol of feathers (correct to 2 significant figures)?
 - 3.0×10^{23}
 - 6.0×10^{23}
 - 1.2×10^{24}
 - Need to know average mass of feathers
- The mass of 1 mol of nitrogen dioxide is closest to which of the following?
 - 30 g
 - 44 g
 - 46 g
 - 6.02×10^{23} g
- What is required to calculate the percentage composition of a substance?
 - Chemical formulas only
 - Relative mass of the compound only
 - Chemical formula and relative mass of the compound
 - Chemical formulas and relative atomic masses of the elements in the compound
- The Avogadro constant is equal to:
 - 6.02×10^{23} units.
 - 6.02×10^{23} g.
 - 12.01 atoms.
 - 12.01 g.
- Which of the following describes the mass of Avogadro's number of particles?
 - 16 g of oxygen gas
 - 35 g of methane
 - 36.5 g of hydrogen chloride
 - 40 g of lithium fluoride
- How many moles of water are produced when 0.03 moles of aluminium hydroxide decompose according to the following equation?

$$2\text{Al}(\text{OH})_3(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + 3\text{H}_2\text{O}(\text{l})$$
 - 0.015 mol
 - 0.03 mol
 - 0.045 mol
 - 0.06 mol

9. Which of the following is both an empirical and a molecular formula?

- A H_2O_2
- B C_2H_4
- C $\text{C}_6\text{H}_{12}\text{O}_6$
- D $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

10. The total number of formula units in 4.5 g of KCl is:

- A 0.06 particles.
- B 2.7×10^{24} particles.
- C 3.6×10^{22} particles.
- D 6.02×10^{23} particles.

SHORT RESPONSE

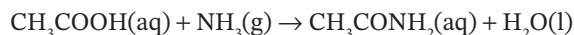
11. Respiration can be represented by the following equation:



- a **Calculate** the molecular mass of glucose.
- b **Calculate** how many moles of carbon dioxide are produced per mole of glucose.
- c **Calculate** how many molecules of water are produced per mole of oxygen.

©VCAA © VCAA VCAA NHT 2023 MCQ Q22, modified

12. The following reaction takes place between acetic acid (CH_3COOH) and ammonia (NH_3) to produce ethanamide (CH_3CONH_2):



- 3.00 g of acetic acid (CH_3COOH), reacts to produce 2.40 g of (CH_3CONH_2)
- Molar mass of acetic acid = 60.06 g mol^{-1}
- Molar mass of ammonia = 17.04 g mol^{-1}
- Molar mass of ethanamide = 59.08 g mol^{-1}

Calculate the percentage yield for this reaction.

CROSS-CHAPTER QUESTION

13. Silver nitrate reacts with iron(III) chloride to give silver chloride and iron(III) nitrate. In a particular experiment, it was planned to mix a solution containing 25.0 g of silver nitrate with another solution containing 45.0 g of iron(III) chloride.

- a
 - i **Determine** the chemical formulas for the substances in the reaction.
 - ii **Identify** the type(s) of compounds used in the reaction.
 - iii **State** the properties of the compounds used in the reaction.
- b Write the balanced chemical equation for the reaction.
- c **Identify** which reactant is the limiting reactant.
- d **Determine** the maximum number of moles of AgCl that could be obtained from this mixture.
- e **Calculate** the maximum number of grams of AgCl that could be obtained.
- f **Calculate** how many grams of the reactant in excess will remain after the reaction is over.

DATA ANALYSIS

14. Analyse data

A student conducted an experiment to find the formula of a compound of copper and oxygen. The sample was heated until only pure metallic copper remained. The following results were obtained:

Items weighed	Mass (g)
Crucible	35.030
Crucible + copper	38.205
Crucible + copper compound	39.005

- a Determine** the empirical formula of the compound.

A student replicated the experiment using the same methodology. The results are shown.

Items weighed	Mass (g)
Crucible	20.500
Crucible + copper	23.675
Crucible + copper compound	23.475

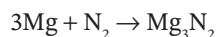
- b** When comparing their results with the original experiment, the student noticed that all the values were different. As a result, they thought that their results were inaccurate. **Explain** why this is not the case. Include any calculations to support your explanation.

15. Interpret evidence

Sometimes in Practical activity 9.2.1, the ratio of Mg to O is determined to be too low (e.g. 0.9 Mg to 1.0 O). This means that there was too much oxygen relative to the mass of magnesium. At other times, the ratio is too large (e.g. 1.2 Mg to 1.0 O), in which case there must have been too little oxygen (or too little weight at the end of the experiment, which registers as too little oxygen).

As you would know from previous experimental work, a yellow safety flame of the Bunsen burner left on while heating the crucible causes soot to be deposited on the outside of the crucible. If the crucible lid is not put back onto the crucible once the magnesium is lit, dense white smoke escapes from the crucible during the burning.

Since magnesium is a reactive metal, it combines readily with both oxygen and nitrogen when it burns in air, forming both the oxide and the nitride, Mg_3N_2 , according to the equation:



A teacher and lab technician repeated this experiment. Explore each scenario below relating to their Mg:O ratios. Suggest possible outcomes to deduce the cause(s) of incorrect ratios.

- a** The teacher's data provided a ratio with too little Mg for the amount of O (e.g. 0.85 Mg:1.0 O).
- b** The lab technician's data provided a ratio with too much Mg for the amount of O (e.g. 1.2 Mg:1.0 O).

Exothermic and endothermic reactions



Maksim Ladouski/Shutterstock.com

SYLLABUS DOT POINTS

SCIENCE UNDERSTANDING

- 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 INQUIRY

Investigate:

- the enthalpy change of a reaction, e.g. calorimetry or Hess's law.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

In a chemical reaction, bonds are broken in the reactants, and bonds are formed to make the products. These changes involve energy.

This topic will explore how endothermic and exothermic reactions relate to the law of conservation of energy. When bonds break or re-form, heat energy is released into or absorbed from the surrounding environment. Heat is a form of energy and temperature is a measure of the average kinetic energy of the particles. Enthalpy level diagrams will be used to apply the relationship between temperature and enthalpy changes to identify thermochemical reactions as exothermic or endothermic. Thermochemical equations will be used to identify the relative stabilities of reactants and products, and the sign of the enthalpy change (ΔH) for a reaction. The reason reactions are exothermic or endothermic will be understood in terms of average bond enthalpies. Calorimetry measurements will be carried out to estimate the heat changes for a substance, given the mass, specific heat capacity and measured temperature change. Calorimetry can be applied to obtain estimates for the enthalpy change for a chemical reaction.

Practicals

- Determining the heat of combustion of alkanols by calorimetry
- Applying Hess's law

Worksheets

- Bond energies
- Calculating heat of combustion
- Hess's law



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Atoms are composed of a nucleus containing protons and neutrons, and electrons in distinct energy levels.
- ✓ Reactants are written on the left side of an equation and products are written on the right.
- ✓ Balanced chemical equations can be written from descriptions of reactions.
- ✓ The number of moles of a substance can be calculated using the mass of the substance divided by the molar mass of a substance.
- ✓ The law of conservation of mass states that matter cannot be created or destroyed in a chemical reaction.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ distinguish between the terms 'heat' and 'temperature'
- ✓ recall the temperature scales used to measure temperature
- ✓ state the law of conservation of energy and explain how it relates to reactions
- ✓ explain exothermic and endothermic reactions in terms of energy change
- ✓ describe the concept of enthalpy and how it relates to exothermic and endothermic reactions
- ✓ sketch enthalpy level diagrams for exothermic and endothermic reactions
- ✓ determine the relative stabilities of the reactants and products and the sign of enthalpy change for the reaction using enthalpy level diagrams
- ✓ write thermochemical equations
- ✓ use average bond enthalpy to explain whether a reaction is exothermic or endothermic
- ✓ calculate the amount of heat energy gained or lost from a reaction using $Q = mc\Delta T$
- ✓ describe how calorimeters work
- ✓ explain Hess's law
- ✓ use Hess's law to calculate the change in enthalpy
- ✓ analyse data for heat of combustion and heat of neutralisation to determine heat, mass, specific heat capacity, temperature and enthalpy change.



Weblink
Temperature

heat the thermal energy transferred from a hotter system to a cooler system

kinetic energy the energy possessed by moving particles

temperature a measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

10.1 Energy and temperature

Temperature is a common indicator of the change in energy of a substance. For example, when water is heated, it absorbs heat energy and its temperature increases. Conversely, when water is cooled, it releases heat energy and its temperature decreases.

Heat refers to a form of energy. More specifically, it is thermal energy transferred from a hotter system to a cooler system that are in contact. Heat is represented by the symbol q or Q , and is measured in joules (J). Other forms of energy include light energy, electrical energy, sound energy, nuclear energy and dark energy.

When heat is transferred to an object, the particles absorb the heat energy and move faster.

Kinetic energy is the energy of motion. Any object that is moving possesses kinetic energy.

Temperature

Temperature is a measure of the average kinetic energy of the particles in the system. The system refers to the chemical reaction. When a substance is heated, some of the absorbed energy is stored within the particles, while some of the energy increases the motion of the particles. This is registered as an increase in the temperature of the substance.

In gases and liquids, this is easy to conceptualise. In solids, the lattice of atoms and molecules is tightly bonded and the average kinetic energy refers mainly to the kinetic energy of the vibrational motion associated with the atoms and/or molecules in the solid. When a molecule in a solid sublimates, the molecule leaves the surface of the solid, travelling faster, but vibrating much less, to join the surrounding gaseous environment.

At any given temperature, not all of the particles of a substance have the same kinetic energy. The particles have a wide range of kinetic energies. Most of the particles have a kinetic energy near the middle of the range. However, a small number of particles have kinetic energies lower or higher than the average, as shown in **Figure 10.1.1**.

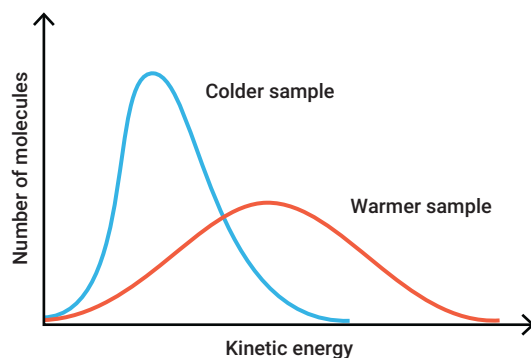


FIGURE 10.1.1 For any given temperature, there is a range of kinetic energy with which particles within the substance can move. Most of the particles have kinetic energy at the top of the curve. The blue curve is for a low temperature, and the red curve is for a high temperature.



Weblink
Kinetic energy

Temperature scales

Temperature can be measured on various scales. In Australia, temperature is commonly reported in degrees Celsius ($^{\circ}\text{C}$) (sometimes also called the centigrade scale). The Celsius temperature scale can have both positive and negative values. On a hot day, the temperatures are higher than 35°C . In some parts of Australia, temperatures can be as low as -10°C . Water freezes at 0°C and, at standard pressure, boils at 100°C (**Figure 10.1.2**). Standard pressure is 100 kPa or 1 bar.

Scientists often use the absolute temperature scale, which is measured in kelvin (K). This scale only has positive values. There is a linear relationship between the Celsius temperature scale and the absolute temperature scale.

KEY FORMULA

$$\text{Temperature } (^{\circ}\text{C}) = \text{temperature (K)} - 273$$

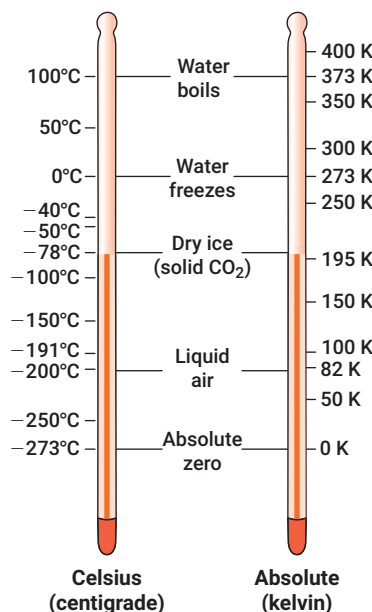


FIGURE 10.1.2 The Celsius and absolute temperature scales

LEARNING CHECK 10.1

DESCRIBING

1 Define:

- a temperature
- b absolute temperature scale
- c kinetic energy.

APPLYING

2 Convert:

- a 14°C to kelvin
- b -11°C to kelvin
- c 206 K to degrees Celsius.

3 Consider the graphs in **Figure 10.1.3**, which shows gases X, Y and Z at three different temperatures.

Which of the gases is at the lowest temperature?

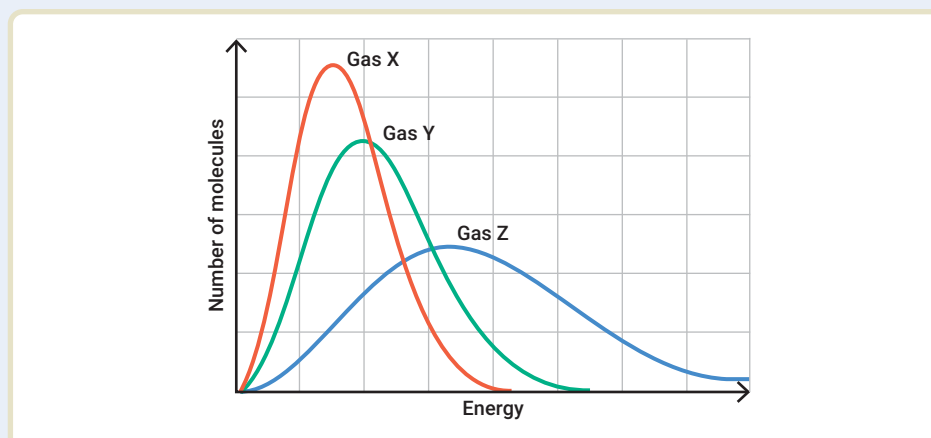


FIGURE 10.1.3 Three gases at different temperatures



Syllabus link
Chapter 8 discussed chemical reactions in further detail.

KEY LAW

The law of conservation of energy

Energy cannot be created or destroyed; it can only be changed from one form to another. The law of conservation of energy dictates that in the breaking and forming of bonds, heat energy is released or absorbed by the system to, or from, the surroundings.

10.2 The law of conservation of energy

Energy is stored in the chemical bonds of reactants and products. The law of conservation of energy states that energy cannot be created or destroyed; it can only be changed from one form to another. When the energy stored in the bonds of the reactants is more than that of the products, the excess energy is released as heat, sound or light during the formation of the products. An example of the law of conservation of energy is the burning of magnesium, which produces a very bright white light.

Recall from Chapter 8 that during a chemical reaction, the reactants transform into products. For this to occur, bonds in the reactants have to break and new bonds in the products form. Consider the reaction involving the burning of magnesium. Once the reaction is kick-started by placing the magnesium ribbon in the flame, no further energy input is required. The bonds in the magnesium metal and oxygen gas are broken and different bonds are formed in magnesium oxide. The burning of magnesium is an

exothermic reaction because energy is released into the surroundings, as shown in **Figure 10.2.1** and the following equation:



The surroundings refers to the rest of the universe around a chemical reaction (the system). In exothermic reactions, the reactants contain more chemical energy than the products. Another exothermic reaction is the reaction between magnesium and hydrochloric acid. This differs from the burning of magnesium in two ways:

1. Only heat is produced.
2. The reaction is spontaneous at room temperature; no energy is needed to start the reaction:



Scientists write the equations for these reactions to include both the chemicals involved and an energy term (energy gained or lost, as heat or light). They are called thermochemical equations.

Physical processes, such as changes of state, can also involve the release of energy. For example, water vapour (steam) cools down when it forms liquid water – it releases energy as heat to the surroundings. This phase change is an exothermic process (but not a reaction). Here, the hot, fast-moving gaseous water molecules (steam) lose kinetic energy in collisions with colder air molecules (mainly nitrogen and oxygen), or a colder surface, progressively slowing down. Once the kinetic energy of the water molecules is less than the forces of attraction between them (intermolecular forces), the water molecules combine, at first just in pairs (dimers), then trimers, tetramers, pentamers ... and eventually they form droplets. The droplets grow and finally combine to form a continuous liquid. The equation representing the phase change is:

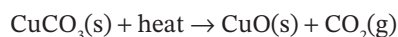


Other chemical reactions are **endothermic**, which means that the products contain more chemical energy than the reactants. This means that energy must be continuously added from the surroundings to allow the chemical reaction to proceed. For example, the reaction between ammonium thiocyanate and barium hydroxide absorbs heat from the surrounding environment:



This reaction is the chemistry behind many instant ice packs. When the reactants are mixed, the temperature drops by a significant amount relatively quickly. As a result, we can usually feel the cold sensation not long after the ice packs are activated.

The decomposition of copper(II) carbonate (CuCO_3) is another example of this type of reaction. This reaction only occurs while the copper(II) carbonate is being heated. The products contain more chemical energy than the reactants so extra energy is put into the reaction. Heat energy is converted to chemical energy.



Physical processes, such as the boiling of water to produce steam, also involve the input of energy. This makes the boiling of water an endothermic process. The water molecules gain kinetic energy from the heat supplied, and as a result they collide with the hot surface of the heated container and rebound with increased kinetic energy. The increased kinetic energy permits water molecules with sufficient kinetic energy to overcome the intermolecular forces holding the liquid together and the molecules leave as fast-moving gaseous molecules and are dispersed into the surrounding air. The energy required to change a liquid to a gas is called the **latent heat of vaporisation**.

The energy required to change a solid to a liquid (the process of melting) is called the **latent heat of fusion**.



FIGURE 10.2.1 Burning magnesium releases energy in the form of heat and light.

exothermic reaction

a reaction in which energy is released to the surroundings, causing temperature to increase

endothermic reaction

a reaction in which energy is absorbed from the surroundings, causing temperature to decrease



Weblink

Exothermic vs endothermic reactions

latent heat of vaporisation

the energy required to change a specific quantity of a substance from liquid to a gas

latent heat of fusion

the energy required to change a specific quantity of a substance from a solid to a liquid

enthalpy (H) the total energy content of a chemical substance

enthalpy change (ΔH) the heat gained or lost per mole of reactant in a chemical reaction:
enthalpy change = (enthalpy of products) – (enthalpy of reactants)

Enthalpy (H) is the energy content of a chemical substance at constant pressure. This includes potential energy, which is the summation of the different energies (e.g. chemical energy) stored in the chemical substance. The **enthalpy change (ΔH)** that occurs as a result of a chemical reaction, also referred to as the 'heat of reaction', uses the symbol ΔH . It is measured in kilojoules (kJ) and can sometimes be expressed in kJ/mol (kJ mol^{-1}) or kJ/g (kJ g^{-1}) of reactant. An endothermic reaction is one in which heat is absorbed by the reactants during a chemical reaction, so, by definition, ΔH is positive because the energy of the chemical system has increased. A negative ΔH indicates that energy has been released during a chemical reaction and is exothermic. As a result, the energy of the system has decreased. The enthalpy change during reactions is expressed in thermochemical equations to show the amount of energy absorbed or released as a result of the reaction. Examples of how to write thermochemical equations are shown in Worked example 10.2.1.

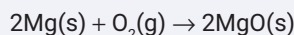
WORKED EXAMPLE 10.2.1

Write thermochemical equations for the following reactions.

- a** Magnesium reacts with oxygen to form solid magnesium oxide. This is an exothermic reaction during which 1204 kJ are released when the reaction proceeds according to the balanced equation.
- b** Solid copper(II) carbonate decomposes to form solid copper(II) oxide and carbon dioxide. This is an endothermic reaction during which 44 kJ mol^{-1} are absorbed per mole of copper(II) carbonate.

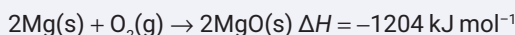
ANSWERS

- a 1 Write a balanced equation for the reaction.**

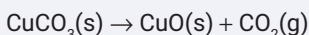


- 2 Include the enthalpy change to the equation to write the thermochemical equation.**

This is an exothermic reaction, so the ΔH is negative.

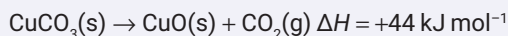


- b 1 Write a balanced equation for the reaction.**



- 2 Include the enthalpy change to the equation to write the thermochemical equation.**

This is an endothermic reaction, so the ΔH value is positive.



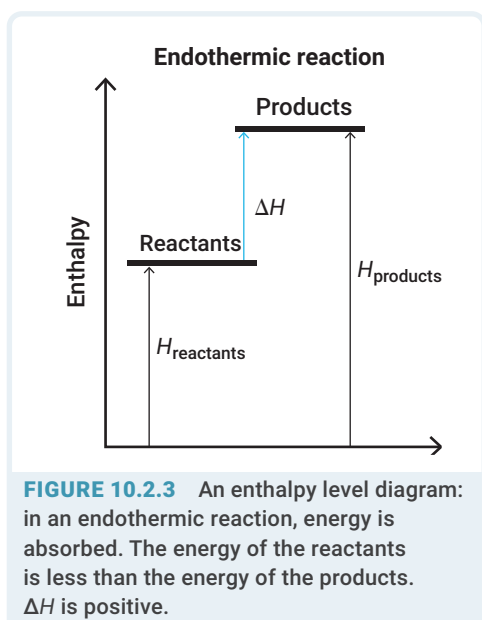
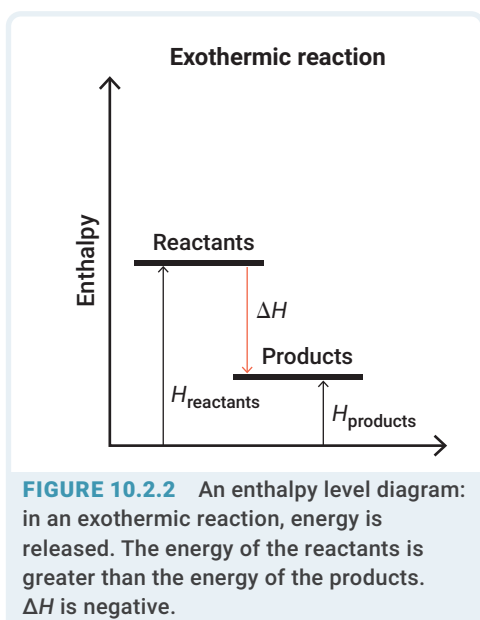
KEY FORMULA

$$\Delta H = H_{\text{(products)}} - H_{\text{(reactants)}}$$

It is important to understand that in thermochemical equations, the ΔH value in kJ mol^{-1} reflects the energy released as per coefficients of the balanced equation, and is not *per mol* of any particular substance.

The difference between the relative energies of reactants and products, and which is the higher or lower, allows you to determine if reactions are exothermic or endothermic. Enthalpy level diagrams show the energy change in reactions. The y-axis of an energy level diagram represents the enthalpy of the substance. The x-axis represents the progression of the reaction from reactants to products. They are simplified illustrations that only show the initial and final energy of substances during a reaction. In the enthalpy level diagram in **Figure 10.2.2**, the reaction is exothermic, energy is released, so ΔH is negative. The red arrow indicates that energy has been lost to the surroundings, which become hotter. There is more energy stored in the bonds of the reactants than the products, so energy is lost.

In **Figure 10.2.3**, the reaction is endothermic, energy is absorbed, so ΔH is positive. The blue arrow indicates that the reactants absorb energy from the surroundings. As a result, the surroundings lose heat and become cooler. There is less energy stored in the reactants, so energy is required to form the products, which have more energy stored in them.



Enthalpy level diagrams can also be used to deduce the relative stability of the reactants and products in a reaction. The higher the enthalpy, the less stable the substance. In exothermic reactions, the products have a lower enthalpy than the reactants, so the products are more stable. The products have less energy, and they have a lower capability to react with another substance. The reactants are more reactive than the products. The reaction releases energy into the surroundings and the change in enthalpy is negative. However, in an endothermic reaction, the products are less stable than the reactants because they have more energy and a higher capability to react with another substance. The products are more reactive than the reactants.

When classifying a process as exothermic or endothermic, scientists measure changes in the temperature of the surroundings. An exothermic process releases heat, so the temperature of the immediate surroundings rises. An endothermic process absorbs heat from the surroundings, so the surrounding temperature decreases.

Table 10.2.1 lists some exothermic and endothermic processes.

TABLE 10.2.1 Exothermic and endothermic processes

Exothermic processes	Endothermic processes
Making ice cubes in a freezer	Melting ice cubes or snow
Snow formation in clouds	Converting frost to water vapour
Water vapour condensing to liquid water (e.g. clouds turning to rain)	Evaporating water
Burning gas or petrol (combustion)	Producing glucose and O_2 by photosynthesis
Mixing sodium hydroxide pellets with water	Mixing water with ammonium nitrate
Mixing water with strong acids	Mixing water with table salt
Nuclear fission	Splitting apart a gas molecule (e.g. O_2 or N_2)
The rusting of iron in air	Baking bread
Burning sugar	Cooking an egg

WORKED EXAMPLE 10.2.2

When heated, solid magnesium carbonate decomposes to produce solid magnesium oxide and carbon dioxide gas. For this reaction to proceed, $117.3 \text{ kJ mol}^{-1}$ of heat energy is required.

- Sketch the enthalpy level diagram for this reaction.
- Determine the relative stabilities of the reactants and products and the sign of enthalpy change for the reaction.

ANSWERS

- a 1 Determine the reactants and the products of the reaction.**

Reactants: magnesium carbonate (MgCO_3)

Products: magnesium oxide (MgO) and carbon dioxide (CO_2)

- 2 Determine if the reaction is endothermic or exothermic.**

Heat is added to start the reaction, so it is an endothermic reaction.

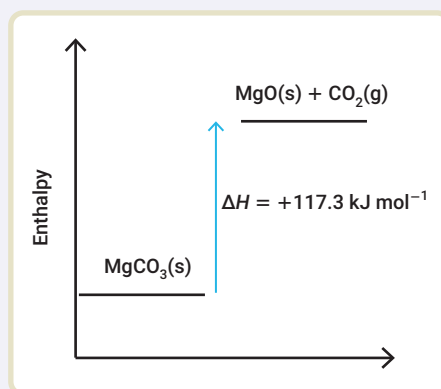
- 3 Sketch the enthalpy level diagram.**

Because it is endothermic, the reactants will be at a lower level than the products. Add the arrow and amount of heat required for the reaction to proceed.

- b Determine the stabilities and sign of enthalpy change.**

The products have a higher enthalpy level than the reactants.

Therefore, the products are more unstable than the reactants. The reaction required energy to be absorbed, so the change in enthalpy is positive.



WORKED EXAMPLE 10.2.3

Carbon dioxide gas forms when solid carbon is reacted with oxygen gas, releasing $393.5 \text{ kJ mol}^{-1}$ of heat energy.

- Sketch the enthalpy level diagram for this reaction.
- Determine the relative stabilities of the reactants and products.

ANSWERS

- a 1 Determine the reactants and the products of the reaction.**

Reactants: carbon (C) and oxygen (O_2)

Products: carbon dioxide (CO_2)

- 2 Determine if the reaction is endothermic or exothermic.**

As the heat is released during the reaction, it is an exothermic reaction.

- 3 Sketch the enthalpy level diagram.**

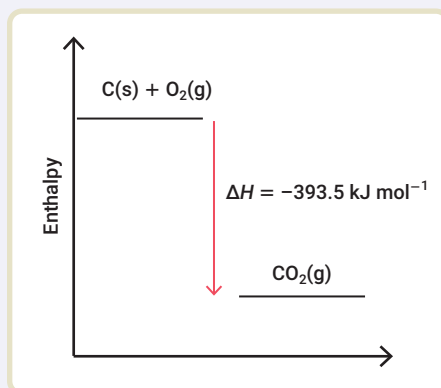
Because the reaction is exothermic, the reactants will be at a higher level than the products. Add the arrow and amount of heat released from the reaction.

- b 1 Compare the relative enthalpies of the reactants and products.**

The reactants have a higher enthalpy level than the products.

- 2 Determine the relative stability.**

Since the reactants have a higher enthalpy than the products, the reactants would be considered less stable than the products.



Chemical bond energy

The **chemical bond energy**, also known as bond enthalpy, is the energy associated with the bond in a chemical substance. Chemical bond energy is the energy required to break a chemical bond and is measured in kilojoules per mole (kJ mol^{-1}). Since energy is required to break bonds, chemical bond energy is always a positive value.

Table 10.2.2 shows some average bond enthalpies for some common bonds. You will find more average bond enthalpies at 298 K in the *Formula and data book*.

chemical bond energy

the amount of energy required to break 1 mole of the stated bond to give separated atoms, usually expressed in units of kJ mol^{-1} ; also known as bond enthalpy

TABLE 10.2.2 Average bond enthalpies for some common bonds at 298 K

Bond	Average bond enthalpy (kJ mol^{-1})
H—H	436
C—H	414
C—C	346
C=C	614
C—O	358
C=O	804



As seen from Table 10.2.2, it takes 414 kJ mol^{-1} to break 1 mol of C—H bonds. Strong bonds have higher bond enthalpies because more energy is required to break a strong bond. When a chemical reaction occurs, the chemical bonds in reactants break and new ones form to make products. The new arrangement does not have the same total energy as the bonds in the reactants. Energy is always required to break a bond; therefore, ΔH is positive and it is endothermic. Energy is released when a bond is made; therefore, ΔH is negative and it is exothermic. Worked example 10.2.4 demonstrates how to determine the overall bond enthalpy given the number of moles of a substance. Worked example 10.2.5 demonstrates how to determine the enthalpy of a reaction using average bond enthalpies. However, this simple method of calculating the enthalpy change of a reaction using average bond energies has certain limitations:

- The values for bond energy in the *Formula and data book* are average values; hence, calculations are estimates only.
- The calculation is for gaseous species (reactants and products) only; hence, the calculated answer for reactions that involve solids and liquids will be even less accurate.

$$\Delta H = (\text{sum of energy required to break bonds}) - (\text{sum of energy released when bonds are formed})$$

WORKED EXAMPLE 10.2.4

If the bond enthalpy for H—Cl is 431 kJ mol^{-1} , determine the overall bond enthalpy of 2 mol of HCl.

ANSWER

- 1 Determine the energy required to break the H—Cl bond.**

431 kJ mol^{-1}

- 2 Calculate the energy required to break 2 mol of H—Cl bonds.**

Multiply the average bond enthalpy of H—Cl by 2.

Overall bond enthalpy for 2 mol HCl = $2 \times 431 = +862 \text{ kJ}$

WORKED EXAMPLE 10.2.5

Calculate the enthalpy change for this reaction and determine if it is endothermic or exothermic:



ANSWER

1 Identify the bonds in the reactants and products of the reaction.

Reactants: one H—H bond and one I—I bond.

Products: two H—I bonds.

2 Identify the average bond enthalpies from the *Formula and data book*.

H—H bonds: 436 kJ mol^{-1}

I—I bonds: 151 kJ mol^{-1}

H—I bonds: 298 kJ mol^{-1}

3 Find the sum of the enthalpies of the reactants and products. Remember, bond energy is always positive.

Reactants: $436 + 151 = 587 \text{ kJ mol}^{-1}$

Products: $2 \times 298 = 596 \text{ kJ mol}^{-1}$

Subtract the value of the bonds formed from the value of the bonds broken.

$$\begin{aligned}\Delta H &= (\text{sum of energy required to break bonds}) - (\text{sum of energy required to form bonds}) \\ &= 587 - 596 = -9 \text{ kJ mol}^{-1}\end{aligned}$$

This is how much energy is released when the bonds on the product side are formed. The net change of the reaction is therefore -9 kJ mol^{-1} . Since this is negative, the reaction is exothermic.

LEARNING CHECK 10.2

DESCRIBING

1 Define:

- a endothermic reaction
- b exothermic reaction
- c enthalpy
- d chemical bond energy.

2 Determine whether the following statements are true or false.

- a Exothermic reactions have negative ΔH values.
- b For endothermic reactions, the products have less energy than the reactants.

APPLYING

3 Write thermochemical equations for the following reactions.

- a Liquid water is formed from hydrogen gas and oxygen gas in a reaction that releases 286 kJ per mole of water.
- b Ammonia gas (NH_3) absorbs 92 kJ mol^{-1} of heat to form nitrogen and hydrogen gases.
- c Calcium oxide solid is formed when 180 kJ of heat is absorbed per mole of solid calcium carbonate. The other product is carbon dioxide.
- d Ammonia gas reacts with oxygen to form water vapour and nitrogen monoxide (NO) gas. The reaction releases 905 kJ of energy.
- e Liquid water freezes to form ice. The latent heat of fusion is 6.02 kJ mol^{-1} .

4 Sketch enthalpy level diagrams for each of the reactions in Question 3.

- 5 The chemical energies of three compounds, X, Y and Z, are labelled in **Figure 10.2.4**.

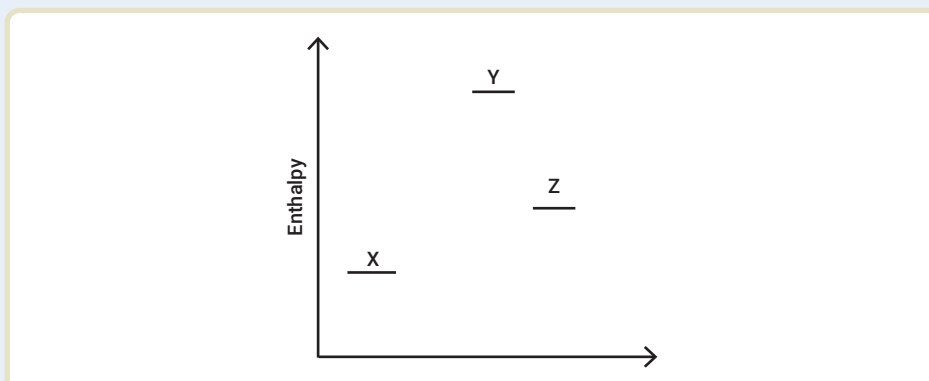


FIGURE 10.2.4 Enthalpy level diagram

- Identify** which of the labelled compounds is highest in energy.
 - Identify** which of the labelled compounds is lowest in energy.
 - Identify** which of the labelled compounds is most stable.
 - Identify** which of the labelled compounds is least stable.
- 6 **Analyse** the enthalpy level diagrams in **Figure 10.2.5** to **determine** the relative stabilities of the reactants and products, and the sign of the enthalpy change (ΔH) for the reactions.

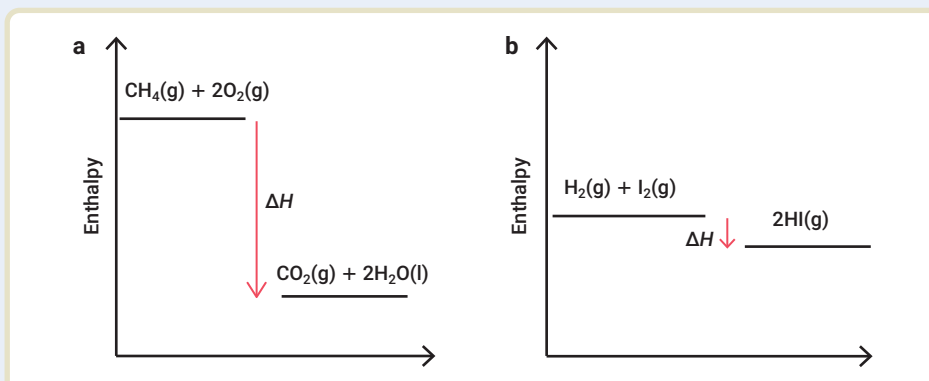
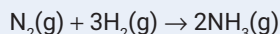
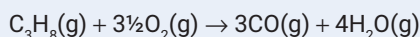


FIGURE 10.2.5 Enthalpy level diagrams

- 7 Use the table of average bond enthalpies in the *Formula and data book* to **calculate** the enthalpy change for the reaction in which gaseous nitrogen reacts with gaseous hydrogen to produce gaseous ammonia as shown in the balanced chemical equation:



- 8 The following equation represents a combustion reaction of propane ($\text{C}_3\text{H}_8(\text{g})$) when the oxygen supply is limited:



Determine the enthalpy change of the reaction, in kJ mol^{-1} , using average bond enthalpy data from *Formula and data book*. The bond enthalpy for the carbon–oxygen triple bond in carbon monoxide (CO) is 1072 kJ mol^{-1} .

10.3 Relationship between temperature and enthalpy

During a chemical reaction, heat energy is either gained by, or absorbed from, the surroundings.

The amount of heat energy (Q) gained or lost depends on the:

- capacity of the substance to absorb heat, the ‘specific heat capacity’ (c), as defined later in this section
- mass of the substances reacting (m), (the mass of the *limiting reactant* in the reaction)
- change in temperature (ΔT) (final temperature – initial temperature).

The amount of heat (Q) gained or lost is directly proportional to each of these three factors, so it depends on the product of these three factors. This can be expressed mathematically as:

$$Q = mc\Delta T \text{ (units J)}$$

This formula can be found in the *Formula and data book*. The heat gained or lost in a chemical reaction can be measured experimentally for a known mass of substance, with known specific heat capacity, by measuring the temperature change from the beginning to the end of the reaction. This method is known as calorimetry. By using this method, enthalpy change is obtained by dividing Q by the number of moles of measured reactant (see Chapter 9 for calculation of moles):

$$\Delta H = \frac{Q}{n} \text{ (units J mol}^{-1}\text{)}$$

It can also be expressed as the heat gained or lost per gram (or kilogram) of reactant, in which case:

$$\Delta H = \frac{Q}{m} \text{ (units J g}^{-1}\text{ or kJ kg}^{-1}\text{)}$$

The units of the above formulas assume that you have used the value calculated for Q . The same formula may be used to calculate ΔH in kJ mol^{-1} ; however, you would need to first convert the value of Q from J to kJ before dividing by the number of moles or mass.

Specific heat capacity

Every substance has a different ability to hold heat. On a hot day, sand feels hot underfoot but grass feels cool (**Figure 10.3.1**). Even though the Sun is shining on both of them, and they are receiving the same amount of heat energy, they feel very different. This is because they have different capacities to hold heat.

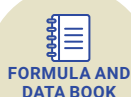
The capacity of different substances to hold heat is defined in terms of a standard unit of measure called the **specific heat capacity (c)**, quoted in units of $\text{J g}^{-1} \text{K}^{-1}$ (or also the same in the alternative units, $\text{kJ kg}^{-1} \text{K}^{-1}$).

Specific heat capacity values give us a way to directly compare the heat retention (or loss) abilities of substances, because they consider the same mass and temperature increase. The only variable is the material’s capacity to hold (or lose) heat.

The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{K}^{-1}$. This means that 4.18 joules of energy is needed to increase the temperature of 1 gram of water by 1 kelvin (or $4.18 \text{ kJ kg}^{-1} \text{K}^{-1}$).



FIGURE 10.3.1 Sand and grass have different capacities to retain heat.



FORMULA AND
DATA BOOK

KEY FORMULA

$$\Delta H = \frac{Q}{n} \text{ (units J mol}^{-1}\text{)}$$

KEY FORMULA

$$\Delta H = \frac{Q}{m} \text{ (units J g}^{-1}\text{ or kJ kg}^{-1}\text{)}$$

specific heat capacity (c)
the heat energy required
to raise the temperature
of 1 gram of a substance
by 1 kelvin



Weblink

Specific heat capacity

Similar to the heat energy associated with a chemical reaction, the total amount of energy needed to heat a substance (without it reacting) depends on the:

- mass of the substance (usually expressed in grams, g)
- specific heat capacity of the substance (usually in the units $\text{J g}^{-1} \text{K}^{-1}$)
- increase in temperature required (expressed in K; since this is a temperature change, the magnitude of the change is the same regardless of whether we calculate the difference of two temperatures in $^{\circ}\text{C}$ or in K).

As highlighted earlier, this may be written as:

$$Q = mc\Delta T \text{ (in J)}$$

A positive temperature change indicates an increase in temperature, meaning that the substance has absorbed heat. A negative temperature change indicates a decrease in temperature, meaning that the substance has released heat.

KEY FORMULA

$$Q = mc\Delta T$$

where: Q = heat change (J)

m = mass (g)

c = specific heat capacity ($\text{J g}^{-1} \text{K}^{-1}$)

ΔT = change in temperature ($^{\circ}\text{C}$ or K).

WORKED EXAMPLE 10.3.1

Calculate the amount of heat, in kJ, added to 100 g water to cause a temperature increase from 18°C to 60°C .

ANSWER

- 1 Extract the data from the question and the *Formula and data book*.**

$$m = 100 \text{ g}, c = 4.18 \text{ J g}^{-1} \text{K}^{-1}, \Delta T = 60 - 18 = 42^{\circ}\text{C} = 42 \text{ K}$$

(This is a temperature *difference*, so the value for the difference is the same in $^{\circ}\text{C}$ as in K.

Check for yourself from the temperature scales illustrated in Figure 10.1.2)

- 2 Substitute the data into the equation.**

$$Q = mc\Delta T$$

$$= 100 \times 4.18 \times 42$$

$$= 17\,556 \text{ J}$$

$$= \frac{17\,556}{1000} = 18 \text{ kJ (correct to 2 significant figures)}$$

The amount of heat added was 18 kJ.

WORKED EXAMPLE 10.3.2

In an experiment, 26.8 kJ of heat was used to increase the temperature of 180 g of water. The initial temperature of the water was 22°C . What was the final temperature of the water?

ANSWER

- 1 Extract the data from the question and the *Formula and data book*.**

$$m = 180 \text{ g}, c = 4.18 \text{ J g}^{-1} \text{K}^{-1}, Q = 26.8 \text{ kJ} = 26\,800 \text{ J}$$

- 2 Rearrange the equation and substitute data to determine the change in temperature.**

$$\Delta T = \frac{Q}{m \times c}$$

$$= \frac{26\,800}{180 \times 4.18}$$

$$= 35.6^{\circ}\text{C}$$

3 Determine the final temperature of the water.

$$\Delta T = T(\text{final}) - T(\text{initial})$$

$$\begin{aligned}\text{where: } T(\text{final}) &= \Delta T + T(\text{initial}) \\ &= 35.6 + 22 = 57.6^{\circ}\text{C}\end{aligned}$$

The final temperature of water was 57.6°C , rounded up to 58°C .

LEARNING CHECK 10.3

DESCRIBING

1 Define:

- a enthalpy change
- b specific heat capacity.

APPLYING

- 2 Calculate how much heat was added if the temperature of 100 g water increased from 10°C to 50°C .
- 3 Calculate the final temperature of 200 g water if 53.5 kJ of heat was added. The initial temperature of the water was 10°C .
- 4 Use the following data to answer the questions.

Substance	Specific heat capacity ($\text{J g}^{-1} \text{K}^{-1}$)
Ethanol	2.46
Vegetable oil	2.00

- a Calculate how much heat is needed to increase the temperature of 15 g of ethanol by 40°C .
- b Determine the mass of vegetable oil that was heated by 2346 J, if the temperature of the oil increased from 30°C to 60°C .

10.4 Calculating energy change

In a high school laboratory, enthalpy changes are often measured in a 'coffee cup calorimeter', as shown in Figure 10.4.1. A calorimeter is a device used to measure the heat released or absorbed during a mechanical, electrical or chemical reaction and for calculating the heat capacity of materials. In a research laboratory, bomb calorimeters are used; they are more complex and fewer errors are associated with the experiment. An example of a bomb calorimeter is shown in Figure 10.4.2.

Calorimetry is the process of measuring the amount of heat released or absorbed during a chemical reaction and can be used in many different experiments. The changes in heat

calorimetry the process of measuring the amount of heat released or absorbed during a chemical reaction

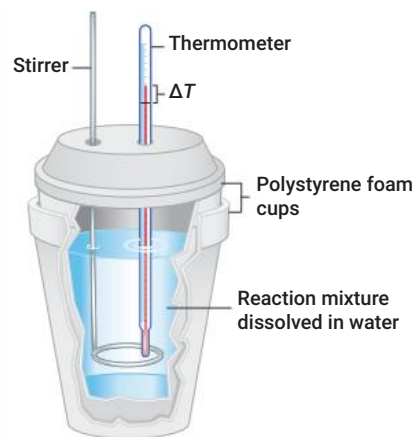
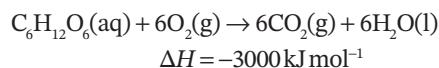


FIGURE 10.4.1 A coffee cup calorimeter can be used to measure the amount of heat involved in the reaction taking place in the cup. The temperature change is measured on the thermometer and used to calculate the amount of heat released or absorbed.

measured in a calorimeter can help us calculate the enthalpy of different types of reactions.

Combustion reactions

As you learnt in Chapter 8, in **combustion** reactions, substances burn in the presence of oxygen and release large amounts of energy. Generally, combustion refers to burning fuels to release the energy needed for our daily activities (**Figure 10.4.3**). The general equation used for respiration (Worked example 8.1.1) is a combustion reaction that occurs in our body cells. Glucose ($\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$) and oxygen gas ($\text{O}_2(\text{g})$) react to produce carbon dioxide gas ($\text{CO}_2(\text{g})$) and water ($\text{H}_2\text{O}(\text{l})$):



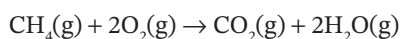
During respiration, energy is released to enable us to breathe, talk, walk and carry out all of our everyday activities.

Since the Industrial Revolution, we have burnt fossil fuels on a large scale to release energy. Power stations convert coal to electricity. Refined petroleum fuels our cars. The source of fossil fuels is decayed animal and plant matter, which takes millions of years to produce. Coal (predominantly elemental carbon), crude oil (hydrocarbons) and natural gas (predominantly methane (CH_4)) are fossil fuels. They are the predominant sources of energy used in Australia.

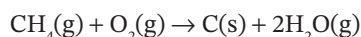
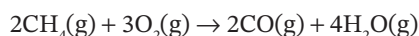
Many fuels contain hydrocarbons, but fuels can also be organic molecules made of carbon, hydrogen and oxygen, such as ethanol ($\text{C}_2\text{H}_5\text{OH}$). Complete combustion occurs when these fuels react with sufficient oxygen (O_2) at standard atmospheric pressure to produce the oxides of carbon and hydrogen – carbon dioxide (CO_2) and water (H_2O). However, sometimes incomplete combustion occurs and carbon monoxide (CO) or carbon, as well as water, are produced. Incomplete combustion occurs when there is not enough oxygen for complete combustion.

For example, methane (CH_4), a major component of natural gas, undergoes complete or incomplete combustion depending on the amount of oxygen available.

Complete combustion produces carbon dioxide and water:



Incomplete combustion produces carbon monoxide or carbon and water:



The formation of carbon or carbon monoxide requires less oxygen per mole of methane than the formation of carbon dioxide does. In the incomplete combustion of methane, the carbon produced ends up as small particles made up of clusters of many carbon atoms, called 'soot'.

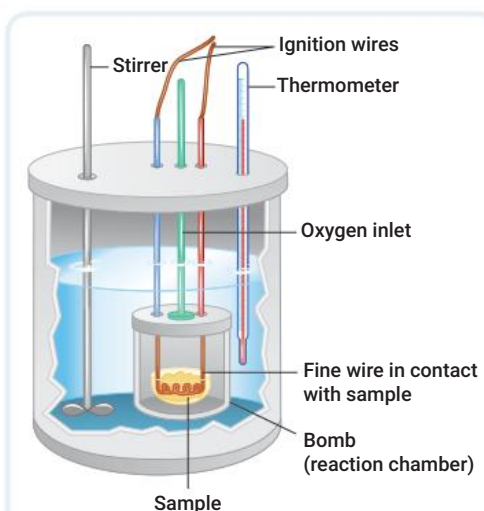


FIGURE 10.4.2 A bomb calorimeter: electricity starts a reaction inside a container that is surrounded by water inside an insulated metal container. The water heats up. By measuring the temperature increase, you can calculate the amount of heat released.

combustion an exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that releases energy

Syllabus link
Chapter 8 discusses combustion.



FIGURE 10.4.3 Burning coal is an example of the combustion of a fossil fuel.

Heat of combustion

heat of combustion (ΔH_c) the heat released when a fuel undergoes complete combustion at standard atmospheric pressure

standard atmospheric pressure atmospheric pressure at sea level, described as 1 atm, or 100 kPa



Worksheet
Calculating heat
of combustion

The enthalpy change involved in combustion reactions is usually called the **heat of combustion (ΔH_c)**. This is an enthalpy of reaction, but it is given a special name. The 'heat of combustion' is further defined as the heat released when a fuel undergoes complete combustion at **standard atmospheric pressure**. Combustion reactions are always considered exothermic reactions. Directly measuring the heat released when a fuel undergoes combustion is difficult, so an indirect method is used by applying the law of conservation of energy. This is the method of calorimetry. The procedure uses the relationship described in section 10.3, connecting heat released in a reaction with mass, specific heat capacity and temperature change.

Calorimetric experiments measure changes in temperature, which is used to calculate changes in heat. In these experiments a known quantity of water is heated by the combustion of a fuel, the temperature change of the water is determined and Q may be calculated. Changes in heat can be measured using the equation:

$$\text{Heat absorbed by water} = \text{heat released by combustion of fuel}$$

The calorimetric method for determining an enthalpy change experimentally is reliable, provided there is no heat loss to the surroundings from the water used to absorb the heat released during the combustion reaction.

WORKED EXAMPLE 10.4.2

A student clamped a conical flask containing 150 g water above a spirit burner that contained hexane (**Figure 10.4.4**). The temperature of the water was measured before and after the experiment and was found to increase by 20°C . The spirit burner containing hexane was weighed before and after the experiment. The mass of hexane that underwent combustion was 0.26 g. Calculate how much heat, in kJ, is released by the complete combustion of 1 g of hexane.

ANSWER

- 1 Write down all of the data for water and determine the heat absorbed by the water.**

$$m = 150 \text{ g}, c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}, \Delta T = 20^\circ\text{C}$$

$$Q = mc\Delta T$$

$$= 150 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 20^\circ\text{C}$$

$$= 12\,540 \text{ J}$$

- 2 Write down the data for hexane and use ratios to determine the heat released by 1 g of hexane.**

The law of conservation of energy:

$$\text{Heat absorbed by the water} = \text{heat released by combustion of hexane}$$

Therefore, the heat released by combustion of hexane = 12 540 J

Mass of hexane that was combusted = 0.26 g

$$\begin{aligned} 1 \text{ g of hexane releases } \frac{1}{0.26} \times 12\,540 \text{ J} &= 48\,231 \text{ J} \\ &= 48 \text{ kJ} \end{aligned}$$

Therefore, 1 g of hexane releases 48 kJ.

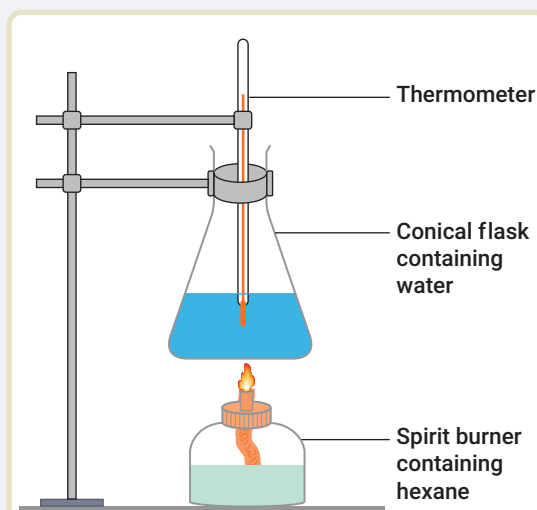


FIGURE 10.4.4 The experimental set-up to determine the heat of combustion for hexane

PRACTICAL ACTIVITY 10.4.1

DETERMINING THE HEAT OF COMBUSTION OF ALKANOLS BY CALORIMETRY

Introduction

Calorimetry was introduced in sections 10.3 and 10.4. Here, a method similar to that outlined in Worked example 10.4.2 is applied to determine the heat of combustion of three alkanols (alcohols) – 1-propanol ($\text{C}_3\text{H}_7\text{OH}$), 1-butanol ($\text{C}_4\text{H}_9\text{OH}$) and 1-pentanol ($\text{C}_5\text{H}_{11}\text{OH}$) – using the calorimetry method. The heat of combustion (ΔH_c) will be stated as the heat (kJ) released by 1 g of the alkanol.

You must compare your measured average value with the values quoted in the scientific literature (NIST 1972): 1-propanol $33.61 (\pm 0.05) \text{ kJ g}^{-1}$, 1-butanol $36.10 (\pm 0.05) \text{ kJ g}^{-1}$ and 1-pentanol $37.77 (\pm 0.05) \text{ kJ g}^{-1}$. You must discuss the significance of any difference found between the literature and your experimental values. The values obtained for ΔH_c can be graphed against the number of C—H bonds in each alcohol in order to seek a correlation and rationalise the results in terms of energy stored in chemical bonds as well as the theoretical calculation of enthalpy change for these combustion reactions.

Research question

How accurate is the heat of combustion of alkanols determined experimentally compared to what is accepted in the scientific literature?

Aim

To determine the heat of combustion for 1 g each of 1-propanol ($\text{C}_3\text{H}_7\text{OH}$), 1-butanol ($\text{C}_4\text{H}_9\text{OH}$) and 1-pentanol ($\text{C}_5\text{H}_{11}\text{OH}$) using the method of calorimetry

Equipment per group

- 100 mL each of 1-propanol ($\text{C}_3\text{H}_7\text{OH}$), 1-butanol ($\text{C}_4\text{H}_9\text{OH}$) and 1-pentanol ($\text{C}_5\text{H}_{11}\text{OH}$)
- 1 L tap water
- 100 mL measuring cylinder
- 50 mL measuring cylinder
- calorimeter (beaker or copper pot) to contain water
- retort stand and clamps or tripod stand, gauze mat
- spirit burner and cap
- 9 aluminium sheets – $15 \text{ cm} \times 15 \text{ cm}$ (for insulating the copper pot or beaker)
- thermometer or temperature probe
- wood spacer blocks to raise the spirit burner if a tripod stand is used for the calorimeter
- matches
- electronic balance ($\pm 0.01 \text{ g}$)
- heat mat



What are the risks in doing this experiment?

- Alcohols are flammable liquids.
- Equipment may get hot during this investigation.
- The spirit burner could leak and spilled alcohol could burn.

How can you manage these risks to stay safe?

- Keep alcohols away from flames; avoid spillage.
- Avoid touching hot equipment.
- Wear safety glasses, gloves and aprons.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your risk assessment before you proceed.

Procedure

- 1 Weigh the dry copper pot/beaker (calorimeter) and record its mass.
- 2 Use a 100 mL measuring cylinder to pour 50 mL tap water into the calorimeter.
- 3 Reweigh the calorimeter and water and record the mass.
- 4 Set up the equipment using the diagram in Figure 10.4.4 as a guide.
- 5 Use a 50 mL measuring cylinder to pour approximately 30 mL of 1-propanol into the spirit burner (or enough to cover the wick). Weigh and record the mass of the spirit burner and 1-propanol.

- 6 Place the spirit burner under the calorimeter, allowing a 2 cm space between the wick and the bottom of the calorimeter.
- 7 Use the aluminium foil to cover the water sample in the calorimeter, and wrap it around the sides and the top of the calorimeter to help insulate the calorimeter from heat loss.
- 8 Pierce a small hole in the aluminium foil. Carefully place the thermometer through the hole and clamp the thermometer so that its bulb is approximately in the middle of the 50 mL volume of water.
- 9 Record the initial temperature.
- 10 Use any strategy available to reduce the effect of air currents on the spirit burner flame.
- 11 Light the spirit burner and wait until the water temperature has risen by approximately 35°C. Keep watch on the temperature. The thermometer can be used to gently stir the water to ensure uniformity.
- 12 Cap the spirit burner to extinguish the flame.
- 13 Record the highest temperature displayed by the thermometer.
- 14 Wait until the spirit burner has cooled. Weigh and record the mass of the spirit burner and remaining propanol to determine the mass of 1-propanol that has undergone combustion.
- 15 Remove the aluminium foil, empty the calorimeter, rinse it with cold water to restore its temperature to room temperature and dry it.
- 16 Refill the spirit burner with 1-propanol to the same level as previously and reweigh.
- 17 Repeat steps 1–15 to gather data (temperature change and mass propanol combusted) for three separate trials of propanol samples.
- 18 Repeat steps 1–17 for 1-butanol and 1-pentanol.

Results

Construct a table like the one below (a spreadsheet is recommended). Show the measured masses for the calorimeter, and calorimeter and water, the initial and final temperatures of the water, and the masses for the spirit burner (and alcohol) before and after combustion

Experiment number	Mass of empty calorimeter (g)	Mass of calorimeter + water (g)	Mass of water (g)	Initial temperature (°C)	Final temperature (°C)	Temperature change (°C)	Initial mass of burner + alcohol (g)	Final mass of burner + alcohol (g)	Mass of alcohol combusted (g)

Analysis of results

- 1 Write a balanced equation for the combustion of each alcohol in air.
- 2 Calculate the mass of water, temperature change and mass of alcohol combusted for each trial.
- 3 Confirm the calculations by carrying out one calculation by hand (calculator) if using a spreadsheet.
- 4 Determine the heat absorbed by the water in the calorimeter for each alcohol sample using $Q = mc\Delta T$, where the specific heat of water c is taken to be $4.18 \text{ J g}^{-1} \text{ K}^{-1}$ (see Worked example 10.3.1). This equals the heat released by the mass of alcohol combusted. Use the procedure in Worked example 10.4.2 to calculate the heat released by 1 g of each of the alcohols. Once again, use your spreadsheet for your calculations but demonstrate the validity of one calculation by hand.
- 5 Using your results for the nine trials (three for each alcohol), determine the average experimental value for the heat of combustion of each alcohol (units: kJ g^{-1}). Include an estimate of the uncertainty of the experimental measurement (refer to the digital chapter for advice on how to arrive at estimates for uncertainty in an experimental procedure).

- 6 Plot both experimental and literature values for ΔH_c for each alcohol versus the number of C—C bonds in the alcohol (all on the same graph). Identify any trends, patterns or relationships observed. Interpret the trends, patterns and relationships to come to a conclusion on how the number of C—C bonds affects the heat of combustion.

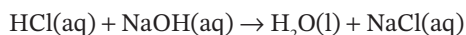
Interpretation

- 7 Compare your experimental value for the heat of combustion of each alcohol with the accepted literature values given in the introduction. Include uncertainties (as percentages or absolute uncertainties) for both.
- 8 Evaluate the reliability of the results for the nine trials and decide whether it is justifiable to retain the three estimates for the heat of combustion for each alcohol. If not, revise your result for the average by eliminating any outlier data (refer to the digital chapter to see how to identify outliers).
- 9 Identify possible sources of error and uncertainty and describe how each would affect the final result. All the heat from the spirit flame was probably not absorbed by the calorimeter (why may this be so?). Explain how this may help to identify differences between your experimental values and the literature values.
- 10 Suggest ways to improve and extend your investigation to minimise each error and uncertainty.
- 11 State the outcome of the experiment.

Heat of neutralisation

The heat of neutralisation is the heat evolved when an acid and a base react to form a salt and water. The enthalpy change depends on the concentration of acid and base involved in the reaction. Neutralisation reactions are generally exothermic because the energy released from the formation of water is greater than the energy absorbed in the reaction. The amount of energy released is proportional to the number of water molecules formed. Molar heat of neutralisation is the energy released per mole of water formed during a neutralisation reaction. The heat of neutralisation can be calculated the same way as the heat of combustion except that there is no burning of fuel to heat the water.

When the acid is added to a base in the reaction, the total mass of the solution in the calorimeter must be found. If we are using dilute aqueous solutions, the density of each solution and hence the density of the combined solution can be assumed to be the same as water. Consider the reaction between hydrochloric acid and sodium hydroxide:



When 50 mL of 1.0 mol L^{-1} of NaOH(aq) is placed in the coffee cup calorimeter, the temperature of NaOH is recorded. Then 50 mL of 1.0 mol L^{-1} of HCl at the same temperature is added and the mixture is stirred. The maximum temperature of the solution is recorded and the temperature changes are calculated. The energy change for the reaction can then be calculated. As the same number of moles of acid and moles of base have reacted, it can be assumed that complete neutralisation has occurred. If the initial temperature of the solutions before addition is 18°C and the final maximum temperature is 24°C , the heat of neutralisation can be calculated:

$$\begin{aligned} Q &= mc\Delta T \\ &= 100\text{g} \times 4.18\text{J g}^{-1}\text{K}^{-1} \times 6^\circ\text{C} \\ &= 2508\text{ J of energy is produced} \end{aligned}$$



Syllabus link

See Chapter 14 for information on concentration. See Chapter 17 for further information on neutralisation.

If 0.05 mol of water are produced, the heat released per mol of water equals:

$$\begin{aligned}\Delta H &= \frac{Q}{n} \\ &= \frac{2508 \text{ J}}{0.05 \text{ mol}} \\ &= -50\,160 \text{ J mol}^{-1} \text{ or } -50 \text{ kJ mol}^{-1}\end{aligned}$$

As you can see, the process is very similar to calculating the heat of combustion. This process will also work with any aqueous solutions. For example, determining the change in enthalpy for sodium hydroxide dissolving in water.

LEARNING CHECK 10.4

DESCRIBING

1 Define:

- a combustion
- b heat of combustion
- c standard atmospheric pressure
- d heat of neutralisation.

APPLYING

- 2** A student clamped a flask with 100 mL water above a spirit burner that contained 1-propanol ($\text{C}_3\text{H}_7\text{OH}$). The temperature of the water increased by 30°C . A mass of 0.372 g of 1-propanol was used during this experiment. **Calculate** how much heat is released by 1 g of 1-propanol.
- 3** A student conducted an investigation to determine the heat of combustion (kJ g^{-1}) of 1-butanol. The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$. The student collected the following data:
- mass of water heated = 150 g
 - initial temperature of water = 16.3°C
 - final temperature of water = 39.6°C
 - mass of 1-butanol used = 0.48 g
- a** Use the student's data to **calculate** the heat of combustion (kJ g^{-1}) for 1-butanol.
- b** The accepted value for the heat of combustion for 1-butanol published in the scientific literature is 36.1 kJ g^{-1} . **Explain** why the student's value could be different from the literature value.
- 4** The heat released when an aqueous solution of ethanoic acid (CH_3COOH) reacted with aqueous sodium hydroxide (NaOH) was measured in the laboratory using a coffee cup calorimeter. The calorimeter consisted of two polystyrene foam coffee cups, one inside the other, which was covered with a polystyrene foam lid through which a thermometer and a stirring rod were placed. The two reacting aqueous solutions, ethanoic acid and sodium hydroxide, were placed in the inner cup. The following data was obtained.

Volume of $1.0 \text{ mol L}^{-1} \text{ CH}_3\text{COOH(aq)}$ used	50.0 mL
Volume of $1.0 \text{ mol L}^{-1} \text{ NaOH(aq)}$ used	50.0 mL
Initial temperature of $\text{CH}_3\text{COOH(aq)}$	25.0°C
Initial temperature of NaOH(aq)	25.0°C
Maximum temperature of mixture	31.0°C

Determine the enthalpy of neutralisation for this reaction in kJ.

10.5 Hess's law and enthalpy change calculations

In calculations involving enthalpy changes for reactions or physical processes, you can apply Hess's law when there is more than one possible mechanism (reaction path) for the reactions that lead to a specific set of products.

Hess's law allows us to determine the enthalpy of a reaction indirectly because sometimes it is too difficult to measure the enthalpy change of a reaction if it is occurring too quickly or too slowly. Hess's law states that an enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs. In other words, if a reaction is carried out in a series of steps, ΔH for the overall reaction equals the sum of the enthalpy changes for the individual steps. The amount of heat energy released or absorbed in a chemical reaction is constant, irrespective of the number of steps or the kind of steps by which the reaction is carried out, provided the same reactants and products are involved.

Figure 10.5.1 shows an example of the two pathways involved in producing carbon dioxide from carbon and oxygen.

KEY LAW

Hess's law

An enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs.

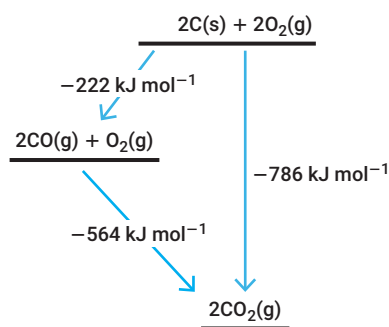


FIGURE 10.5.1 Two pathways for the conversion of carbon to carbon dioxide

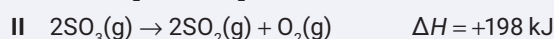
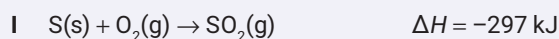
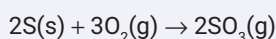
In this example, when carbon and oxygen react to form carbon dioxide, ΔH is -786 kJ mol^{-1} . An alternative pathway to producing carbon dioxide from carbon and oxygen involves first the production of carbon monoxide, which is further reacted with more oxygen to form carbon dioxide. When adding the ΔH of the two steps, the value for the overall production of carbon monoxide is also -786 kJ mol^{-1} .

There are two important rules to remember when applying Hess's law to determine ΔH :

- If a reaction is reversed, so is ΔH .
- If a reaction is doubled/halved, ΔH is also doubled/halved.

WORKED EXAMPLE 10.5.1

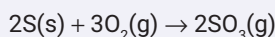
Given the chemical equations I and II, calculate the enthalpy change (ΔH) for the reaction that produces SO_3 :



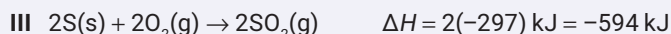
Worksheet
Hess's law

ANSWER

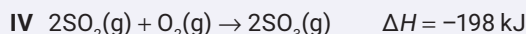
- 1 For each equation, determine what needs to occur in order to achieve the desired reaction.



The desired equation shows 2 mol of sulfur reacting, so equation I must be doubled by multiplying all the coefficients by 2. When you double a reaction, ΔH must also be doubled because twice the energy is released. Applying these changes, you have:

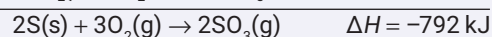
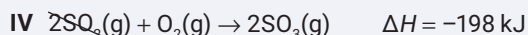


For the desired reaction, SO_3 must be a product, so equation II should be reversed. When an equation is reversed, the sign of ΔH changes. The reverse of equation II is:



- 2 Now, add equations III and IV to obtain the overall equation.

Add the ΔH values to determine ΔH for the overall or net equation. Cancel any terms that are common to both sides of the combined equation. Add the same terms that are on the same side of the equation.



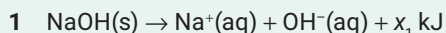
Therefore, the enthalpy change required to form 2 mol of SO_3 is -792 kJ or -396 kJ mol^{-1} . Since 792 kJ of energy is released in the production of 2 mol of SO_3 , it is an exothermic reaction.

PRACTICAL ACTIVITY 10.5.1

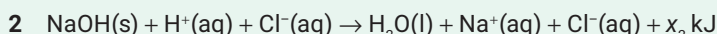
APPLYING HESS'S LAW

Introduction

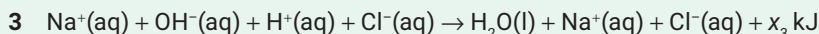
In this experiment, you will measure the amounts of heat released in three exothermic chemical reactions. In reaction 1, solid sodium hydroxide dissolves in water to form an aqueous solution of ions. Water is not included in the equation because it would appear unchanged on both sides.



In reaction 2, solid sodium hydroxide reacts with an aqueous solution of hydrochloric acid to form water and an aqueous solution of sodium chloride.



In reaction 3, aqueous solutions of sodium hydroxide and hydrochloric acid react to give water and an aqueous solution of sodium chloride.



Each of these three reactions takes place in a dilute aqueous solution. For the purposes of this experiment, you may assume that the specific heats of these dilute solutions are the same as the specific heat of pure water, $4.18 \text{ J g}^{-1} \text{ K}^{-1}$. You may also assume that the density of the solution is the same as that of pure water, 1.0 g mL^{-1} .

You will record, as the *initial temperature*, the temperature of the solution before the reactants have been combined. The *final temperature* will be the highest temperature reached by the solution during the reaction. If the mass of the solution and the temperature change associated with the reaction are known for a particular system, the heat of reaction can be calculated.

Research question

Is it possible to experimentally demonstrate Hess's law through reactions involving sodium hydroxide?

Aim

To examine Hess's law experimentally

Equipment per group

- 100 mL $\times$ 0.50 mol L⁻¹ hydrochloric acid (HCl)
- 50 mL of 1.0 mol L⁻¹ hydrochloric acid (HCl)
- sodium hydroxide (NaOH) pellets
- 50 mL of 1.0 mol L⁻¹ sodium hydroxide (NaOH)
- deionised water
- 50 mL beaker
- 400 mL beaker
- 100 mL measuring cylinder
- polystyrene foam cup
- spatula
- thermometer or temperature probe
- electronic balance



What are the risks in doing this experiment?	How can you manage these risks to stay safe?

Copy and complete the risk assessment table in your write-up. Add any risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

Reaction 1

- Measure 100 mL of deionised water into the polystyrene foam cup. Place the polystyrene foam cup in a 400 mL beaker for support.
- Weigh a sample of solid NaOH pellets that is as close to 2.00 g as possible in the 50 mL beaker. Record the mass of the NaOH pellets.
- Measure the temperature of the water in the polystyrene foam cup in the 400 mL beaker to the nearest 0.5°C. Record this initial temperature.
- Add the weighed NaOH pellets to the water in the polystyrene foam cup. Stir the mixture *gently* with the thermometer, until all the solid has dissolved. Record the highest temperature reached during the reaction.
- Pour the solution into the sink and flush it down with water. Rinse the polystyrene foam cup and thermometer with water.

Reaction 2

- Measure 100 mL of 0.5 mol L⁻¹ HCl into the polystyrene foam cup.
- Weigh 2.00 g of solid NaOH pellets as in step 2. Record the mass of the pellets.
- Measure and record the temperature of the HCl solution in the polystyrene foam cup.
- Add the weighed NaOH pellets to the acid solution and stir until the solid is dissolved. Record the highest temperature reached by the solution during the reaction.
- Pour the solution into the sink and flush it down with water. Rinse the polystyrene foam cup and thermometer with water.

Reaction 3

- Measure 50 mL of 1.0 mol L⁻¹ HCl and pour this volume into the polystyrene foam cup.
- Rinse out the measuring cylinder and add to it 50 mL of 1.0 mol L⁻¹ NaOH.
- Determine the temperature of the HCl solution to the nearest 0.5°C. Rinse the thermometer and use it to measure the temperature of the NaOH solution. Record the temperature of each solution.
- Pour the NaOH solution into the HCl solution in the polystyrene foam cup. Stir the mixture gently and record the highest temperature obtained during the reaction.
- Pour the solution into the sink and flush it down with water. Rinse the polystyrene foam cup and thermometer with water.

Results

Results for reaction 1 – water + NaOH(s)

Mass of NaOH(s) (g)	
Initial temperature (°C)	
Final highest temperature (°C)	

Results for reaction 2 – 0.5 mol L⁻¹ HCl(aq) + NaOH(s)

Mass of NaOH(s) (g)	
Initial temperature (°C)	
Final highest temperature (°C)	

Results for reaction 3 – 1.0 mol L⁻¹ HCl(aq) + 1.0 mol L⁻¹ NaOH(aq)

Initial temperature of HCl(aq) (°C)	
Initial temperature of NaOH(aq) (°C)	
Average initial temperature (°C)	
Final highest temperature (°C)	

Analysis of results

- 1 Calculate the change in temperature for each reaction. Enter your results in the analysis table.
- 2 Calculate the mass of the reaction mixture in each reaction. (To do this, first determine the total volume of the solution. Then, calculate the mass of solution, based on the assumption that the density of the solution is the same as that of pure water, 1.0 g mL⁻¹.) Record your results in the analysis table.
- 3 Calculate the total heat released in each reaction, assuming that the specific heat of the solution is the same as that of pure water, 4.18 J g⁻¹ K⁻¹). Record your results in the analysis table.
- 4 Calculate the number of moles of NaOH used in reactions 1 and 2 and record the results in the analysis table.
- 5 Calculate the number of moles of NaOH used in reaction 3, from the concentration of the solution (1.0 mol L⁻¹ or 1.0 mol of NaOH per litre of solution) and the volume used. The calculation is:

$$\frac{1 \text{ mol NaOH}}{1000} = \frac{x \text{ mol NaOH}}{50 \text{ mL}}$$

Enter this result in the analysis table.

- 6 Calculate the energy released per mole of NaOH for each reaction and enter your results in the analysis table.

Analysis table					
Reaction	ΔT	Mass of reaction mixture	Total heat released	Moles NaOH reacted	Heat released per mol NaOH
1					
2					
3					

Interpretation

- 7 Apply Hess's law to demonstrate the relationship between reactions 1, 2 and 3. Include the energy released per mole of NaOH in each equation.

- 8 Examine all the class data. Does the sum of the energy released per mole of NaOH for reactions 1 and 3 equal the energy released per mole of NaOH in reaction 2? Identify factors that might account for any difference.
- 9 Calculate the percentage difference between the heat given off in reaction 2 and the sum of the heats given off in reactions 1 and 3. Assume that the heat given off in reaction 2 is correct.

$$\text{Percentage difference (in evolved heat)} = \frac{|\text{heat}_2 - (\text{heat}_1 + \text{heat}_3)|}{\text{heat}_2} \times 100$$

- 10 Would changing the amount of NaOH used in reaction 1 affect the value obtained for the energy given off per mole of NaOH? Explain your answer.
- 11 Taking into account your answer to Question 4, explain why you were asked to use exactly 2.00 g of NaOH in reactions 1 and 2, and an equivalent number of moles of NaOH in reaction 3.
- 12 Propose what is meant by the additivity of heats of reaction.

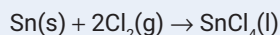
LEARNING CHECK 10.5

DESCRIBING

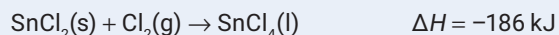
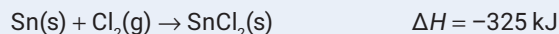
- 1 **Define** 'Hess's law'.

APPLYING

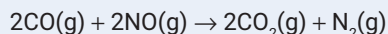
- 2 **Calculate** ΔH for the following reaction:



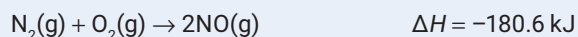
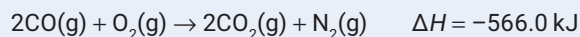
given these two equations:



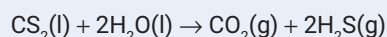
- 3 **Calculate** ΔH for the following reaction:



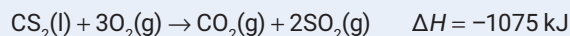
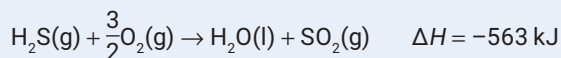
given these two equations:



- 4 **Calculate** ΔH for the following reaction:



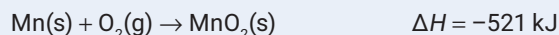
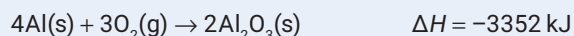
given these two equations:



- 5 **Calculate** ΔH for the following reaction:



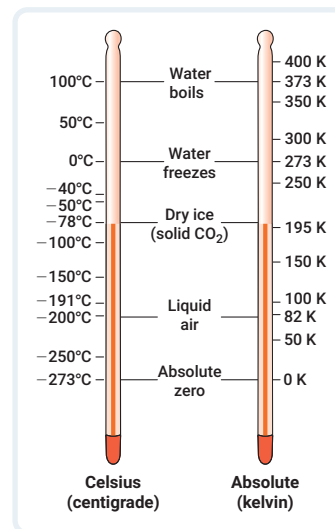
given these two equations:



CHAPTER SUMMARY

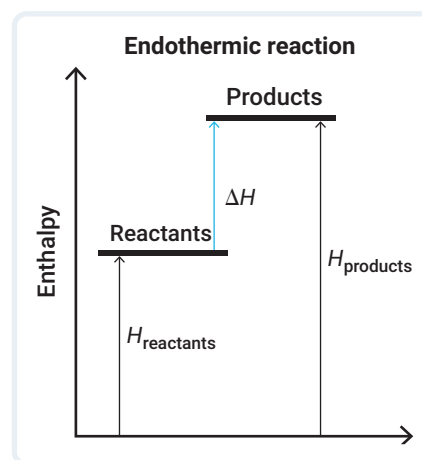
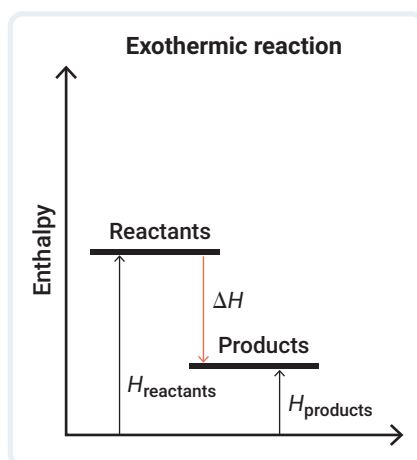
Energy and temperature

- Temperature is a measure of the average kinetic energy of the particles of an object. It is measured on the Celsius ($^{\circ}\text{C}$) or absolute (K) scale. As the temperature increases, the energy of the particles increases.
- Heat (Q) is defined as the energy that flows from one object to another because of a difference in temperature.
- The law of conservation of energy states that energy cannot be created or destroyed; it can only be changed from one form to another.



Enthalpy, endothermic reactions and exothermic reactions

- Enthalpy (H) is the heat absorbed in a chemical reaction at constant pressure.
- The change in enthalpy uses the symbol ΔH and is measured in J or kJ. It is calculated using the equation $\Delta H = H_{\text{products}} - H_{\text{reactants}}$.
- The heat of combustion refers to the enthalpy change involved in combustion reactions and the heat of neutralisation refers to the enthalpy change involved in an acid-base neutralisation reaction.
- Energy is required to break bonds and energy is released when bonds form.
- An exothermic reaction is a reaction in which energy is released to the surroundings, causing a temperature increase. Exothermic reactions have negative ΔH values.
- An endothermic reaction is a reaction in which energy is absorbed from the surroundings, causing a temperature decrease. Endothermic reactions have positive ΔH values.



Relationship between temperature and enthalpy

- The heat change (Q) for a substance can be calculated given the mass, specific heat capacity and temperature change.

KEY FORMULA

$$Q = mc\Delta T$$

where: Q = heat change (J)

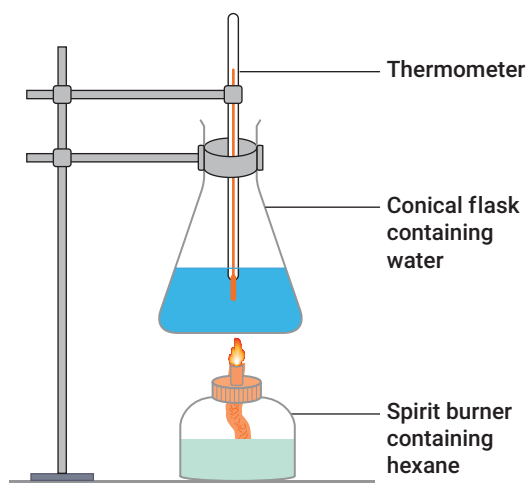
m = mass

c = specific heat capacity of the substance ($\text{J g}^{-1} \text{K}^{-1}$)

T = temperature

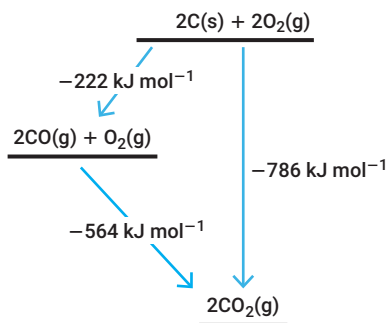
ΔT = change in temperature.

- Specific heat capacity (c) is the amount of heat needed to increase the temperature of 1 g of a substance by 1 K.
- Calorimetric experiments measure the change in heat.



Hess's law and enthalpy change

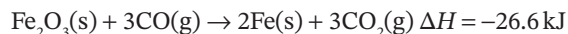
- Hess's law states that an enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs.
- According to Hess's law, a series of reactions can be added to find the enthalpy change for the net reaction. The total enthalpy change is the sum of the enthalpy changes for the individual steps.



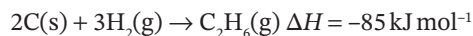
MULTIPLE CHOICE

- Which of the following statements is false?
 - Chemicals react to achieve stability.
 - Mass is conserved in chemical reactions.
 - Bonds are broken and new bonds form in chemical reactions.
 - More energy is required to break bonds than to form bonds in a chemical reaction.
- If energy does not need to be continuously added to a reaction to make it occur at room temperature, then it can be referred to as a:
 - reversible reaction.
 - combustion reaction.
 - exothermic reaction.
 - endothermic reaction.
- A student adds a chemical to water and stirs it with a thermometer until it has dissolved. The student correctly deduces that it is endothermic. What observation could the student have made to deduce this?
 - The chemical dissolved; it would not dissolve if it were exothermic.
 - The temperature did not change, so energy needs to be added.
 - The temperature of the solution was lower than the initial temperature of the water.
 - The temperature of the solution was higher than the initial temperature of the water.
- At what temperature does absolute zero occur?
 - -273 K
 - 0 K
 - 0°C
 - 273°C
- Water has a specific heat capacity of $4.18\text{ J g}^{-1}\text{ K}^{-1}$ and ethanol has a specific heat capacity of $2.46\text{ J g}^{-1}\text{ K}^{-1}$. Which is true?
 - Ethanol is always hotter than water.
 - 2 grams of ethanol has a higher specific heat capacity than water.
 - Less energy is required to raise the temperature of water by 1 K than to raise the temperature of the same mass of ethanol.
 - More energy is required to raise the temperature of water by 1 K than to raise the temperature of the same mass of ethanol.
- A dietician used the phrase 'Food is fuel! It provides us with energy'. Which of the following reactions does this describe?
 - Combustion
 - Respiration
 - Both combustion and respiration
 - None of the above
- Which of the following demonstrates that complete combustion has occurred/is occurring?
 - The flame of a Bunsen burner is blue.
 - The flame of a Bunsen burner is orange.
 - The bottom of a beaker is black after being heated with a spirit burner containing a fuel.
 - Limewater does not turn cloudy when the gas released upon burning a fuel is bubbled through it.

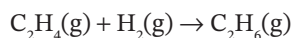
8. Which statement is correct for this reaction?



- A 13.3 kJ are released for every mole of Fe produced.
B 26.6 kJ are absorbed for every mole of Fe produced.
C 26.6 kJ are released for every mole of Fe produced.
D 53.2 kJ are released for every mole of Fe produced.
9. What is the temperature rise when 2100 J of energy is supplied to 100 g of water?
A 0.2°C
B 5°C
C 20°C
D 278 K
10. Consider the enthalpy changes of the following reactions.



What is the enthalpy change, in kJ mol^{-1} , for the reaction between ethene (C_2H_4) and hydrogen?



- A -137 kJ mol^{-1}
B -33 kJ mol^{-1}
C $+33 \text{ kJ mol}^{-1}$
D $+137 \text{ kJ mol}^{-1}$

SHORT RESPONSE

11. Copper(II) nitrate is heated in a crucible over a Bunsen burner, forming solid copper(II) oxide, nitrogen dioxide and oxygen gas.
a Write a balanced chemical equation for this reaction.
b Is this an exothermic or endothermic reaction? **Explain** your answer using bond energies.
12. The complete combustion of ethanoic acid ($\text{CH}_3\text{COOH}(\text{l})$) to form water ($\text{H}_2\text{O}(\text{l})$) and $\text{CO}_2(\text{g})$, at constant pressure releases 875 kJ of heat per mol of acetic acid.
a Write a balanced thermochemical equation for this reaction.
b Draw an enthalpy level diagram for the reaction.

CROSS-CHAPTER QUESTION

13. Ammonia (NH_3) is formed when nitrogen gas is reacted with hydrogen gas.
a Draw the Lewis structure for ammonia.
b **Determine** if ammonia is a covalent molecular substance or a covalent network substance.
c State the properties of ammonia based on your decision in part b.
d Write a balanced chemical equation for this reaction.
- The enthalpy change for this reaction is -46 kJ per mole of ammonia produced.
- e **Identify** if this reaction is exothermic or endothermic.
f **Identify** if the reactants or products are the most stable in this reaction. Provide a reason for your answer.
g **Calculate** how much energy is given off when 222.4 g of N_2 is used in this reaction.

DATA ANALYSIS

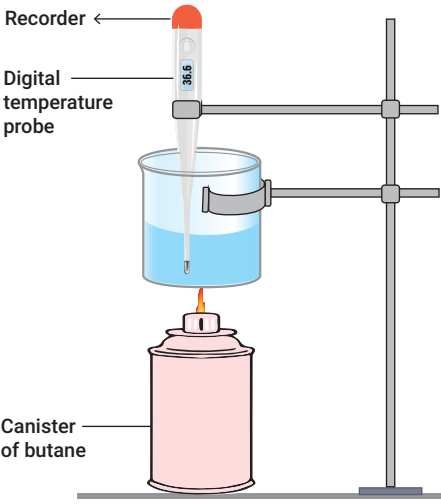
©VCAA VCAA 2016 SA Q10

14. A senior Chemistry student bought a packet of Krispy Krackers from the local farmers' market. The label on the packet had no information on the energy content of the biscuits.

The student decided that they would measure the energy content of a Krispy Krackers biscuit by burning it under a can of water and measuring the temperature rise of the water.

Having performed a similar experiment in class, they knew that when the biscuit was burnt, heat energy would be lost to the environment. To increase the accuracy of the result, some butane gas from a butane canister was first burnt and the temperature rise of the water from that was measured. The heat energy released by burning butane was known, and the percentage energy loss using the equipment could be determined and adjusted to get the result for the biscuit.

The experimental set-up and the results for Part 1 of the experiment are shown below.



Part 1 – The heat content of butane

1. Measure the initial mass of a butane canister.
2. Measure the mass of a metal can, add 250 mL of water and re-weigh.
3. Set up the apparatus as in the diagram and measure the initial temperature of the water.
4. Burn the butane gas for 5 minutes.
5. Immediately measure the final temperature of the water.
6. Measure the final mass of the butane canister when cool.

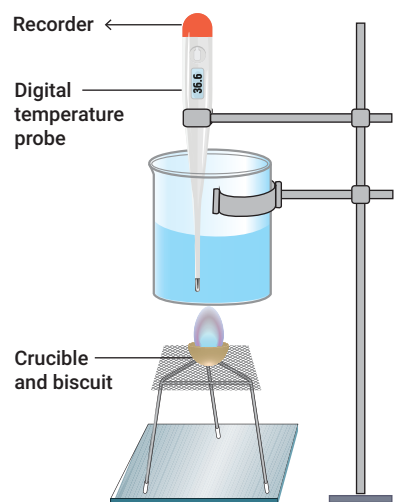
Results table for Part 1

Quantity	Measurement
Mass of empty can	52.14 g
Mass of can + water before combustion	303.37 g
Mass of butane canister before heating	260.15 g
Mass of butane canister after heating	259.79 g
Initial temperature of water	22.1°C
Final temperature of water	32.7°C

- a i Write the balanced thermochemical equation for the complete combustion of butane.
ii **Calculate** the amount of heat energy absorbed by the water when it was heated by burning butane. Give your answer in kilojoules.

- iii **Calculate** the experimental value of the molar heat of combustion of butane. Give your answer in kJ mol^{-1} .
- iv The known energy change for butane is 2877 kJ mol^{-1} . **Calculate** the percentage error for this experiment.
- v **Discuss** ways in which measurement error, instrumental uncertainty and the nature of the methodology influence uncertainty and limitations in the data.
- vi **Identify** if the errors are systematic, random or both. Provide reasoning for your response.

The experimental set-up and results table for Part 2 of the experiment are shown below.



Part 2 – The heat content of a Krispy Kracker

1. Measure the mass of a crucible, add a biscuit and re-weigh.
2. Set up the apparatus as in the diagram, using the same can of water as used in Part 1, and measure the initial temperature of the water.
3. Burn the biscuit until the flame runs out.
4. Immediately measure the final temperature of the water.
5. Measure the final mass of the crucible when cool.

Results table for Part 2

Quantity	Measurement
Mass of crucible	44.33 g
Mass of biscuit + crucible before combustion	46.75 g
Mass of crucible after combustion	44.34 g
Mass of water (from Part 1)	251.23 g
Initial temperature of water	28.5°C
Final temperature of water	34.9°C

- b i **Calculate** the energy content of Krispy Krackers using the data in the results table for Part 2. Give your answer in $\text{kJ}/100 \text{ g}$.
- ii **Explain** why the energy content of Krispy Krackers cannot be given in kJ mol^{-1} .
- c Assume that the same percentage error occurred when burning the Krispy Krackers as when the butane was burnt in Part 1. **Calculate** a more accurate value of the energy content of Krispy Krackers in $\text{kJ}/100 \text{ g}$.
- d Suggest some improvements and extensions to minimise uncertainty, address limitations and improve the overall quality of evidence.

SCIENCE AS A HUMAN ENDEAVOUR

Syllabus dot point

- Evaluate fuels, including fossil fuels and biofuels, in terms of their energy output, their suitability of purpose, and the nature of products of combustion.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Fossil fuels versus biofuels

Fossil fuels are materials such as coal, oil and natural gas that contain hydrocarbons. They are formed naturally in Earth's crust over millions of years from the remains of dead plants and animals. Fossil fuels are mostly used for cooking or heating, to power engines and to generate electricity and are regarded as a non-renewable energy source.

Biofuels are fuels that are produced over a short time span from biomass. They can be produced from plants or from agricultural, domestic or industrial biowaste. Biofuels are mostly used as a fuel for vehicles. Biofuels are regarded as a renewable energy source. In general, they emit fewer greenhouse gas gases and are generally considered to be carbon neutral because the carbon emitted has been captured from the atmosphere by the crops used in production. The use of biofuels has recently been subject to criticism because of the food versus fuel debate, sustainability, deforestation and loss of biodiversity.

Energy output

Fossil fuels, such as diesel, which is a mixture of hydrocarbons with boiling points 150–380°C, are known for their high energy density, meaning they can release a significant amount of energy when burnt. This high energy output has historically made them a preferred choice for various applications.

Biofuels, such as biodiesel, which is chemically similar to regular diesel except it has an ester group on the end of the carbon chain, typically have lower energy densities than fossil fuels. However, they can still provide a considerable amount of energy, particularly when produced efficiently. Advanced biofuels, such as cellulosic ethanol and biodiesel from algae, are being developed to enhance energy output and compete more effectively with fossil fuels.

We can compare the energy output of the two types of fuels by looking at the combustion of both fuels (**Figure 1**).

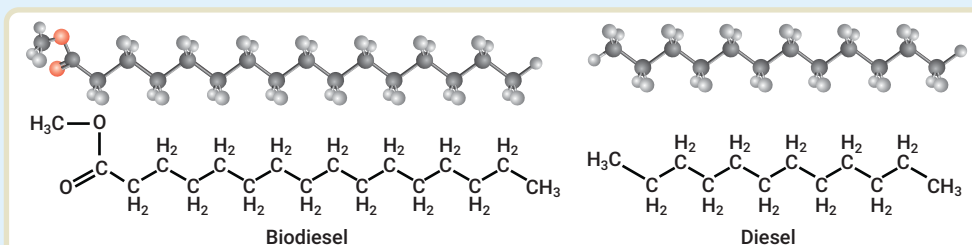
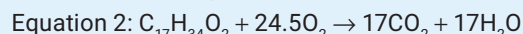
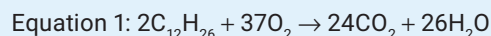
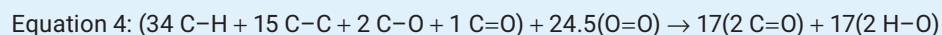
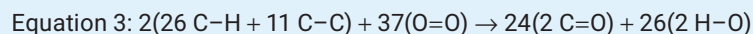


FIGURE 1 Expanded structural formulas for diesel and biodiesel

The reason fossil fuels are such a common source of fuel is that a relatively small amount can produce a lot of energy. The energy of the fuel source is stored within the bonds of the molecules and the amount of energy released can be calculated by looking at bond energies.



By using the bond energies from section 10.2, we can calculate that 1 mol of diesel releases 7551 kJ and that 1 mol of biodiesel releases 10 091 kJ of energy. In terms of quantity, biodiesel releases more energy per mole than diesel.

However, diesel and biodiesel do not have the same molar mass, so comparing the energy content for equivalent masses of each fuel gives a clearer picture of energy comparison. For example, 500 g of diesel releases 22 124 kJ and 500 g of biodiesel releases 18 668 kJ of energy. Therefore, diesel releases more energy by mass than biodiesel.

Suitability for purpose

Fossil fuels have been widely used for decades because of their versatility and reliability. They can be easily transported, stored and used in existing infrastructure, making them suitable for various applications, including electricity generation, heating, transportation (via petrol and diesel) and industrial processes.

Biofuels have advantages in sustainability and reduced greenhouse gas emissions. They are particularly well suited where direct electrification or alternative energy sources are not feasible, such as long-haul transportation (e.g. aviation and shipping) and certain industrial processes (e.g. heat-intensive manufacturing).

Nature of products of combustion

Combustion of fossil fuels releases carbon dioxide (CO_2), sulfur dioxide (SO_2), nitrogen oxides (NO_x), particulate matter and other pollutants into the atmosphere. These emissions contribute to air pollution, smog formation, acid rain and climate change.

Combustion of biofuels also produces CO_2 emissions, but because the carbon in biofuels is derived from recently captured atmospheric CO_2 through photosynthesis, biofuel combustion is considered carbon neutral over time. Furthermore, biofuels typically emit lower levels of pollutants such as sulfur and nitrogen compounds and burns cleaner than fossil fuels, resulting in reduced air pollution and fewer adverse health effects.

In a study by McCarthy et al. (2011), the authors compared the performance and emission characteristics of various biodiesel blends with petroleum diesel in an internal combustion engine. Two types of biodiesel were blended with diesel: biodiesel A was made from beef, pork and sheep fat and canola oil; biodiesel B was made from chicken fat and waste cooking oil.

As shown in the results, while some greenhouse gas emissions were higher than from petroleum diesel, others were lower (e.g. NO_x , Figure 2). Overall, biodiesel A produced lower emissions across all parameters than diesel and biodiesel B.

Although fossil fuels have long been valued for their high energy output and versatility, their combustion releases pollutants and greenhouse gases that contribute to environmental degradation and climate change. Biofuels offer a more sustainable alternative with lower emissions and a reduced environmental impact, although their lower energy density and challenges in scaling up production may limit their suitability for

certain applications. Continued research and development efforts are essential to maximise the benefits of biofuels and to transition towards a more sustainable energy future.

Reference

McCarthy, P., Rasul, M.G., and Moazzem, S. (2011). 'Comparison of the performance and emissions of different biodiesel blends against petroleum diesel'. *International Journal of Low-Carbon Technologies*, 6(4), 255–260.

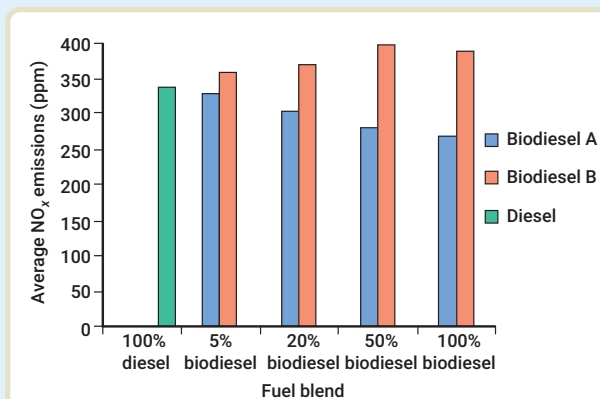


FIGURE 2 Comparison of NO_x emissions of diesel and two different biodiesels in a variety of fuel blends

UNIT 2

Molecular interactions and reactions



iStock.com/Claudia Prommegger

Topic 1: Intermolecular forces and gases

CHAPTERS RELATED TO THIS TOPIC AREA: 11–13

Topic 2: Aqueous solutions and acidity

CHAPTERS RELATED TO THIS TOPIC AREA: 14–18

Topic 3: Rates of chemical reactions

CHAPTER RELATED TO THIS TOPIC AREA: 19

Water is essential for life on Earth and, despite being a common substance, it is unusual compared with other compounds. It is an excellent solvent because of the hydrogen bonds it can form with solutes. The acidity or alkalinity of solutions can be measured using the pH scale, which also determines the reactions in which a solution takes part. It is important to understand how quickly these reactions occur by measuring the rate of a chemical reaction, particularly when external conditions are varied, and to be able to represent and explain these changes.

UNIT OBJECTIVES

By the end of this unit, students should be able to:

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

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Intermolecular forces



Freefile/Shutterstock.com

SYLLABUS DOT POINTS

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. (Hybridisation 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 INQUIRY

Investigate:

- 3D models of linear, bent, trigonal planar, tetrahedral and pyramidal molecules.*

*Note: Simulations may be used.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Chemical bonding and molecular structure are crucial to our understanding of the natural world. However, bonding between atoms is only part of the story. Scientists also need to understand the concept of intermolecular forces – the forces that act between molecules, binding them together and influencing their physical properties. Without this, the idea that lumps of solid water could float in liquid water would be unthinkable! Scientists would not be able to understand why oil and water do not mix. We would be shocked to see a balloon filled with radon gas falling to the ground like a lump of lead.

Practical

- Molecular model building

Worksheets

- Determining the shape and polarity of molecules
- Bonding and polarity
- Intermolecular forces



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ Electronegativity refers to the ability of an atom to attract electrons to itself, where fluorine is the most electronegative element on the periodic table.
- ✓ Covalent bonds are formed from the sharing of electrons.
- ✓ The Lewis structure shows the valence electrons of covalently bonded atoms.
- ✓ Recall how to draw Lewis structures.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ identify bonding and non-bonding electrons in a Lewis structure
- ✓ predict the shape of a molecular structure using the VSEPR theory
- ✓ state bonding angles of simple molecular shapes: linear, bent, trigonal planar, trigonal pyramidal, tetrahedral
- ✓ identify the polarity of a bond by using the difference in electronegativities of two elements in a covalent bond
- ✓ determine the overall polarity of a molecule
- ✓ name the three types of intermolecular forces and compare their relative strengths
- ✓ list factors that affect the strength of each type of intermolecular force
- ✓ identify the types of intermolecular forces that occur between molecules based on the polarity and the shape of molecules
- ✓ understand how the different intermolecular forces affect observable properties of substances such as boiling point and melting point, solubility and vapour pressure.

11.1 Molecular structure: valence shell electron pair repulsion (VSEPR) theory

The shape of simple molecules can be predicted by inspection of the outer or valence electron pairs around the central atom of the molecule. As we learned in Chapter 5, these pairs may be bonding or non-bonding. In bonding pairs, electrons form a covalent bond with another atom. In non-bonding pairs, the electrons from the central atom have not formed covalent bonds (**Figure 11.1.1**).

Non-bonding electrons (lone pair)

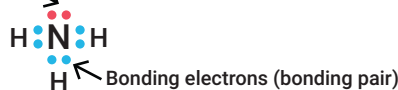


FIGURE 11.1.1 Electrons around the central atom can be bonding or non-bonding.

Predicting the shape of molecules

The pairs of electrons around the central atom, being negatively charged, repel each other, and try to get as far away as possible from each other. This repulsion of the pairs of electrons contributes to determining the shape of a molecule. This theory is known as the valence shell electron pair

repulsion (VSEPR) theory. In this theory, multiple bonds are counted as one bond. For example, the C=O double bond in carbon dioxide counts as one bonding pair, even though four electrons are present in the bond. The different types of electron pairs have different repelling effects, as shown in **Figure 11.1.2**. The greatest repulsion occurs between lone pairs, an intermediate repulsion between lone pair, and bonding pair, and the least repulsion between bonding pairs.

The bonding and non-bonding pairs of electrons arrange themselves around the central atom to minimise the electrostatic repulsion between them, forming the various shapes of molecules in nature shown in **Table 11.1.1**. These shapes are modified by the presence of non-bonding pairs of electrons.

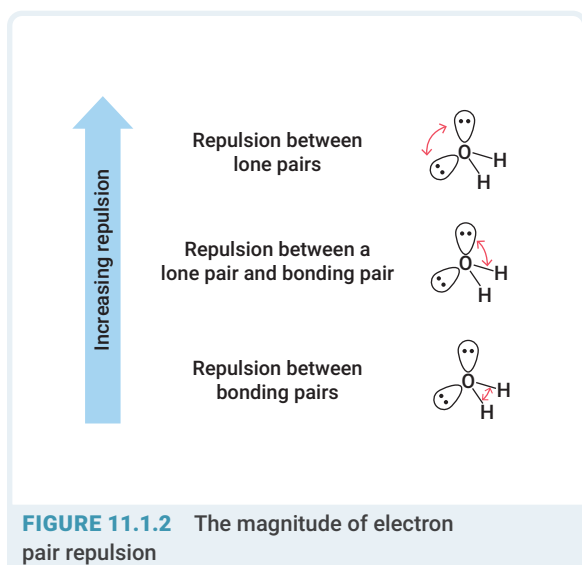


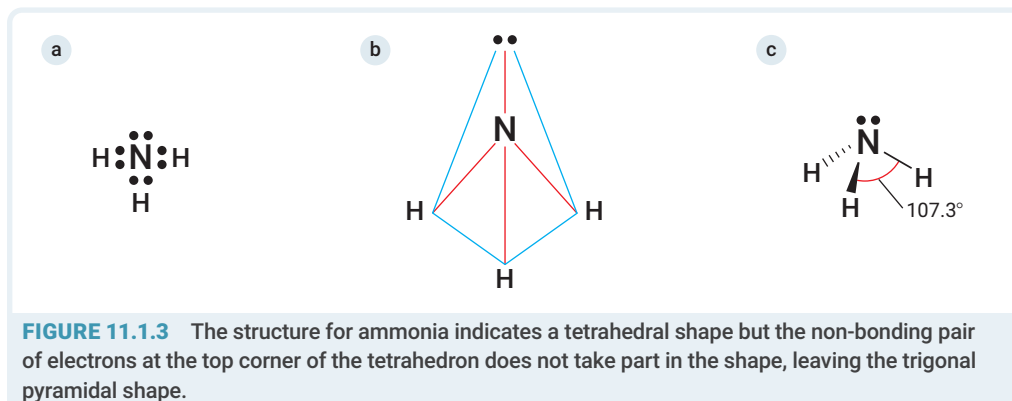
TABLE 11.1.1 Shapes associated with different molecules

Shape	Visual representation
Linear	
Bent	
Trigonal planar	
Tetrahedral	
Trigonal pyramidal	

Ammonia

The shape of the ammonia molecule (NH_3) is modified by a non-bonding pair of electrons.

At first glance, a formula of NH_3 suggests that the molecule has a trigonal planar shape because there are three N—H bonds, so three bonding pairs of electrons. However, the Lewis structure shows an extra pair of electrons, a non-bonding pair, around the central nitrogen atom (**Figure 11.1.3a**). As shown in the structural formula in **Figure 11.1.3b**, the lone pair of electrons is at the apex of the tetrahedron and the three nuclei of the hydrogen atoms are at the other three corners with the nitrogen atom in the centre. The non-bonding pair at the top has greater repulsion force, pushing the bonding pairs of electrons away, resulting in an angle of 107.3° between the bonding pairs. This forms a trigonal pyramidal-shaped molecule as shown in the valence structure (**Figure 11.1.3c**).



Weblink
The shape of molecules



Syllabus link
Table 5.4.2 in Chapter 5 showed the different diagrams used to represent covalent bonds.

Water

Another example of how non-bonding pairs of electrons can affect the shape of a molecule is seen in water (H_2O). H_2O has two bonding pairs of electrons, so it may be anticipated that it would form a linear shape: $\text{H}-\text{O}-\text{H}$.

However, the Lewis structure (Figure 11.1.4a) shows four pairs of electrons around the central oxygen atom: two bonding pairs with the hydrogen atoms, and two non-bonding pairs. The two non-bonding pairs at the top and bottom have greater repulsion force, pushing the bonding pairs of electrons away so that there is an angle of 104.5° between the bonding pairs. This forms a bent-shaped molecule (Figure 11.1.4c).

Syllabus link
Chapter 14 discusses the structure and bonding in water in more detail.

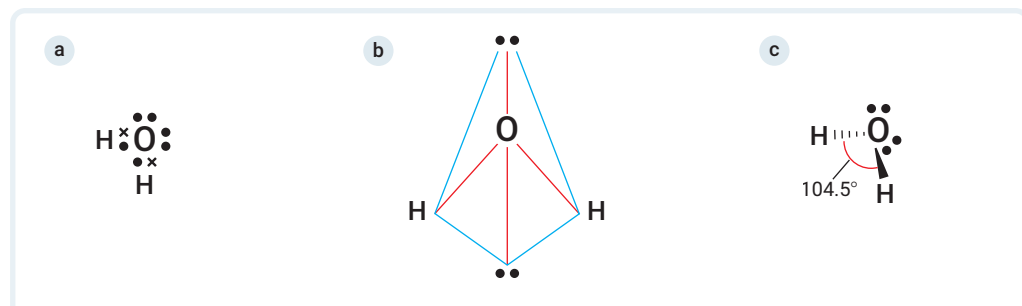


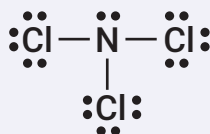
FIGURE 11.1.4 The structure for water indicates a tetrahedral shape but the two non-bonding pairs of electrons at two of the corners of the tetrahedron do not take part in the shape, leaving the bent shape.

WORKED EXAMPLE 11.1.1

Predict the shape of a molecule of nitrogen trichloride (NCl_3).

ANSWER

- 1 Draw the valence structure of the molecule.



- 2 Determine the number of bonding and non-bonding pairs on the central atom.

Nitrogen is the central atom; it has three bonding pairs and one non-bonding pair.

- 3 Use VSEPR theory to determine how the pairs are arranged around the central atom.

The four pairs of electrons around the nitrogen atom are distributed in a tetrahedral shape, with the non-bonding pair repelling the bonding pairs of electrons.

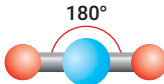
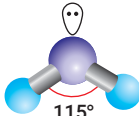
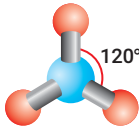
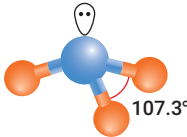
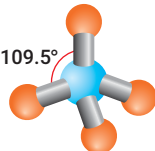
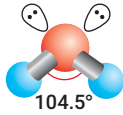
- 4 Deduce the shape by considering just the atoms in the molecule.

Looking at just the three chlorine atoms, it can be determined that the shape of the molecule is trigonal pyramidal.

Bond angles

As we have seen, non-bonding pairs of electrons affect the bond angles in a molecule. A bond angle is the angle between the atoms bonded to the central atom. This is because these non-bonding pairs exert stronger repulsive forces between each other, and between a non-bonding pair and a bonding pair, than the repulsive forces *between* bonding pairs. Table 11.1.2 shows the shapes of molecules associated with bonding and non-bonding pairs of electrons, along with the bond angles.

TABLE 11.1.2 Shapes of molecules associated with bonding and non-bonding pairs of electrons
M = central atom, X = bonding pairs, E = non-bonding pairs

General formula	Number of bonding electron pairs	Number of non-bonding electron pairs on central atom	Molecular shape
MX_2	2	0	Linear – 180° (e.g. CO_2) 
MX_2E	2	1	Bent – 115° (e.g. NO_2) 
MX_3	3	0	Trigonal planar – 120° (e.g. BF_3) 
MX_3E	3	1	Trigonal pyramidal – 107.3° (e.g. NH_3) 
MX_4	4	0	Tetrahedral – 109.5° (e.g. CH_4) 
MX_2E_2	2	2	Bent – 104.5° (e.g. H_2O) 



Weblinks
Molecule shapes
VSEPR and bond angles

Being able to understand the shape of molecules is useful for understanding and predicting the polarity of molecules (see section 11.2). However, there are some limitations with the VSEPR theory. For example, VSEPR theory lacks the precision and scope to handle all aspects of molecular shapes, particularly in the presence of complex bonding situations, electron delocalisation, and larger or transition metal-containing molecules.

LEARNING CHECK 11.1

DESCRIBING

1 Define:

- a bonding electron pair b non-bonding electron pair c valence electrons.

APPLYING

2 Draw the Lewis structure for the following molecules. Identify the shape of the molecule using VSEPR theory.

- a Oxygen gas b Nitrogen gas
c Methane d Phosphorus trichloride
e Trichloromethane f Formaldehyde (CH_2O)

3 Carbon dioxide (CO_2) and sulfur dioxide (SO_2) both have their central atoms double bonded to two oxygen atoms. Draw their shapes and explain the differences, if any, between them.

4 Ethyne (acetylene) has the formula C_2H_2 . It contains a C to C triple bond. Use VSEPR theory to predict its shape.

11.2 The polarity of molecules

In elemental molecules such as nitrogen (N_2), hydrogen (H_2) and chlorine (Cl_2), the electrons are evenly shared between the atoms that make up the covalent bond. However, with molecules such as hydrogen chloride (HCl), ammonia (NH_3) and water (H_2O) that involve different elements, the electrons are not evenly shared. This is because of the difference in electronegativity of the atoms involved in the bond.

Electronegativity

Recall that electronegativity is the relative ability of an atom in a molecule to attract electrons to itself. This relative ability is measured using the Pauling scale of electronegativity. Table 11.2.1 shows some electronegativity values of important elements. Helium, neon and argon do not have values because, as noble gases, they have full outer electron shells and no electron-attracting power.



Syllabus link
See Chapters 3 and 5 to revise electronegativity.



FORMULA AND
DATA BOOK



Worksheet
Determining the shape and
polarity of molecules

TABLE 11.2.1 Electronegativity values of some important elements

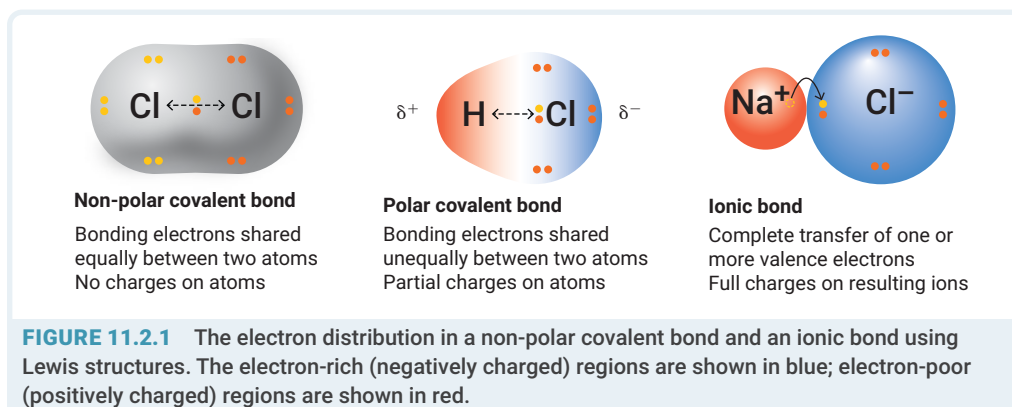
Element	Electronegativity
H	2.2
He	—
Li	1.0
C	2.6
N	3.0
O	3.4
F	4.0
Ne	—
Cl	3.2
Br	3.0
Ar	—

Since elements have different electronegativity values, bonds involving atoms of different elements have a slightly uneven sharing of the bonded electrons. As a result, one atom can attract electrons towards itself better than the other. This difference in electronegativity affects the **polarity** of a covalent bond. The greater the difference, the greater the polarity of the bond. As a general rule, bonds where there is a difference in electronegativity of less than 0.4 are considered **non-polar** covalent bonds, whereas bonds that have a difference of 0.5 to 1.0 are considered slightly **polar** covalent bonds. A bond that has a difference in electronegativity of 1.0–1.7 is considered to be a strongly polar covalent bond. Due to the uneven distribution of bonding electrons in polar covalent bonds, each atom in the bond carries a partial charge (Figure 11.2.1). This partial charge is represented by the symbol δ^+ for a partial positive charge and δ^- for a partial negative charge.

polarity a measure of the unequal sharing of electrons in a covalent bond; the result of the difference in electronegativity of bonded atoms

non-polar describes bonds or molecules that have relatively equal sharing of electrons

polar describes bonds or molecules that have uneven sharing of electrons



Hydrogen

The hydrogen molecule (H_2) consists of two hydrogen atoms bonded together by a single covalent bond. The two atoms in the bond are identical, so they have the same electronegativity value of 2.2. This results in no difference in electronegativity between the two atoms. This means that the $H-H$ bond is non-polar and the hydrogen molecule is non-polar. The electrons are shared equally between the two hydrogen atoms.

Hydrogen chloride

The hydrogen chloride (HCl) molecule consists of a hydrogen atom bonded to a chlorine atom by a single covalent bond. The two atoms making up this bond are different; they have different electronegativity values where the electronegativity of hydrogen and chlorine is 2.2 and 3.2 respectively. This results in a difference in electronegativity of 1.0, which means that the $H-Cl$ bond is polar (Figure 11.2.2).

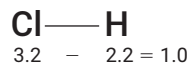


FIGURE 11.2.2 The polarity of the $H-Cl$ bond, showing the difference in electronegativity between chlorine and hydrogen

The difference in electronegativity means that the bond has a slightly positive pole (δ^+) and a slightly negative pole (δ^-). The separation of these partial positive and negative charges is known as a **dipole**. It can help to think of a covalent bond as a tug of war between the two elements.

dipole oppositely charged poles at each end of a bond or molecule

In **Figure 11.2.3**, the chlorine atom is more electronegative than the hydrogen atom and can ‘pull’ the electrons towards itself, thereby becoming slightly more negative. The hydrogen atom becomes slightly more positive. As a result, hydrogen chloride is a polar molecule.

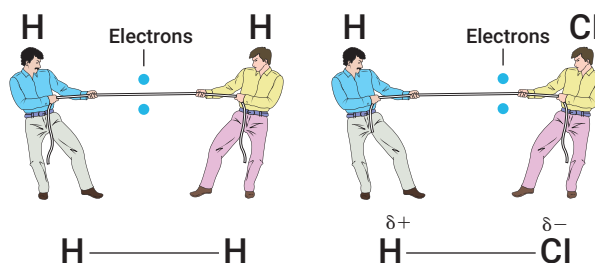


FIGURE 11.2.3 The ‘tug of war’ between two elements, showing the polarity of the H—Cl bond compared to the H—H bond

WORKED EXAMPLE 11.2.1

Determine the polarity of a C—O bond.

ANSWER

1 Determine electronegativity values of each element.

Carbon = 2.6

Oxygen = 3.4

2 Calculate the difference in electronegativity (EN) of the two elements by subtracting the lowest value from the highest value.

$$\Delta\text{EN} = 3.4 - 2.6 = 0.8$$

3 Determine the polarity of the bond.

The oxygen is more electronegative than the carbon, so it has a slightly negative charge. The carbon has a slightly positive charge. With a difference of 0.8, it is a polar bond.



Worksheet
Bonding and polarity

Polar and non-polar polyatomic molecules

For molecules with two or more different atoms, the presence of a polar covalent bond does not necessarily mean that the whole molecule is polar. Molecules can be non-polar despite the presence of polar bonds. For the molecule to be polar, a polar bond must be present and there must be asymmetry in the molecular shape.

Carbon dioxide (CO_2) is a linear molecule with a central carbon atom and an oxygen atom on each side. The molecule is symmetrical (**Figure 11.2.4a**). The C=O bond is polar, with the dipole pointing towards the oxygen because it is more electronegative, but there is no net separation of charge in the overall molecule. Each oxygen atom attracts the electrons but,

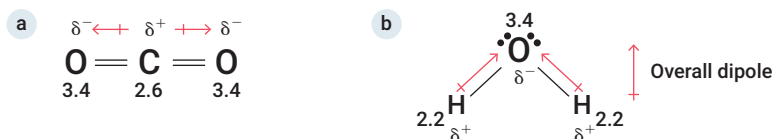


FIGURE 11.2.4 (a) The non-polar carbon dioxide molecule (CO_2) and (b) the polar water molecule. The arrows point from the positive end to the negative end of the dipole.

because they are on either end of the molecule, the charge is distributed over the whole molecule. It is like having an equally strong person on each end of a rope in tug of war; the rope does not move because the pull from each end is equal. The dipoles cancel each other out. Although the molecule has polar covalent bonds, there is no net separation of the charges, so no net dipole. Therefore, the molecule is non-polar.

Contrast the polarity of carbon dioxide with water as shown in **Figure 11.2.4b**. In water, the O—H bond is polar. Oxygen is central and the two hydrogens are at the bottom of the bent shape. This means that the hydrogen end of the molecule is slightly positive. The water molecule has an asymmetrical shape. The presence of polar covalent O—H bonds and a bent shape means that the water molecule is polar. The dipoles cancel out in carbon dioxide but add together in water.

The ammonia (NH_3) and tetrachloromethane (carbon tetrachloride, CCl_4) molecules have different shapes and polar bonds. Ammonia, a trigonal pyramidal-shaped molecule, is polar. The nitrogen atom at the top of the pyramid is slightly negative due to its greater electronegativity. The hydrogen end of the molecule is slightly positive. Because of the molecule's asymmetry, it is polar (**Figure 11.2.5**). Tetrachloromethane has polar bonds owing to differences in electronegativity between the carbon and chlorine atoms – it is tetrahedral in shape with a central carbon atom. The dipole in each of the C—Cl bonds points to the chlorine atom. However, the negative regions on chlorine in CCl_4 cancel out because of the symmetry of the molecule, which is why CCl_4 molecules are non-polar (**Figure 11.2.5**). **Table 11.2.2** shows examples of molecules of different molecular shapes and their overall polarities.



Weblink
Polar and non-polar
molecules

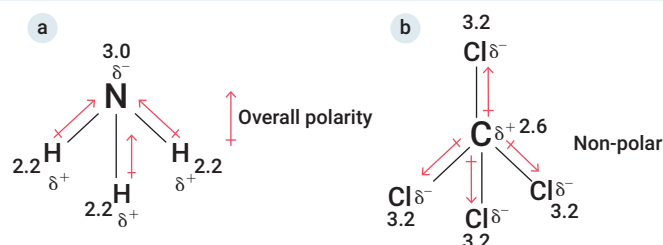
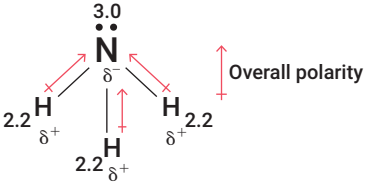
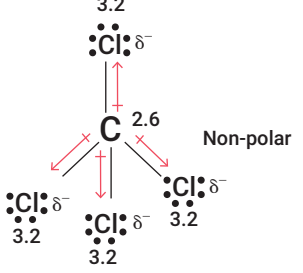
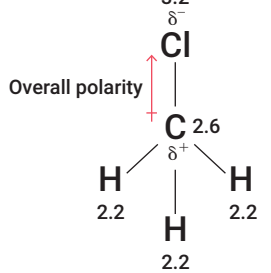
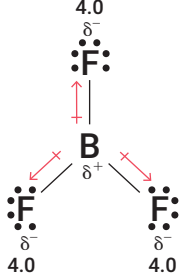


FIGURE 11.2.5 (a) The molecular shape of ammonia (NH_3) and (b) tetrachloromethane (carbon tetrachloride, CCl_4)

TABLE 11.2.2 Molecule polarities of molecules of different molecular shapes

Molecule	Bond polarity (difference in electronegativity)	Molecule shape	Symmetry	Molecule polarity
Hydrogen	$2.2 - 2.2 = 0$ Non-polar bond	Linear $\text{H}-\text{H}$ 2.2 2.2	Symmetrical	Non-polar molecule
Hydrogen chloride	$3.2 - 2.2 = 1.0$ Polar bond	Linear $\text{H}-\text{Cl}$ 2.2 3.2	Asymmetrical	Polar molecule
Carbon dioxide	$3.4 - 2.6 = 0.8$ Polar bond	Linear $\text{O}=\text{C}=\text{O}$ 3.4 2.6 3.4	Symmetrical	Non-polar molecule
Water	$3.4 - 2.2 = 1.2$ Polar bond	Bent $\text{H}-\text{O}-\text{H}$ 2.2 3.4 2.2	Asymmetrical	Polar molecule

Ammonia	$3.0 - 2.2 = 0.8$ Polar bond	Trigonal pyramidal 	Asymmetrical	Polar molecule
Carbon tetrachloride	$3.2 - 2.6 = 0.6$ Polar bond	Tetrahedral 	Symmetrical	Non-polar molecule
Chloromethane (CH ₃ Cl)	$(\text{C}-\text{H}) = 2.6 - 2.2 = 0.4$ Non-polar bond $(\text{C}-\text{Cl}) = 3.2 - 2.6 = 0.6$ Polar bond	Tetrahedral 	Asymmetrical	Polar molecule
Boron trifluoride	$4.0 - 2.0 = 2.0$ Polar bond	Trigonal planar 	Symmetrical	Non-polar molecule

PRACTICAL ACTIVITY 11.2.1

MOLECULAR MODEL BUILDING

Research question

How can the shape of covalent molecules be represented by using models?

Aims

- To use VSEPR theory to predict the shapes of simple molecules and apply this to build models of the molecules to demonstrate their three-dimensional shape
- To determine the polarity and intermolecular forces of simple molecules

Procedure

- 1 Copy the following table and use your understanding of VSEPR theory to complete it.
- 2 Conduct research to find out the bond angles in each molecule.
- 3 Build each model using a molecular model kit or straws and beads.
- 4 Determine the polarity of each molecule once built.
- 5 State the intermolecular forces that occur between molecules.

Molecular formula	Lewis structure	Predicted shape using VSEPR (including bond angles)	Polarity of molecule	Intermolecular forces present between molecules
HF				
BCl ₃				
PH ₃				
CO ₂				
H ₂ S				
N ₂				
CH ₄				
CHF ₃				

Interpretation

- 1 How did your models compare with your predictions?
- 2 Compare the bond angles for H₂S and PH₃. Can you explain any differences?
- 3 Compare the bond angles for CH₄ and CH₃F. Can you explain any differences?
- 4 Using the information you have gathered in this activity, write a conclusion addressing the relative repulsive power of single bonds, double bonds and lone pairs of electrons.

LEARNING CHECK 11.2

DESCRIBING

- 1 **Define:**
 - a dipole
 - polar covalent bond
 - non-polar covalent bond.

APPLYING

- 2 The most electronegative elements, in order of decreasing strength are:
F > O > Cl > N = Br > I > C = S > P = H > Si
Consider the following pairs of elements:
O—H, Cl—H, C—H, P—H, N—O, N—Cl, N—H, S—H, C—S, F—O, Cl—O, C—O,
P—O, C—Cl
 - a With the covalent bonds between the above elements, **classify** each element as slightly more negative or slightly more positive by labelling with δ^- and δ^+ as appropriate.
 - b **Determine** the electronegativity difference for each pair of elements and use the values to **classify** each pair as strongly polar, slightly polar or non-polar.



3 Determine whether the following are polar molecules.

- a Iodine chloride
- b Dichloromethane (CH_2Cl_2)
- c Nitrogen trichloride
- d Sulfur difluoride
- e Boron trihydride



Weblink
Intramolecular and
intermolecular forces

11.3

Observable properties and intermolecular forces

An understanding of why some molecules are polar while others are not is crucially important when explaining the observable properties of substances.

These properties include:

- melting point
- boiling point
- solubility in polar and non-polar solvents
- vapour pressure.

The degree of polarity of molecules determines the intermolecular forces within substances.

Intermolecular forces

intermolecular force
an attractive force that causes molecules to aggregate to form liquids and solids

Intermolecular forces occur between molecules and are the forces that cause molecules to attract each other. They can occur between molecules of the same or different substances. Intermolecular forces should not be confused with intramolecular forces. Intramolecular forces are covalent bonds and are much stronger than intermolecular forces (**Figure 11.3.1**). VSEPR theory provides a basis for predicting the shapes of molecules based on the repulsive forces between electron pairs, both bonding pairs and non-bonding pairs. These are intramolecular forces. However, VSEPR theory also permits the understanding of the distribution of electrons within a molecule, which allows estimation of the molecular polarity. The polarity of a molecule in turn affects the strength of intermolecular forces.

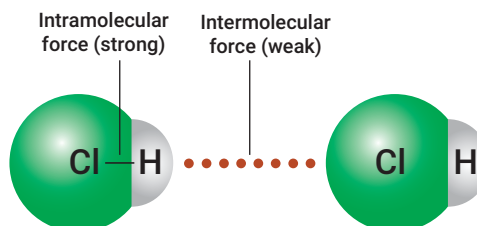


FIGURE 11.3.1 Intramolecular forces keep a molecule intact. Intermolecular forces hold multiple molecules together and determine many of a substance's properties.

Three types of intermolecular force, in order of increasing strength, are:

- dispersion forces
- dipole–dipole attractions
- hydrogen bonding.

Dispersion forces

Polar molecules can form dipole–dipole or hydrogen bonds between the molecules. Non-polar molecules (e.g. dry ice, liquid nitrogen and oils and fats) cannot, so there must be another force that enables non-polar molecules to form liquids and solids. This force is a **dispersion force**.

The simplest molecule, hydrogen (H_2), can be used to explain a dispersion force (Figure 11.3.2). The force is very weak and thus hydrogen has an extremely low melting point (about -259.2°C). Hydrogen molecules have a cloud of electron density surrounding them. When the electrons, which can be thought of as constantly ‘orbiting’ the molecule, are on one side of the molecule there is, for an instant, a temporary dipole. One end of the molecule is very slightly more positive and the other end is very slightly more negative. The temporary positive end attracts a neighbouring molecule’s electrons, inducing them to come towards the positive area. This forms the weak electrostatic attractions of dispersion forces. The fluctuating changes in charge result in a very weak force between molecules.

Dispersion forces exist in every molecule because every molecule has rapidly fluctuating electrons orbiting the molecule. The force is due to the rapid movement of electrons in the cloud of electron density surrounding the molecule, which results in the formation of instantaneous temporary dipoles.

Dispersion forces are much weaker than other intermolecular forces. When dipole–dipole or hydrogen bonding is present, the effect of dispersion forces is discounted unless the molecule is considerably larger.

dispersion force a weak intermolecular force that arises from electrostatic attractions between instantaneous dipoles in neighbouring molecules

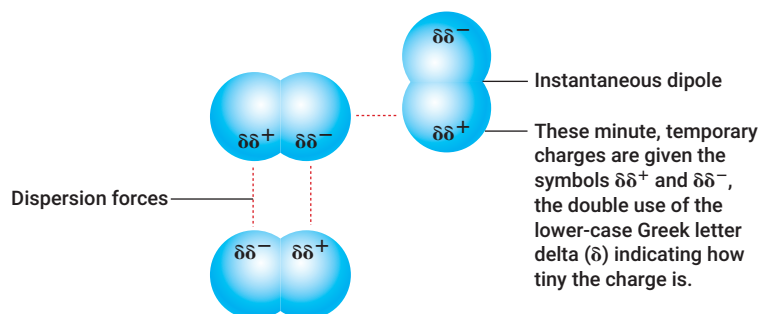


FIGURE 11.3.2 The dispersion force between hydrogen molecules in solid hydrogen, showing the instantaneous dipole–dipole attractions

Dispersion forces are also present in the noble gases (Figure 11.3.3). Electrons in two adjacent atoms occupy positions so that the atoms form temporary dipoles. With more electrons, there are more chances for dispersion forces to occur. Molecules are electrically neutral, so when there are more electrons, there are more protons. Increasing molecular weight of molecules means an increased chance of dispersion forces. An increase in dispersion forces means more energy is required to break these interactions between molecules. This results in a higher boiling point (and melting point) (Table 11.3.1).

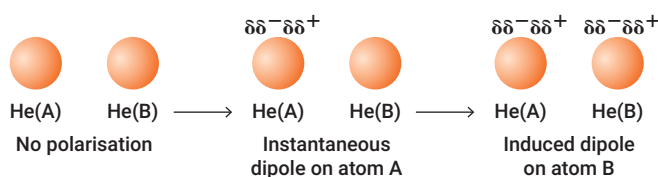


FIGURE 11.3.3 Noble gases can explain how dispersion forces occur through instantaneous dipoles, induced due to the movement of electrons.

The electrons in larger atoms are also further from the nucleus and more readily form temporary dipoles. This explains why the larger noble gas argon has a boiling point of -186°C , much higher than that of helium, -269°C . A similar trend of increasing dispersion forces with increasing molecular weight is seen with the halogens (Table 11.3.1). The more electrons present (the larger the molecular weight), the bigger the dispersion force between the molecules.

TABLE 11.3.1 Boiling points of noble gases and halogens

Noble gas	Boiling point ($^{\circ}\text{C}$)	Halogen	Boiling point ($^{\circ}\text{C}$)
Helium	-269	Fluorine	-188
Neon	-246	Chlorine	-34
Argon	-186	Bromine	59
Krypton	-152	Iodine	184
Xenon	-108	Astatine	335
Radon	-62		

Dipole–dipole attractions

In polar molecules, there is a net dipole where one end of the molecule is slightly negative and the other end is slightly positive. The dipole–dipole attraction occurs when a permanent dipole is present. Hydrochloric acid (HCl) is a polar molecule. The presence of dipoles results in an attraction between neighbouring molecules, a **dipole–dipole attraction**, which is an attraction between two dipoles in two individual molecules. Molecules line up so that the slightly positive (δ^{+}) end of one molecule attracts the slightly negative end (δ^{-}) of another molecule (Figure 11.3.4). The dipole–dipole attraction is relatively strong because the dipole is permanent.

dipole–dipole attraction
the attraction between molecules with permanent dipoles



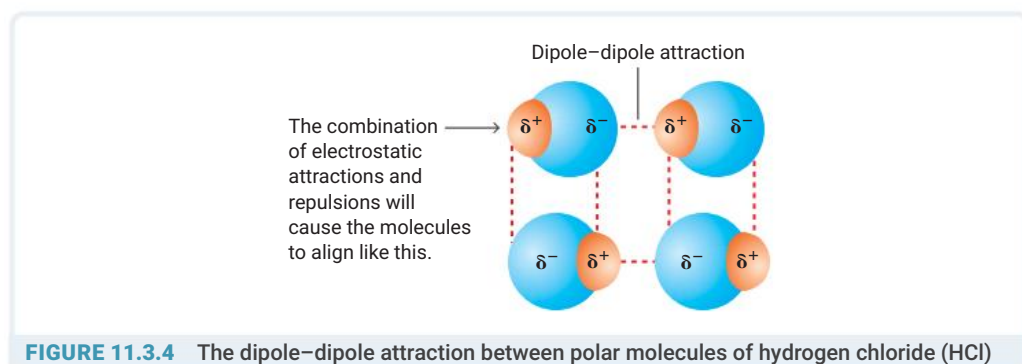
Worksheet

Intermolecular forces

Weblinks

Animation to show the formation of intermolecular forces

Intermolecular forces



Hydrogen bonds

Boiling points and melting points give an estimation of the strength of the intermolecular forces between molecules. Figure 11.3.5 shows a general increase in boiling points of hydrides as you go down the periodic table groups, if you could exclude period 2 hydrides.

When nitrogen, oxygen or fluorine is attached to a hydrogen atom, the resulting molecule has an unusually high boiling point. When these very electronegative elements bond to hydrogen, a special type of dipole–dipole attraction occurs called a **hydrogen bond**. This is the strongest of the intermolecular forces, being about one-tenth the strength of a covalent bond. In diagrams, it is shown as a dotted line between the molecules.

Hydrogen bonds are due to the hydrogen nucleus (a single proton) being extremely small. When hydrogen is bonded to the most electronegative elements (fluorine, oxygen and nitrogen),

hydrogen bond when a H atom that is bonded to an O, N or F atom in one molecule is attracted to the lone pair of electrons of N, O or F of an adjacent molecule

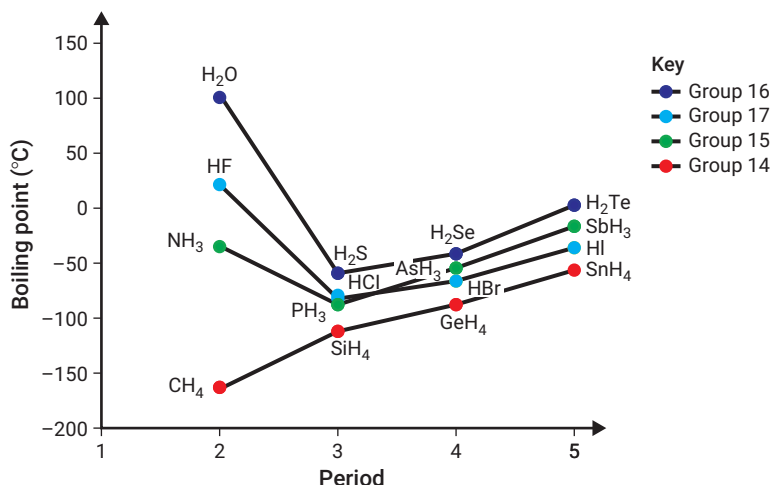


FIGURE 11.3.5 The boiling points of hydrides of groups 14–17

the charge difference over the polar covalent bond is at a maximum. Water displays hydrogen bonding. Oxygen's strong ability to attract electrons means that the electrons in the O—H covalent bond are more attracted to oxygen. Oxygen becomes slightly negatively charged and hydrogen is left with a slightly positive charge. The hydrogen atom is so small that it also means that two adjacent water molecules can get very close to each other. Similarly, in ammonia, the nitrogen attracts the electrons of the covalent bonds, leaving the hydrogen atoms more positive. These hydrogens can attract the nitrogen atoms of adjacent molecules. Hydrogen bonding then occurs between the hydrogen in —OH, —NH or —FH and the lone pairs of nitrogen, oxygen or fluorine atoms on the adjacent molecule (Figure 11.3.6).

In summary, the relative strengths of intermolecular forces in small molecules are approximately: H bond $10\times$ > dipole–dipole attractions $10\times$ > dispersion forces alone. On this relative scale, a covalent bond would be 10 times stronger than a hydrogen bond.

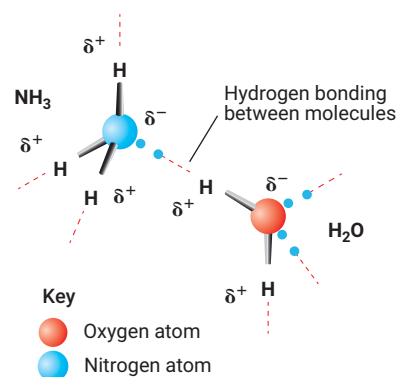


FIGURE 11.3.6 Hydrogen bonding occurring between a hydrogen (H) on one molecule and the lone pair of electrons on nitrogen (N) in another molecule

The importance of bonding

Bonding is responsible for many of the properties of materials and biomolecules. The nucleic acids in the DNA double helix are held together by hydrogen bonds. This allows them to unzip easily, enabling replication and transcription. Hydrogen bonding helps maintain the structures of proteins, enzymes, ribosomes and cell membranes.

Analytical techniques such as chromatography work because of surface interactions and the intermolecular bonds of different substances to different surfaces or solvents. Chemists use this knowledge to determine the organic substances in mixtures.



Syllabus link
Chapter 12 discusses chromatography and the role of intermolecular forces in more detail.

Intermolecular forces and melting and boiling points

Recall looking at melting points and boiling points in Chapter 7. The melting point is the temperature at which the substance melts and the boiling point is the temperature at which the substance boils. It was found that covalent molecular substances had low melting points and boiling points, whereas covalent network substances had very high melting points and boiling points owing to the lattice structure. Covalent molecular structures have a large variety of melting

points and boiling points because of the various sizes, polarities and types of intermolecular forces present between molecules. Those substances that have dispersion forces only have low melting points and boiling points, whereas substances containing dipole–dipole attractions have higher melting points and boiling points. Substances that contain hydrogen bonds have the highest melting points and boiling points because of this strong force between certain molecules that contain a hydrogen atom bonded to either a nitrogen, oxygen or fluorine atom.

The size of a molecule also has an effect on the melting points and boiling points. If two non-polar molecules are compared, the molecule that has the larger molecular mass, and therefore more electrons, has higher melting and boiling points. This is due to more electrons being available to form more temporary dipoles between molecules. Worked examples 11.3.1 and 11.3.2 demonstrate how to predict which molecules have higher melting and boiling points.

WORKED EXAMPLE 11.3.1

Predict which would have higher melting and boiling points: CO_2 or CS_2 .

ANSWER

1 Compare the molecular masses of the molecules.

CS_2 is a much heavier molecule than CO_2 .

2 Determine the intermolecular forces between each molecule.

Both CO_2 and CS_2 are non-polar molecules, so only dispersion forces are present.

3 Determine which substance has higher melting and boiling points based on the strength of their intermolecular forces.

The dispersion forces are stronger in CS_2 than in CO_2 because it has a larger molecular mass. Therefore, there are more electrons to form instantaneous, temporary dipole–dipole attractions. As a result, CS_2 has the higher melting and boiling points.

WORKED EXAMPLE 11.3.2

Predict which would have higher melting and boiling points: CH_4 or CH_3Cl .

ANSWER

1 Compare the molecular masses of the molecules.

CH_3Cl is a much heavier molecule than CH_4 .

2 Determine the intermolecular forces between each molecule.

CH_4 is a non-polar molecule so only dispersion forces are present. CH_3Cl is a polar molecule because of the presence of the more electronegative chlorine atom, and therefore has both dispersion forces and dipole–dipole attractions.

3 Determine which substance has higher melting and boiling points based on the strength of their intermolecular forces.

CH_3Cl has higher melting and boiling points because it has dipole–dipole attractions as well as dispersion forces, which are stronger than the dispersion forces in CH_4 .

Intermolecular forces and vapour pressure

The energy of particles in a liquid depends on its temperature. At a higher temperature, particles have more energy on average. However, some particles have more energy and some have less energy than the average. At the surface of the liquid, sometimes there are particles that have enough energy that they can move fast enough to break free from the forces of attraction holding

the particles to other liquid particles. This process is called evaporation, which is different from boiling. When a liquid boils, there is enough energy to break the intermolecular bonds throughout the whole liquid, not just at the surface.

In a closed container, the evaporated particles stay trapped inside the container. Over time, the number of particles leaving the liquid and returning to the liquid become balanced. As particles move around in the space above the liquid, they collide with the walls of the container and exert pressure. This pressure is called the **vapour pressure** of a liquid.

Different substances have different vapour pressures at a given temperature. This is because of the strength of the intermolecular bonds holding the molecules together in the liquid phase. In substances that have weaker intermolecular forces, more particles can escape the liquid, resulting in a higher vapour pressure. Substances that have a high vapour pressure are described as being volatile because they can easily evaporate. The opposite is true for substances with strong intermolecular forces, where fewer particles can enter the gas phase at the same temperature, resulting in a lower vapour pressure (**Figure 11.3.7**).

vapour pressure the pressure of a substance in the gas phase (vapour) that exists above a sample of liquid of the same substance in a closed container

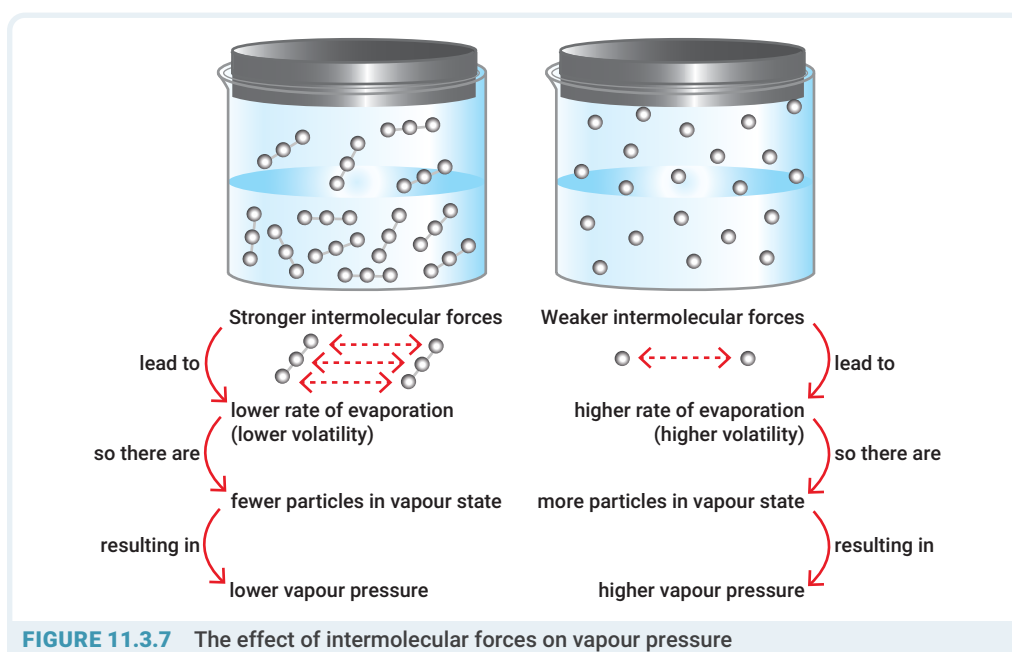


FIGURE 11.3.7 The effect of intermolecular forces on vapour pressure

WORKED EXAMPLE 11.3.3

Predict which would have the higher vapour pressure: HCl or NH_3 .

ANSWER

1 Determine the intermolecular forces between each molecule.

Both are polar molecules, so they have both dispersion forces and dipole–dipole attractions between their molecules. However, NH_3 also has hydrogen bonds due to the presence of a nitrogen atom attached to a hydrogen atom in each molecule.

2 Determine which substance contains the strongest intermolecular forces.

NH_3 has hydrogen bonds, which are stronger than dipole–dipole attractions.

3 Determine which substance has the higher vapour pressure.

HCl has the higher vapour pressure because it contains the weakest intermolecular forces.

**Syllabus link**

Solubility will be explored in greater depth in Chapter 16.

Intermolecular forces and solubility

Solubility measures how well a solute can dissolve in a solvent. When considering the solubility of substances (solutes) in other substances a general rule applies:

Like dissolves like.

The 'like dissolves like' rule means that polar substances (including ionic compounds) dissolve in polar solvents, and non-polar substances dissolve in non-polar solvents.

Water, being a polar substance, can take part in hydrogen bonding or dipole–dipole attraction with polar solutes. For example, water can form hydrogen bonds with ammonia. Therefore, ammonia can be dissolved in water. Methane has dispersion forces only, so is unable to dissolve in water. However, methane can dissolve in a non-polar solvent such as hexane (C_6H_{12}).

WORKED EXAMPLE 11.3.4

Predict whether the following substances would be soluble in pentane (C_5H_{12}) or water: Cl_2 , HCl, NaCl.

ANSWER

1 Determine the intermolecular forces in each solute (substance to be dissolved).

Cl_2 is non-polar so only contains dispersion forces. HCl is polar and contains dispersion forces and dipole–dipole attractions. NaCl is an ionic compound.

2 Determine the intermolecular forces in each solvent (substance doing the dissolving).

Pentane has only dispersion forces so is non-polar. Water is polar and has dispersion forces, dipole–dipole attractions and hydrogen bonding.

3 Determine which solute dissolves in which solvent by using the 'like dissolves like' rule.

Cl_2 dissolves in pentane. Both HCl and NaCl will dissolve in water. NaCl is ionic so it dissolves in the polar water due to the positive and negative ions being attracted to the negative and positive ends of the water molecule.

LEARNING CHECK 11.3

DESCRIBING

- 1 List three strong bonding forces.
- 2 Identify intermolecular forces.
- 3 State the common groups that exhibit hydrogen bonds.

APPLYING

- 4 Rank the three intermolecular forces from weakest to strongest. Explain how polarity influences this order.
- 5 Explain the rule 'like dissolves like'.
- 6 Explain why dihydrogen sulfide (H_2S) has a lower boiling point than water.

ANALYSING

- 7 Use Figure 11.3.8 to explain the trend in boiling points in group 17 hydrides.

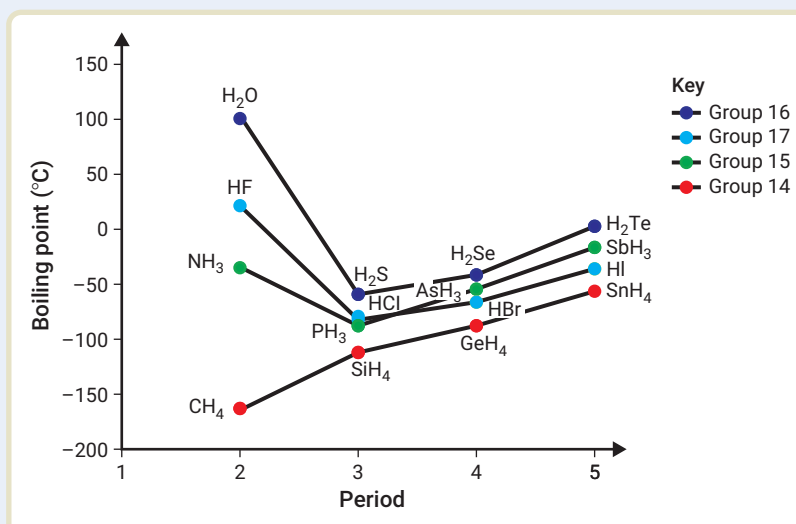


FIGURE 11.3.8 The boiling points of hydrides of groups 14–17

CHAPTER SUMMARY

Molecular shape and polarity

- Shapes of molecules can be predicted by using the valence shell electron pair repulsion (VSEPR) theory.
- Non-bonding pairs of electrons more strongly repel other pairs of electrons than bonding pairs. Bonding pairs repel each other the least.
- The shape and symmetry of a molecule determine the polarity of the molecule that contains polar bonds.
- Molecules that contain polar bonds but are symmetrical are overall non-polar molecules; for example, CCl_4 .

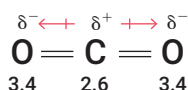
Identifying polarity of covalent compounds

Non-polar covalent compound



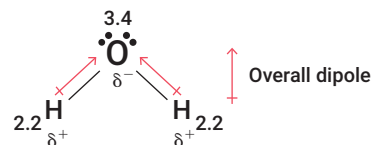
Contains only non-polar covalent bond and is therefore overall non-polar

Non-polar covalent compound



Contains polar covalent bonds; however, is symmetrical and therefore overall non-polar

Polar covalent compound

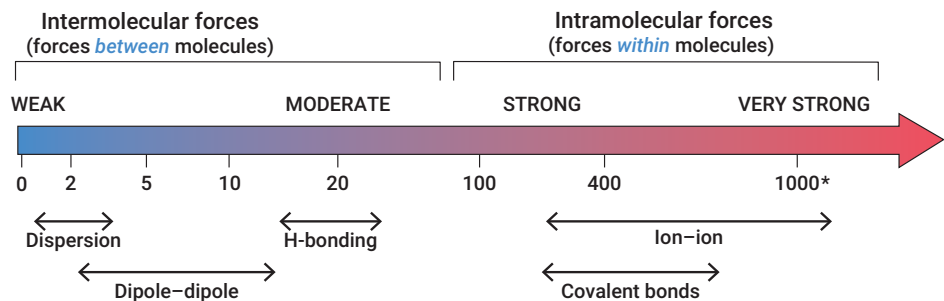


Contains polar covalent bonds and is asymmetrical and therefore overall polar

Intermolecular vs intramolecular forces

- Intramolecular forces occur within molecules, whereas intermolecular forces occur between molecules.
- Intramolecular forces (bonds) are much stronger than intermolecular bonds.

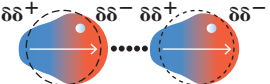
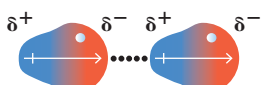
Comparison of strengths of intramolecular bonds and intermolecular forces



* Units are in kJ mol^{-1} of interactions

Intermolecular forces

- For covalent molecules, differences in electronegativity between the bonding atoms can result in polar covalent bonds.
- As a general rule, bonds where there is a difference in electronegativity of less than 0.4 are considered non-polar covalent bonds, whereas bonds that have a difference of 0.5–1.0 are considered slightly polar covalent bonds. Bonds that have a difference in electronegativity of 1.0–1.7 are considered to be strongly polar covalent bonds.
- The polarity of molecules affects the types of intermolecular bonds that can form.
- There are three different types of intermolecular forces that help determine the properties of substances: dispersion forces, dipole–dipole attractions and hydrogen bonding.
- Intermolecular forces determine vapour pressure, melting point and boiling point and solubility of substances.
- The ‘like dissolves like’ rule helps to determine whether a substance will dissolve in a particular solvent. Polar substances will generally dissolve in polar solvents; non-polar substances will generally dissolve in non-polar solvents.

Intermolecular force	Present in	Diagram
Dispersion forces	All molecules and atoms	
Dipole–dipole attractions	Polar molecules	
Hydrogen bonding	Molecules contain an H bonded to F, O or N	

Increasing strength

CHAPTER EXAM

MULTIPLE CHOICE

- Choose the type of geometry in the water molecule.
A Bent
B Linear
C Pyramidal
D Trigonal planar
- The central atom in NF_3 is surrounded by:
A three single bonds only.
B three single bonds and a non-bonding pair of electrons.
C three single bonds and two non-bonding pairs of electrons.
D a double bond, a single bond and a non-bonding pair of electrons.
- The boiling point of water is much higher than would be predicted from the boiling points of the other group 16 hydrides such as H_2S , H_2Se and H_2Te . Which of the following statements explains this?
A Water has a lower molecular mass.
B Water contains hydrogen bonding.
C Water is less polar than H_2S , H_2Se and H_2Te .
D The covalent bonds in water are stronger than in H_2S , H_2Se and H_2Te .
- In which molecule is there at least one polar bond but no overall dipole?
A CF_4
B SO_2
C CHCl_3
D NH_3
- What type of intermolecular forces are due to the attraction between temporary dipoles and their induced temporary dipoles?
A Covalent bonds
B Dispersion forces
C Hydrogen bonds
D Ionic bonds
- Which of the following boils at the *highest* temperature at the same atmospheric pressure?
A CH_4
B C_2H_6
C C_3H_8
D C_4H_{10}
- Which has the *lowest* boiling point at the same atmospheric pressure?
A HF
B HCl
C HBr
D HI
- What is the molecular shape if the central atom is from group 13 and only contains single bonds?
A Bent
B Linear
C Trigonal planar
D Trigonal pyramidal

9. Which one of the following substances is more likely to dissolve in CCl_4 ?

- A CBr_4
- B HBr
- C HCl
- D NaCl

10. What types of intermolecular forces exist between HF and H_2O ?

- A Dispersion forces and hydrogen bonding
- B Dispersion forces and dipole–dipole attractions
- C Dispersion forces, dipole–dipole attractions, and hydrogen bonding
- D Hydrogen bonding and dipole–dipole attractions

SHORT RESPONSE

11. From the following list of elements, select the pair(s) that would form the strongest dipole:

Cl , S , C , H , F , N , Ar , Br and P

12. **Explain** why the boiling point of CH_4 (-161.5°C) is lower than that of CCl_4 (76.7°C).

CROSS-CHAPTER QUESTION

13. Consider the boiling points of the following compounds:

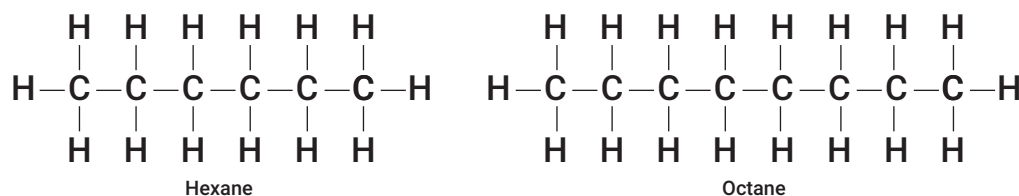
- Silicon tetrafluoride (SF_4), 40°C
- Sodium chloride (NaCl), 1465°C

- a **Identify** the types of compounds SF_4 and NaCl are.
- b Using the information presented, **explain** why there are differences between the boiling points of the compounds.

DATA ANALYSIS

14. **Analyse data**

Two substances were mislabelled in the laboratory; however, it was known that one of them was octane and the other was hexane. The structures of the two substances are shown below.



To identify each substance, a student conducted an experiment to determine their boiling points. The results are shown below.

	Unknown substance A	Unknown substance B
Boiling point ($^\circ\text{C}$)	126	69

- a Based on the information above, **identify** substance A and substance B.
- b **Justify** your answer for part a.

15. Analyse data

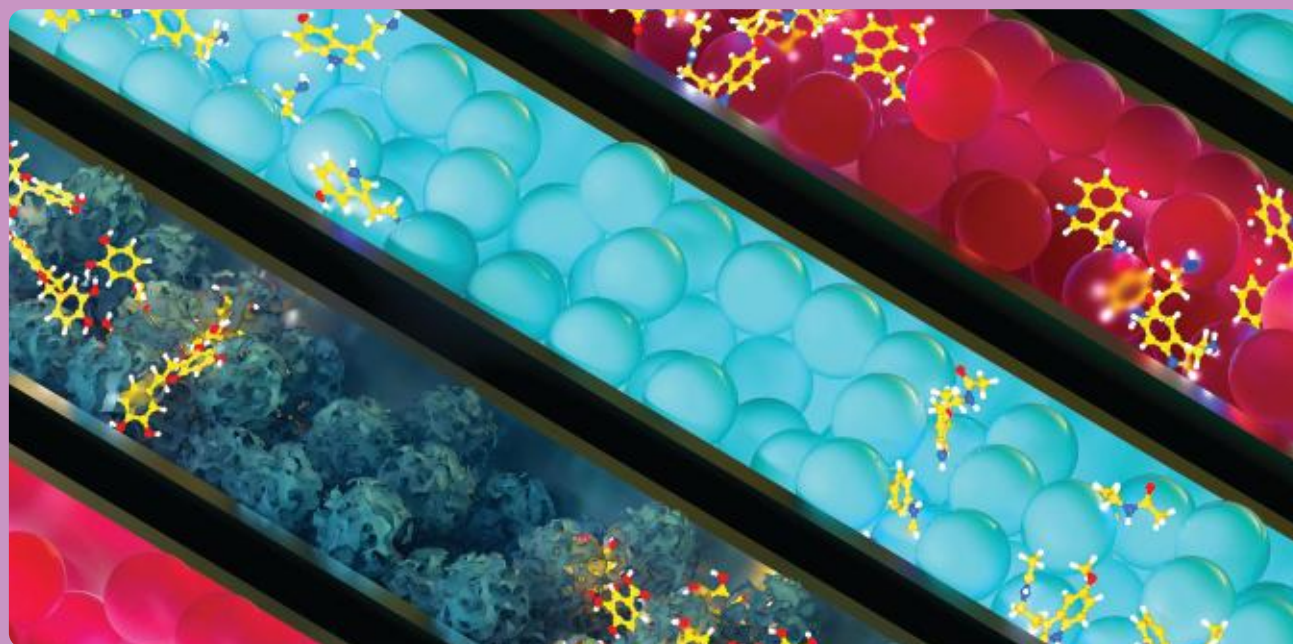
The following table provides information about three organic compounds, X, Y and Z.

Compound	Structural formula	Molar mass (g mol^{-1})	Boiling point ($^{\circ}\text{C}$)
X	$ \begin{array}{ccccccc} & \text{H} & & \text{H} & & \text{H} & \\ & & & & & & \\ \text{H} & - \text{C} & - & \text{C} & - & \text{C} & - \text{O} - \text{H} \\ & & & & & & \\ & \text{H} & & \text{H} & & \text{H} & \end{array} $	60	97
Y	$ \begin{array}{ccc} & \text{H} & & \text{O} \\ & & & // \\ \text{H} & - \text{C} & - & \text{C} \\ & & & \backslash \\ & \text{H} & & \text{O} - \text{H} \end{array} $	60	118
Z	$ \begin{array}{ccc} & \text{H} & & \text{O} \\ & & & // \\ \text{H} & - \text{C} & - \text{O} - & \text{C} - \text{H} \\ & & & \\ & \text{H} & & \end{array} $	60	?

- Identify** the intermolecular forces in compounds X, Y and Z.
- Compare** the structures of compounds X and Y to account for the difference in their boiling points.
- Predict** the boiling point of compound Z based on its chemical structure. Provide reasons for your response.

CHAPTER 12

Chromatography



Science Photo Library/Alamy Stock Photo

SYLLABUS DOT POINTS

SCIENCE UNDERSTANDING

- Identify that paper and thin layer chromatography 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 (TLC) chromatographs to determine the composition and purity of substances, including calculating R_f values.

SCIENCE INQUIRY

Investigate:

- the separation of a mixture using paper or thin layer chromatography* (TLC).

*Note: Simulations may be used.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Chromatography is one of the most powerful and, in many ways, one of the simplest analytical tools at our disposal. It is used to separate mixtures, identify unknown substances and provide information on amounts of substances present in a mixture.

The word 'chromatography' is derived from the Greek words for 'colour', *khroma*, and 'to write', *graphein*. The term was first used by Russian scientist Mikhail Tsvet in 1900 in his work on separating plant pigments such as chlorophyll, carotenes and xanthophylls, which are coloured.

Many new forms of chromatography have been developed, and with technological improvements in recent decades, these new techniques are becoming indispensable to the analytical chemist.

Practicals

- Separating a mixture of dyes
- Chromatography of plant pigments (online-only resource)

Worksheets

- Summarising chromatography methods



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ The shape and polarity of bonds in a molecule determine if the molecule is polar or non-polar.
- ✓ Types of intermolecular forces are dispersion forces, dipole–dipole attractions and hydrogen bonding.
- ✓ The strength of intermolecular forces is in the order: hydrogen bonding > dipole–dipole attractions > dispersion forces.
- ✓ Mixtures can be separated into their component elements and/or compounds by physical means.
- ✓ A solution is made up of a solute dissolved in a solvent.

LEARNING OBJECTIVES

By the end of this chapter, you should be able to:

- ✓ describe how chromatography can be used to determine the composition of a mixture and purity of a substance
- ✓ identify that chromatography has a stationary phase and mobile phase, which may be the same or different states
- ✓ describe how components of a mixture adsorb and desorb from the stationary phase into the mobile phase
- ✓ identify the differences in the stationary phase of paper and thin-layer chromatography
- ✓ determine the retardation factor (R_f) and use it to compare and identify mixture components
- ✓ explain how the distance moved by a component is related to intermolecular attractions between the component, and the stationary and mobile phases
- ✓ analyse paper and thin-layer chromatograms to determine mixture composition and purity of components

chromatography a group of techniques that separate substances based upon differential distribution between a stationary phase and a mobile phase

stationary phase the immobile surface that different components of a mixture can adhere to when carried along by the mobile phase

mobile phase the fluid used to carry the components of a mixture up the stationary phase so they can separate

adsorb to be attracted to the surface of a material; the opposite of desorb

desorb the action of a substance moving from the stationary phase to the mobile phase

12.1 Chromatography techniques

Chromatography is an important analytical technique used to separate the components of a mixture, allowing the components of the mixture to be identified. It may also be used to determine the purity of a substance by indicating the presence or absence of impurities.

Knowledge of how solutes dissolve in solvents and interact with the surface of materials is the basis of this analytical technique. There are several techniques that can be used, including paper chromatography, thin-layer chromatography, gas chromatography and high-performance liquid chromatography. Paper chromatography and thin-layer chromatography (TLC) require relatively simple equipment, whereas gas chromatography and high-performance liquid chromatography require sophisticated equipment. However, the chemical basis for all these techniques is that different substances are attracted to and **adsorb** onto a surface (called the **stationary phase**) and detach (**desorb**) into a solvent (the **mobile phase**) at different rates. In fact, the component molecules may attach to the stationary phase, then dissolve in the solvent and travel a little further, then attach to the stationary phase again, and so on (**Figure 12.1.1**).

Paper and thin-layer chromatography

Paper and thin-layer chromatography allow us to determine how many components are present in the mixture. We can also identify the components if the data for the particular solvent system is available.

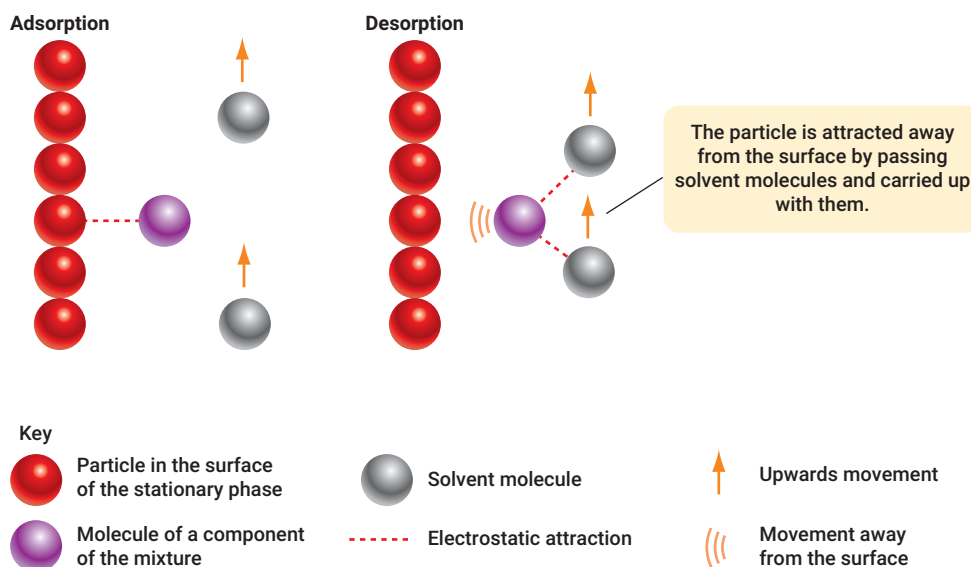


FIGURE 12.1.1 Absorption and desorption

Separation of the components is achieved by the solvent (mobile phase) moving up the stationary phase. This movement is due to capillary action. As the solvent moves through the stationary phase towards the top, a faint wavy line, called the solvent front, can be seen. The components travel with the solvent. If all the components spent the same amount of time in the solvent, then these components would all travel at the same rate. However, different components have different chemical structures, which alter the adsorption and desorption rates. The molecules most attracted to the stationary phase move the slowest. The molecules that are most strongly attracted to the solvent move the fastest. The process is outlined in **Figure 12.1.2**.



Weblinks
[Chromatography](#)
[What is paper chromatography?](#)
[Animation of paper chromatography](#)

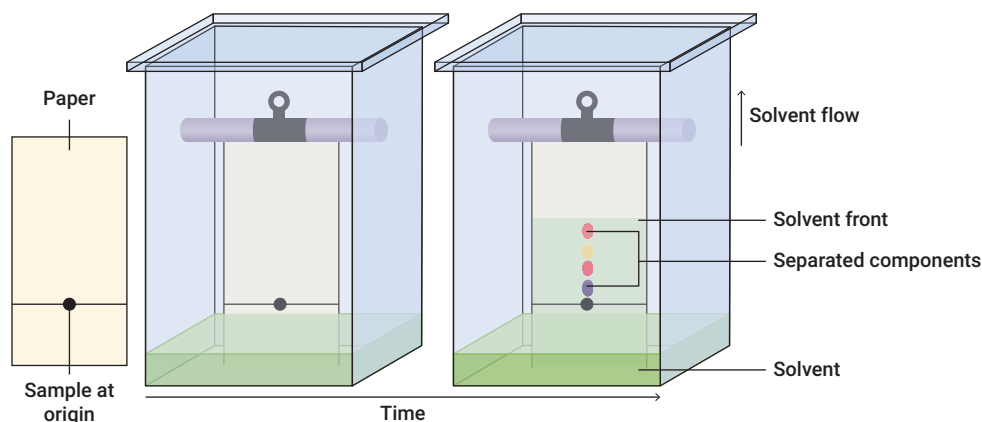


FIGURE 12.1.2 The separation of a sample in paper chromatography

In both paper and thin-layer chromatography, you measure the distance the individual component moves from the origin compared to the distance the solvent moves from the origin. This ratio is called the **retardation factor (R_f)** (also called ‘retention factor’), as shown in **Figure 12.1.3**.

retardation factor (R_f)
 the distance travelled by the sample from its origin divided by the distance moved by the solvent from the origin

The R_f value must be between 0 and 1 and is defined as the distance travelled by the sample components from the origin divided by the distance of the solvent front from the sample origin.

$$R_f = \frac{\text{distance moved by component}}{\text{distance moved by solvent front}}$$

A substance that does not migrate from the sample origin has an $R_f = 0$, whereas one that is not adsorbed (i.e. has migrated with the front) has an $R_f = 1$. The R_f value is characteristic for a particular substance within a solvent system. The R_f value cannot be greater than 1.

For example, in an experiment, if the distance travelled by the solvent is 3.4 cm and the distance travelled by the sample is 1.9 cm, then the R_f value would be:

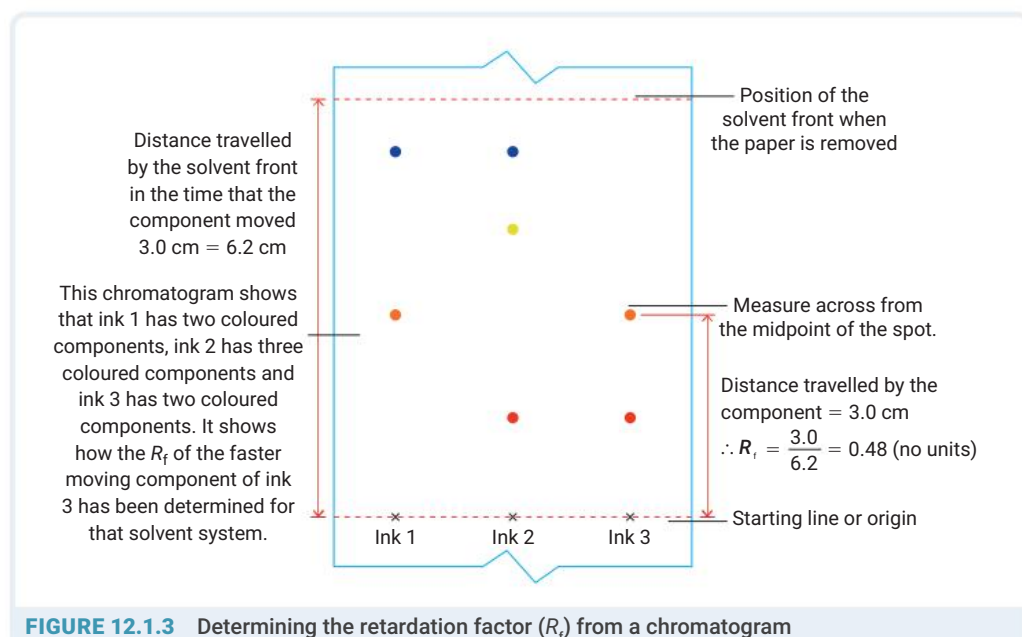
$$R_f = \frac{1.9}{3.4} = 0.5588$$

An R_f value has no units and is usually quoted to 2 decimal places, so the R_f value calculated above should be:

$$R_f = 0.56$$

As long as the conditions under which this experiment is carried out remain constant, such as temperature, solvent and type of stationary phase, the R_f value will be the same.

After the stationary phase has dried of solvent, the result is called a chromatogram (Figure 12.1.3).



The degree at which the components can separate partly depends on the length of time the stationary phase is in contact with the solvent. If it is too short, then there is only a small separation. If it is too long, then all the components eventually make it to the top. A longer plate and a reasonable time provide the best separation. It is not always possible to identify all the components of a mixture with one experiment. Some components may have very similar R_f values. Different solvent mixtures and stationary phases can be tried to achieve better separation.

WORKED EXAMPLE 12.1.1

A painkiller medication is tested to see whether it contains aspirin, paracetamol or ibuprofen. The painkiller is dissolved in ethyl ethanoate (ethyl acetate) and placed on a TLC plate (stationary phase), marked 'U' (unknown). Pure samples of aspirin, paracetamol and ibuprofen are placed on the same plate. Once the solvent front reaches the top of the plate, the plate is removed and dried. An ultraviolet light is shone onto the plate and the spots are marked with iodine solution stain. Figure 12.1.4 shows the results.

Does the unknown painkiller contain aspirin, paracetamol or ibuprofen?

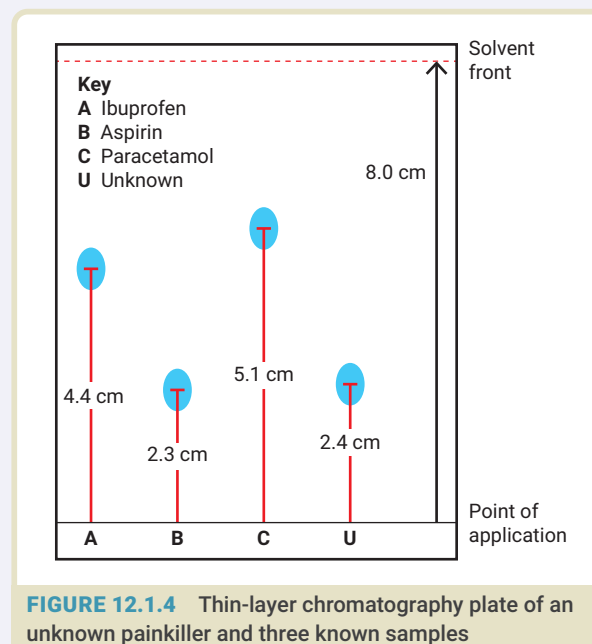


FIGURE 12.1.4 Thin-layer chromatography plate of an unknown painkiller and three known samples

ANSWER

- 1 Calculate the retardation factor (R_f) values for each component, as follows.

$$R_f A = \frac{4.4}{8} = 0.55$$

$$R_f B = \frac{2.3}{8} = 0.29$$

$$R_f C = \frac{5.1}{8} = 0.64$$

$$R_f U = \frac{2.4}{8} = 0.30$$

- 2 Compare the R_f values for the unknown painkiller with the known samples to determine the most likely constituent of the unknown painkiller.

Component B (aspirin) is the most likely constituent with an R_f value of 0.29 compared to a value of 0.30 for the unknown.

Investigating paper chromatography

Paper chromatography is the simplest form of chromatography to perform, but it is still useful in the qualitative analysis of components in a mixture. If the R_f values are available for the system being investigated, then the components can be identified.

In this method, the stationary phase is water molecules held inside the cellulose network of a piece of absorbent paper, and the mobile phase is a solvent in which the mixture being analysed will dissolve. All components of the mixture must be able to dissolve in the solvent, so sometimes a mix of different solvents is required to achieve this.



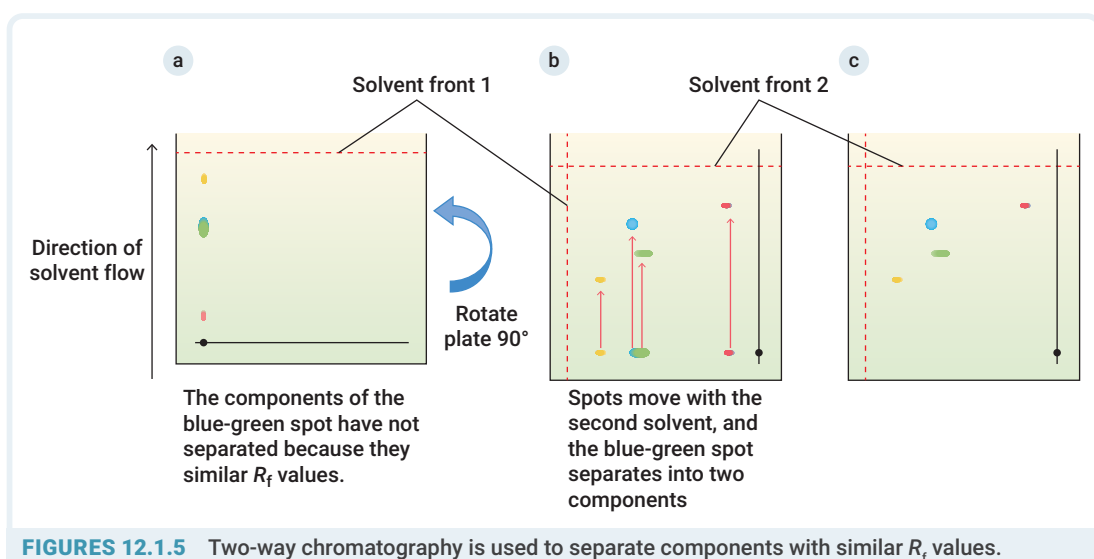
Practical
Chromatography of
plant pigments

Separation occurs because of differences in the relative solubility of each component in the solvent and differences in attraction to the water molecules of the stationary phase.

Sometimes it is difficult to separate components that have similar R_f values. When this is the case, a second separation using a different solvent with the original chromatogram is conducted. This process is called two-way paper chromatography.

The steps of two-way chromatography are shown in **Figure 12.1.5**.

- First, a chromatogram separates a mixture of dyes using the first solvent. In **Figure 12.1.5a**, you can see that the second spot from the bottom appears to be partly blue and partly green. The position of the solvent front is shown.
- The paper is allowed to completely dry and then rotated 90° and placed in a different solvent. **Figure 12.1.5b** shows how the dyes have moved in the second solvent and the position of the second solvent front. The blue-green spot won't be seen in the final chromatogram because it has separated into two distinct spots (**Figure 12.1.5c**).



Investigating thin-layer chromatography

Thin-layer chromatography (TLC) is used to separate and analyse a wide variety of molecular mixtures. It is used to analyse for the presence of particular drugs and amino acids.

TLC has several advantages over paper chromatography:

- It is faster (typically taking around 20 minutes).
- Better separations are achieved.
- It is more sensitive so smaller quantities can be used.
- A wider range of reagents can be used.

The TLC plate typically consists of a 0.1 mm thick layer of absorbent material bonded to a glass or plastic support. The absorbent material consists of a specially prepared, finely ground matrix of microscopic particles of silica gel, alumina or similar materials. These particles provide a large surface area for chromatographic separation. This is the stationary phase. The origin line is drawn in pencil about 1 cm from the base, which is also what is done in paper chromatography. Ink and pen must not be used because they contain a mixture of dyes that may separate and contaminate the sample being analysed. This is where the samples are placed.

PRACTICAL ACTIVITY 12.1.1

SEPARATING A MIXTURE OF DYES

Research question

Which solvent is more effective at separating components of food dyes?

Aim

To use thin-layer chromatography to separate mixtures of dyes to determine their composition

Materials

- 5 coloured food dyes
- methylated spirits
- distilled water
- 2 TLC plates (5 cm × 10 cm)
- 2 × 250 mL beakers
- 5 capillary tubes
- paper towel



What are the risks in doing this investigation?

Methylated spirits is a toxic and flammable liquid.

Glassware

How can you manage these risks to stay safe?

Keep bottles away from flames and heat sources. Do not come into contact with the liquid or breathe the fumes.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

- 1 On both TLC plates use a pencil (not ink) to draw a line across the plate approximately 1 cm from the bottom. Do not press too hard as this will scrape the stationary phase layer off.
- 2 Using a capillary tube, place a small drop of dye solution on the pencil line. Allow the dye to be adsorbed onto the film.
- 3 Repeat this for the second plate.
- 4 Repeat steps 2 and 3 for different dyes. Use a new capillary tube for each dye.
- 5 Pour distilled water into one of the beakers to a depth of approximately 5 mm.
- 6 Place the bottom edge of one plate in the beaker. Check the water level is below the pencil line.
- 7 Repeat steps 5 and 6 using methylated spirits for the other plate.
- 8 Allow the plates to stand in the solvents for approximately 10 minutes.
- 9 Withdraw the plates, place them on a piece of paper towel and use a pencil to mark the solvent front.
- 10 Mark the bottom of each dye component with a pencil.
- 11 Let the plates dry, then measure the distance moved by the solvent and each component of the dyes.
- 12 Record your measurements in a table for each of the dyes used.

Results

- 1 Record the distance moved by each solvent front.
- 2 Copy and complete the following results table.

Dye	Distance for distilled water (cm)	R_f	Distance for methylated spirits (cm)	R_f

Analysis of results

- 1 Calculate the R_f values for each of the dyes and their components and add them to your results table.
- 2 Compare the R_f values of the dyes in the different solvents.

Interpretation

- 3 Compare your results with those of other members of the class. Did you get similar R_f values for the same dyes? Suggest why or why not.
- 4 Identify if different dyes had any common components. Explain how you decided this.
- 5 Which dye was adsorbed most strongly in:
 - a water?
 - b methylated spirits?
- 6 Identify which solvent was the best for separating dyes. Justify your choice.

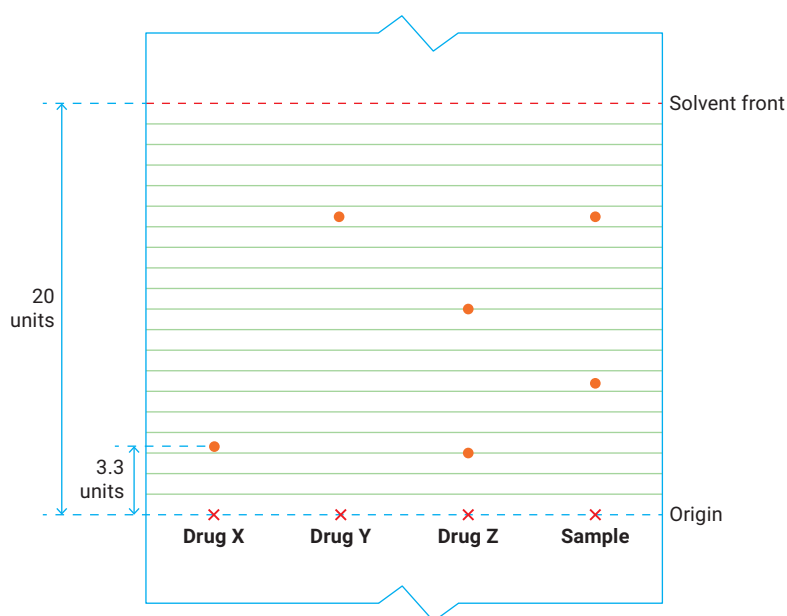
LEARNING CHECK 12.1

DESCRIBING

- 1 **Explain** the purpose of the stationary and mobile phases in chromatography.
- 2 **Identify** what paper and thin-layer chromatography can be used to determine.
- 3 **Compare** paper and thin-layer chromatography.

APPLYING

- 4 **Explain** why TLC would not be suitable for separating a large quantity of mixture.
- 5 **Figure 12.1.6** represents a thin-layer chromatogram of drug samples confiscated by the police. Included on the plate are commercial drugs, labelled X, Y and Z.

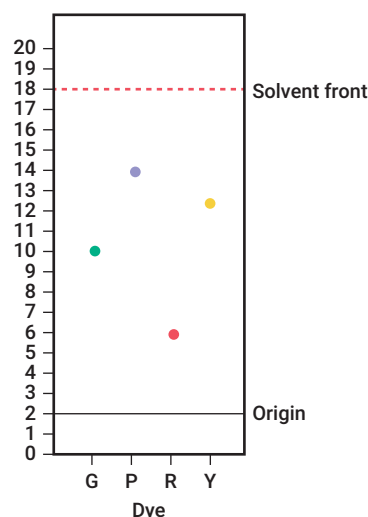


FIGURES 12.1.6 A thin-layer chromatogram of drug samples

- a Explain** why chemists include commercial drugs on the same TLC plate as the samples from the police.
- b Determine** if the confiscated sample contains any of the commercial drugs. **Justify** your answer using the data provided.
- 6 A paper chromatogram of four dyes, G, P, R and Y, is shown in **Figure 12.1.7**.

Calculate the R_f value for each of the dyes.

- 7 A student's note book was messy and it was hard to read the practical instructions for separating the individual dye components of food colouring. Steps 1–10 of the instructions are shown below. Not only was the order strange but there were errors.
- a Identify** the correct order for the instruction steps 1–10.
- b Identify** three fundamental errors in the instructions and in each case **explain** how the error should be corrected.
- 1 Identify the components of the food colouring by comparing the R_f value with those of known dyes.
 - 2 Put more drops of the food colouring over the first to make a concentrated spot that is at least 15 mm in diameter so that it has adequate contact with the paper and can be easily seen.
 - 3 Allow the drop to dry.
 - 4 Carefully place a drop of food colouring on the line.
 - 5 Place a solvent in the bottom of a tall jar.
 - 6 Hang the strip of chromatography paper in the jar, making sure that the spot of colouring is completely covered by the solvent.
 - 7 Allow the paper to remain in contact with the solvent until a solvent front has almost reached the top of the paper.
 - 8 Observe the separation of the coloured components of the food colouring. Note that the component that adsorbs most strongly to the paper pulls itself furthest up the paper.
 - 9 Calculate an R_f value for each spot on the chromatogram.
 - 10 Using a gel pen, draw a reference line 2 cm from the bottom of a strip of chromatography paper.



FIGURES 12.1.7 A paper chromatogram of dye samples

12.2 Separating components

Separating the components of mixtures depends largely on differences in the strength of the intermolecular forces (dipole–dipole attractions, dispersion forces, hydrogen bonding) between the component molecules, the mobile phase and the stationary phase.

Different components in a mixture adsorb onto a surface and desorb into a solvent at different rates. This difference is due to many factors of the substance being tested, including:

- the different types of polar groups
- the amount of charged and polar chemical groups present
- its molecular weight
- its geometry
- the positions and numbers of carbon–carbon double bonds.

How fast the components move up the plate depends on the strength of attraction between the molecules of the component and the:

- solvent molecules
- stationary phase.

The more strongly a component attaches to the stationary phase, the more slowly it moves in the direction of flow of the mobile phase. Components that attach only weakly to the stationary phase but dissolve readily in the mobile phase move up the stationary phase more quickly. Molecules of higher molecular weight with fewer polar sites tend to travel further up the paper, despite the downward pull of gravity.

To understand the role of intermolecular forces in chromatography, it is necessary to look at the different types of chromatography. Paper chromatography is easy to perform, but the chemistry of the stationary phase is more complex than in thin-layer chromatography. Therefore, we will examine thin-layer chromatography because the concepts in this technique also apply to paper chromatography.

Thin-layer chromatography

In thin-layer chromatography (TLC), the most common material used for the stationary phase is silica (a form of silicon dioxide). A very thin layer is deposited on a glass or plastic plate. The important thing about the silica is that, at its surface, silicon atoms are bonded to —OH groups.

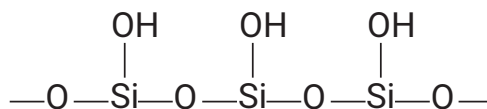


FIGURE 12.2.1 The surface of silica deposited on a TLC plate, showing the —OH groups attached to the silicon atoms

The presence of the —OH groups makes the stationary phase very polar and enables it to form hydrogen bonds with any suitable compounds in the mobile phase.

An important part of TLC is the choice of solvent (the mobile phase). The solvent must be able to dissolve the components of the mixture to be separated and must be attracted, to some degree, to the stationary phase silica.

During TLC, as the chromatogram develops, the solvent begins to soak up the plate. As it travels, the solvent dissolves the mixture of compounds placed at the base line of the plate. As the solvent travels further up the plate, the compounds travel with it. How far they travel depends on how long they stay dissolved in the solvent. This depends on the forces of attraction between the compounds and the:

- solvent
- stationary phase.

For example, a mixture of 1-propanol and propanone was to be separated using TLC. The structures of these compounds are shown in [Figure 12.2.2](#).

Propanol contains the —OH group, which can form hydrogen bonds. Propanone is a polar molecule and can undergo dipole–dipole attractions.

These molecules will both dissolve in a solvent such as ethyl ethanoate (ethyl acetate) and when travelling up a TLC plate coated with silica will travel some distance before one of the compounds becomes stuck to the silica. In this case, the compound that will stick first is 1-propanol because it can form strong hydrogen bonds to the silica. The propanone, being simply polar, will not be as strongly attracted to the silica and will travel further up the plate dissolved in the solvent ([Figure 12.2.3](#)).



Weblink
Investigating thin-layer
chromatography

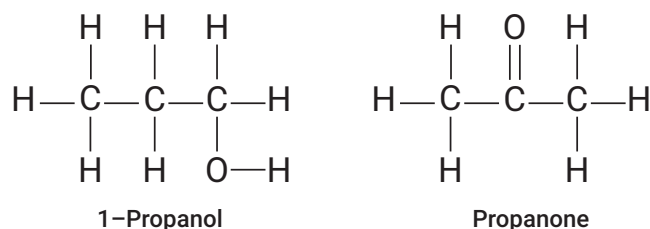


FIGURE 12.2.2 The structures of 1-propanol and propanone

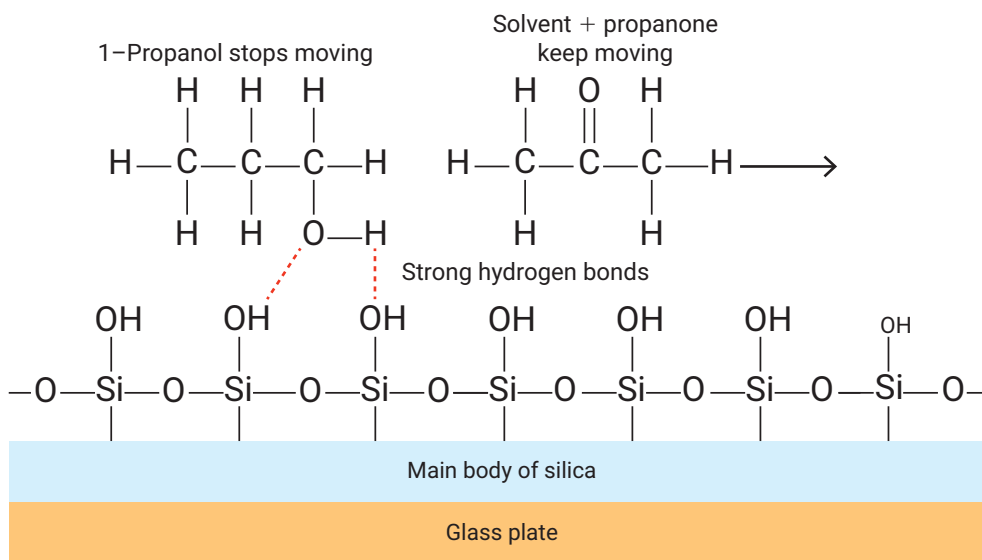


FIGURE 12.2.3 The movement of 1-propanol and propanone along a TLC plate. 1-Propanol forms strong hydrogen bonds with the silica stationary phase, so it will appear on the plate closer to the base line.

Paper chromatography

In paper chromatography, the stationary phase is more complex than in TLC. The paper, typically filter paper or chromatography paper, is composed of cellulose fibres with water trapped in the pores of the paper.

Cellulose is a polymer of the simple sugar, glucose (**Figure 12.2.4**). The cellulose polymer chains have $-\text{OH}$ groups, similar to the silica surface in thin-layer chromatography. However, the cellulose fibres hydrogen bond to water molecules from both the manufacturing process

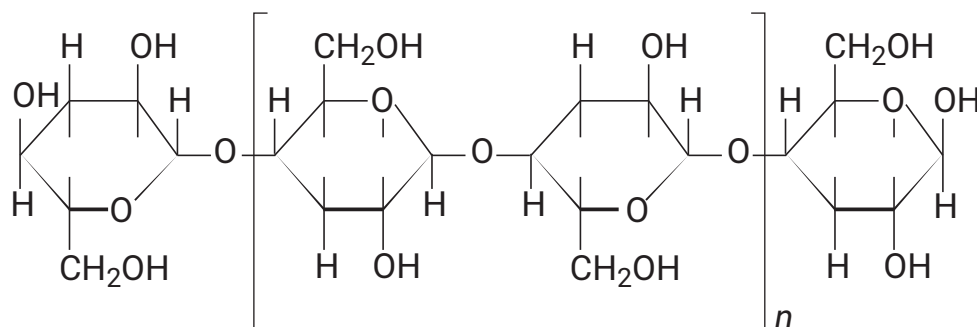


FIGURE 12.2.4 The structure of cellulose

and from the atmosphere. In fact, 'dry' paper is approximately 22% water by weight. Therefore, the paper is cellulose fibres and a thin layer of water molecules bonded to the surface of the fibres. It is the water adsorbed on the paper surface that is the stationary phase in paper chromatography.

Various solvents are used as the mobile phase. Often it is a mixture of water and an alcohol. This makes the mobile phase fairly polar but still less polar than the stationary phase.

The degree to which each component moves depends on its solubility in the particular solvent used and its attraction to the water in the paper through the formation of intermolecular bonds, such as dipole–dipole attraction and hydrogen bonds. As the solvent moves up the paper by capillary action, the more polar components travel more slowly.

Table 12.2.1 lists some common chromatography solvents and their polarity.

TABLE 12.2.1 The polarity of common chromatography solvents

Solvent	Polarity
Petroleum ether	Least polar ↓ Most polar
Cyclohexane	
Methylene chloride	
Diethyl ether	
Ethyl acetate (ethyl ethanoate)	
Acetone	
Ethanol	
Water	

Separations appear to be better obtained when using a mixture of solvents. As the solvent flows through a section of the paper containing the mixture, some of the components leave the stationary phase and enter the mobile solvent phase. When the solvent reaches a section of the paper where there are no mixture components, a component molecule may migrate back to the paper further along. With continuous flow, this may occur several times before the component is firmly deposited.

If a non-polar solvent such as cyclohexane is used, non-polar molecules in the mixture are much less attracted to the water molecules attached to the cellulose and so stay mostly dissolved in the moving solvent. These molecules travel further up the paper and have relatively high R_f values. Polar molecules are highly attracted to the water molecules rather than the non-polar solvent so attach to the thin layer of water around the cellulose fibres and do not travel very far up the paper. These molecules have a low R_f value.



Worksheet
Summarising
chromatography methods

TABLE 12.2.2 A summary of chromatography techniques

Technique	Basis	Stationary phase	Mobile phase	Measures	Uses
Paper chromatography (liquid–liquid chromatography)	Solvent moves over the stationary phase by capillary action. Components adsorb and desorb at different rates because of chemical structure and ability to form intermolecular bonds to the stationary and mobile phases.	Water molecules held inside the pores of the cellulose network of an absorbent host paper such as chromatography paper or filter paper	Solvent – commonly an alcohol or alcohol–water mixture	R_f is characteristic for each sample under the same conditions. Standards are used to confirm identity (qualitative).	Identifying components of organic mixtures such as plant pigments, inks, dyes Determining presence of impurities
Thin-layer chromatography (solid–liquid chromatography)	Solvent moves over the stationary phase by capillary action. Components adsorb and desorb at different rates owing to chemical structure and ability to form intermolecular bonds to the stationary and mobile phases.	Fine silica gel or aluminium oxide powder on glass or plastic	Solvent may be organic or water mixture	R_f is characteristic for each sample under the same conditions. Standards are used to confirm identity (qualitative).	Identifying components of organic mixtures such as plant pigments, amino acids and drugs Identifying impurities Separating less polar substances

LEARNING CHECK 12.2

DESCRIBING

1 Define:

a solvent	b solute
c hydrogen bond	d polar bond
e non-polar bond.	

2 Explain how compounds with hydrogen bonds, dipole–dipole attractions and dispersion forces only would be separated on a TLC plate with a:

- a** non-polar stationary phase
b polar stationary phase.

- 3 **Explain** why paper chromatography is considered to be liquid–liquid chromatography.
- 4 If the R_f value of a component is close to 1, **describe** what this says about the compound's attraction to the stationary phase compared to its attraction to the solvent.

APPLYING

- 5 **Explain** why the spots on a chromatogram can sometimes appear as elongated ovals rather than simple circles.
- 6 It is possible to increase the effectiveness of chromatography by changing the polarity of the solvent. **Consider** if it would be possible to do the same by changing the polarity of the stationary phase.
- 7 **Figure 12.2.5** shows a paper chromatogram prepared using the colours of the coatings on jelly beans. A black, a green, a pink, and a blue jelly bean were used.

- a **State** how many food dyes were used in each colour.
- b For the black jelly bean, **explain** which dye was most attracted to the stationary phase.
- c What could be **inferred** about the molecular structure of the blue and yellow dye?

- 8 Lemon juice was separated by thin-layer chromatography to determine what amino acids it contained. (Amino acids are the building blocks of proteins and will be studied in Unit 4.) Because amino acids are colourless, a developing agent, ninhydrin, was used to reveal the amino acids as purple spots on the chromatogram.

To help identify the amino acids, pure aspartic acid and pure leucine were also spotted onto the chromatography paper.

The results are shown in **Figure 12.2.6**.

- a **Identify** the amino acid that is more strongly attracted to the stationary phase.
- b **Calculate** the R_f value for aspartic acid in this solvent.
- c Draw conclusions about the amino acids in lemon juice.

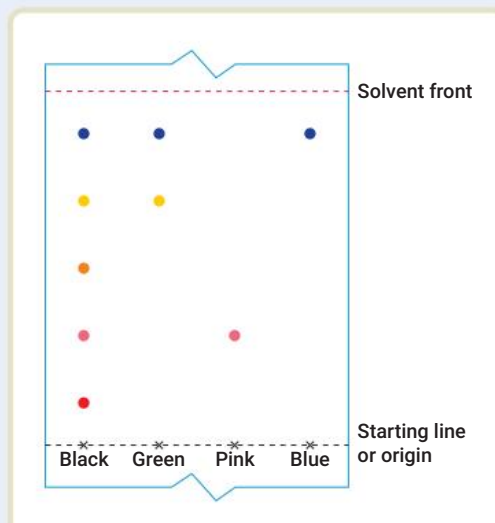


FIGURE 12.2.5 The chromatogram for jelly-bean coatings

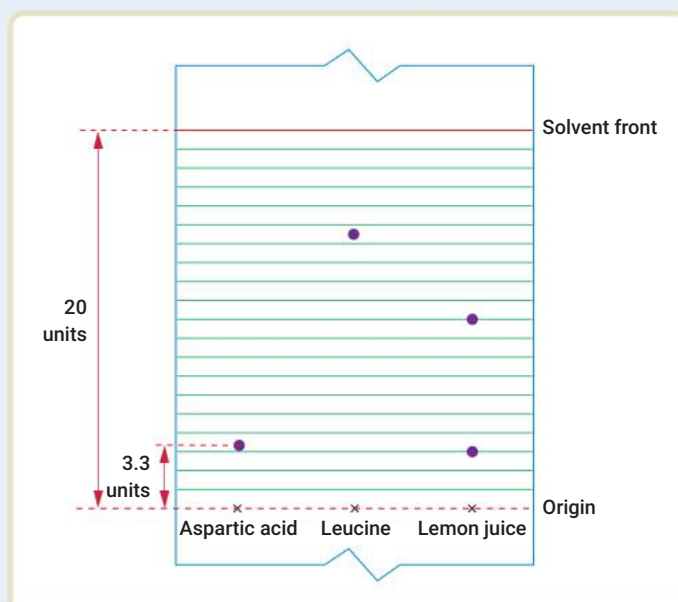
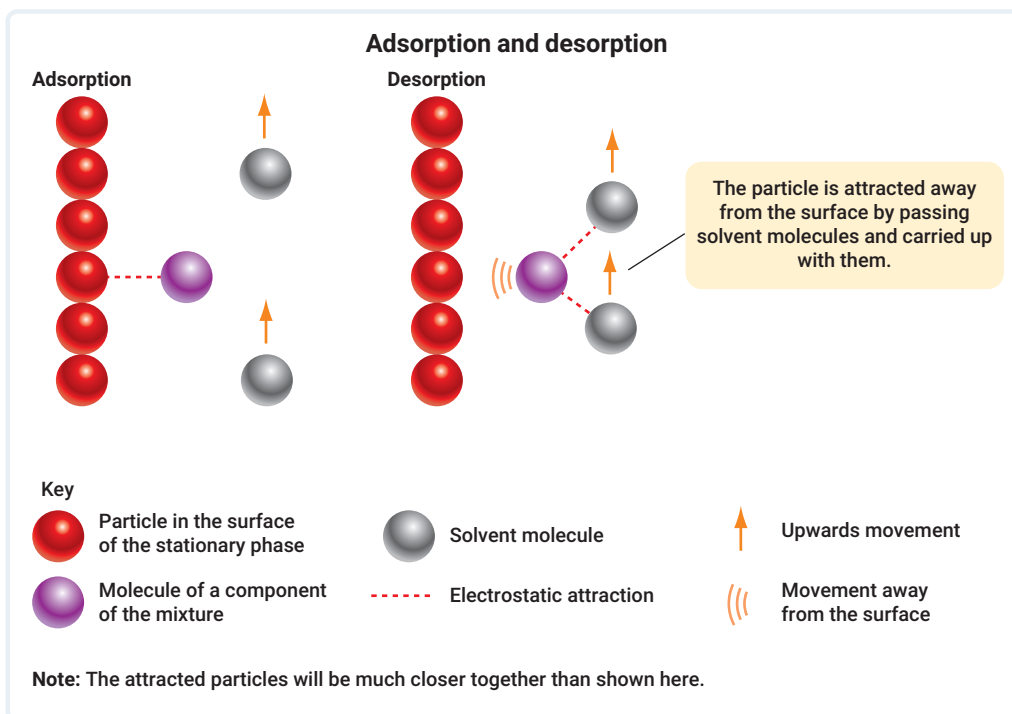


FIGURE 12.2.6 The experiment results

Chromatography

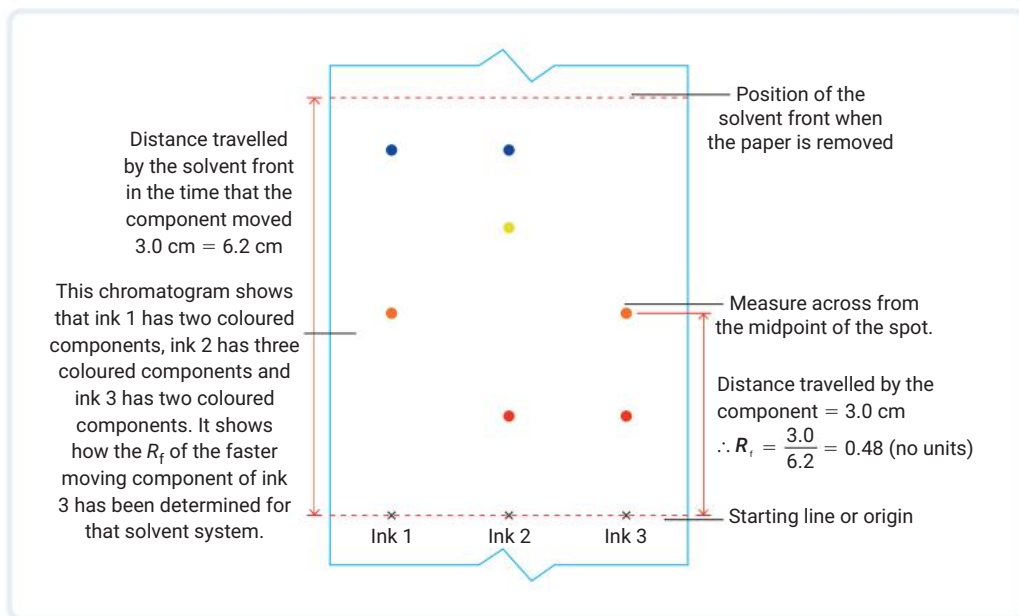
- Paper and thin-layer chromatography are analytical techniques that can be used to determine the composition and purity of substances.
- The chemical basis of chromatography is that different substances adsorb onto a stationary phase and desorb into a mobile phase.



Analysing components using chromatograms

- The rate and distance the component molecules move is determined by how strongly they are attracted to either the stationary or mobile phases.
- The retardation factor (R_f) is the ratio of the distance moved by the component to the distance moved by the solvent front:

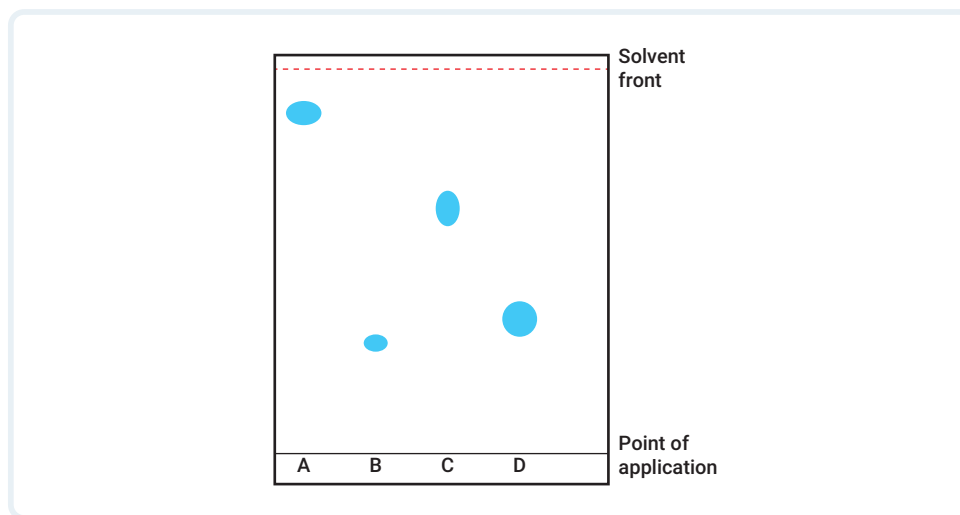
$$R_f = \frac{\text{distance moved by component}}{\text{distance moved by solvent front}}$$



- The R_f value is characteristic for a particular substance with a given absorbent and solvent system. It can be used to identify components of a mixture by comparing experimental values with theoretical values.
- Variations in the strength of intermolecular bonding between the component molecule, stationary phase and mobile phase allow the components to be separated.
- Component molecules that are strongly attracted to the stationary phase move slowly and have a small R_f value, whereas molecules that are less strongly attracted to the stationary phase move more quickly and have a larger R_f value.

MULTIPLE CHOICE

1. In paper chromatography, the stationary phase is:
A a gas.
B a solid.
C a liquid.
D unknown.
2. Which of the following procedures is least likely to affect the analysis of a mixture by paper chromatography?
A Placing the paper in an open dish of solvent
B Handling the paper before the chromatogram is run
C Placing a pencil mark at the starting point of the mixture
D Leaving the paper in contact with the solvent for a very long time
3. A TLC plate is shown in. The solvent used was pentane (a non-polar liquid).



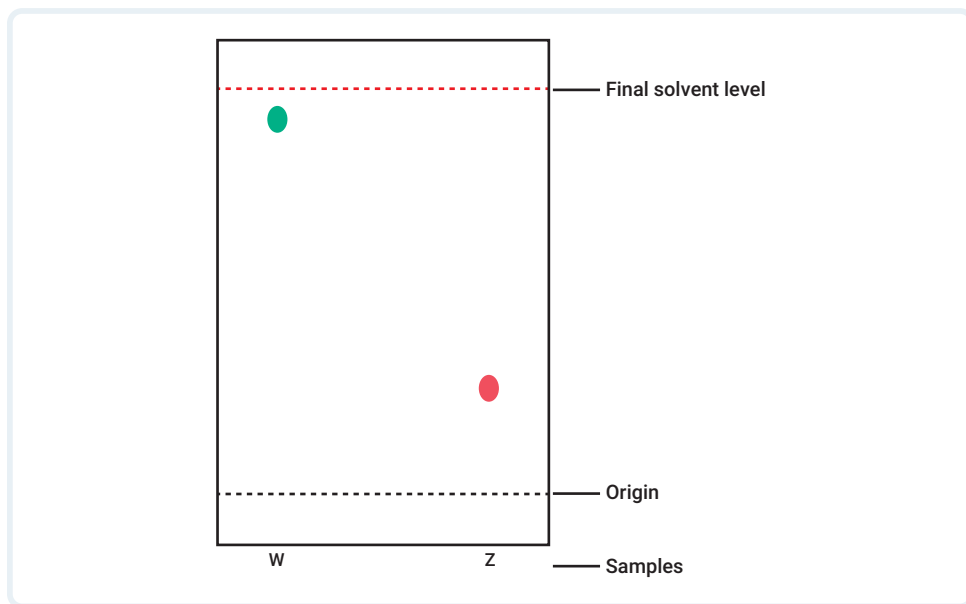
Which sample contains the most polar molecules?

- A A
B B
C C
D D
4. Paper chromatography was used to separate the pigments in a plant leaf. A spot of the pigments was placed on a sheet of chromatography paper and the chromatograph was run inside a closed jar that was partially filled with ethanol.
The mobile phase in this instance is the:
A ethanol.
B paper.
C pigments.
D water in the paper.

5. A student used paper chromatography to separate two components, I and II, in a solution. A spot of the solution was initially placed at the origin. When the spot corresponding to component I ($R_f = 0.50$) had advanced 4.0 cm, the spot corresponding to component II was 1.0 cm behind. Which number is closest to the R_f value of component II?
- A 0.13
B 0.25
C 0.38
D 0.68

©VCAA VCAA 2011 MCQ Q17

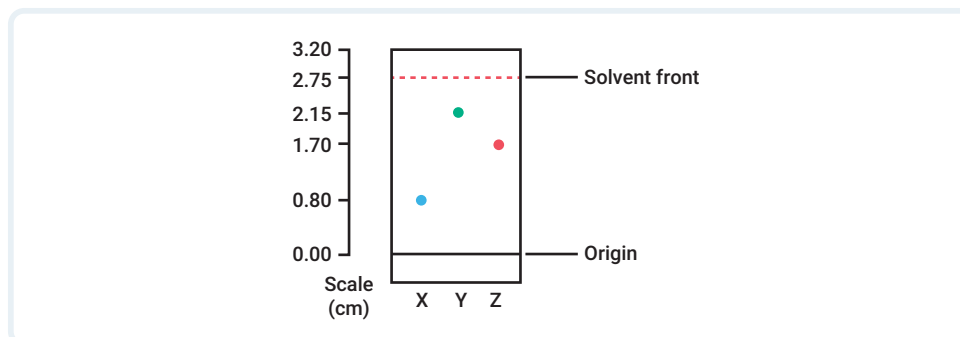
6. Two different food dye samples, W and Z, were compared by using thin-layer chromatography as shown below.



Which of the following statements is correct?

- A Z is more strongly adsorbed than W and has a lower R_f value.
B Z is more strongly adsorbed than W and has a higher R_f value.
C W is more strongly adsorbed than Z and has a lower R_f value.
D W is more strongly adsorbed than Z and has a higher R_f value.
7. Consider the following statements about a thin-layer chromatography that uses a polar solvent and a non-polar stationary phase to analyse a solution:
- Statement I** – Polar molecules in the solution will be attracted to the solvent particles by dipole–dipole attractions.
Statement II – Non-polar molecules in the solution will be attracted to the stationary phase by dispersion forces.
Statement III – Polar molecules in the solution will travel along the TLC plate more rapidly than non-polar molecules.
- Which of these statements are true?
- A I and II only
B I and III only
C II and III only
D I, II and III

8. Consider the following TLC plate of compounds X, Y and Z developed using a suitable mobile phase on a polar stationary phase.



What is the R_f value of the most polar component in this TLC separation?

- A 0.29
B 0.62
C 0.78
D 0.80
9. Two substances, P and Q, were both analysed by paper chromatography and thin-layer chromatography. The solvent used for the mobile phase was methylated spirits. In both forms of chromatography, component P was bonded more strongly to the stationary phase than component Q.
Which of the following statements is correct in terms of the R_f values for P and Q in both paper and thin-layer chromatography?
- A The R_f value for P will be higher than Q for both methods.
B The R_f value for P will be lower than Q for both methods.
C The R_f value for P will be higher than Q for paper chromatography but lower than Q for TLC.
D The R_f value for P will be lower than Q for paper chromatography but higher than Q for TLC.
10. Paper chromatography would be most suitable for:
- A detecting the presence of benzene in methylated spirits.
B separating the compounds present in black dye.
C separating the clay particles from a sample of river water.
D extracting a large quantity of aspirin from tablets containing aspirin and paracetamol.

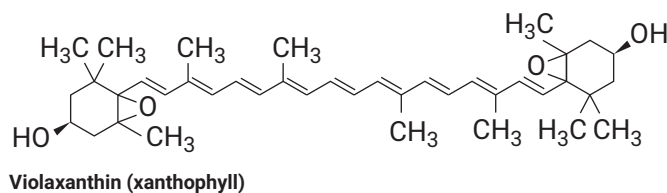
SHORT RESPONSE

11. Define:

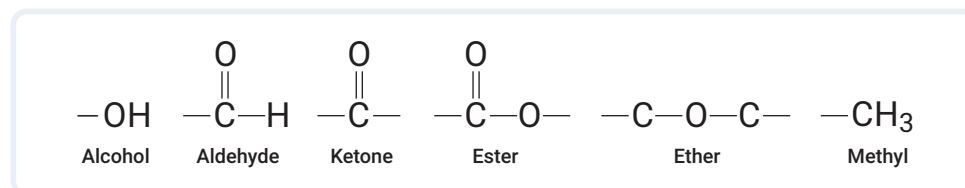
- a adsorption
b desorption
c mobile phase
d stationary phase
e R_f

12. Describe the factors that determine the length of time a component spends in the mobile phase.

13. A student decided to investigate the pigments in leaves by using thin-layer chromatography. They chose to use leaves of *Ficus benjamina*, which is commonly known as weeping fig. They used thin-layer chromatography plates precoated with silica gel and a solvent of hexane (C_6H_{14}), acetone (C_2H_6CO) and trichloromethane ($CHCl_3$) in the ratio 3:1:1. As part of their background research, the student identified possible pigments that would be in the leaf and the chemical structure of the pigments. These are shown below.



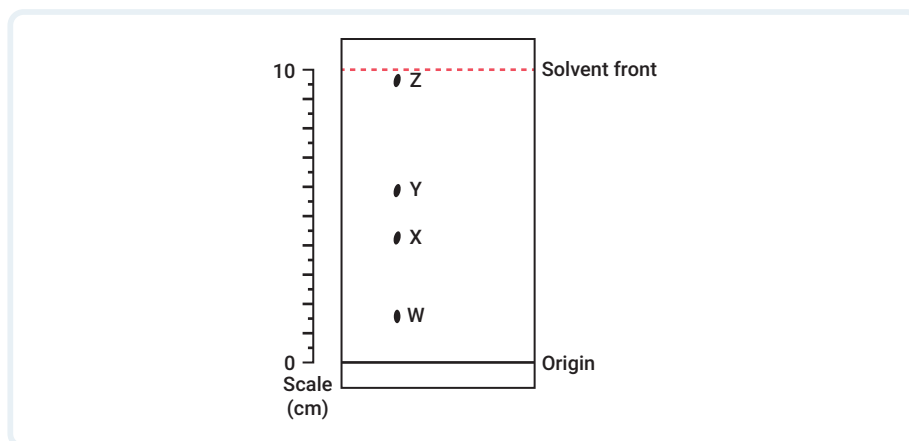
They also learnt that the polarity of the chemical groups on the molecules varied in polarity as follows



Alcohol groups are the most polar, ester and ether groups are the least polar, aldehyde and ketone groups are in between and the methyl group is non-polar.

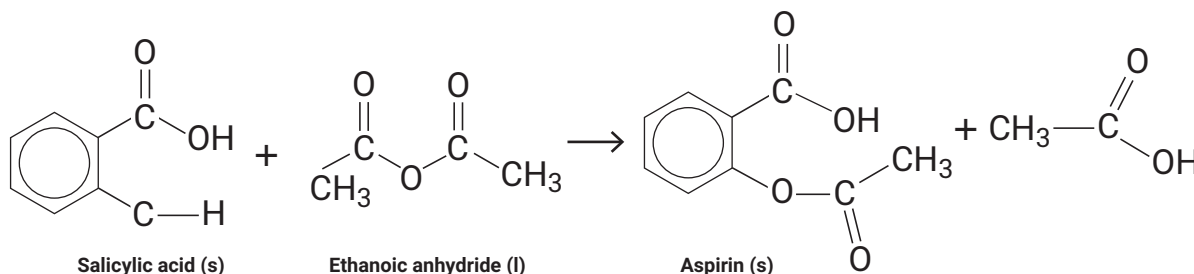
- a Copy the diagrams of the four pigment structures, then **identify** the polarity of different chemical groups by circling the polar groups in blue and the non-polar groups in red.
- b **List** the pigments in order from the most polar to the least polar. **Justify** your order.
- c **Identify** the stationary and mobile phases and **determine** the polarity of each of these.

The student conducted the experiment and obtained the following chromatogram.



- d Match each of the spots on the chromatogram to the corresponding pigment. **Justify** your decision.
- e **Calculate** the R_f value for each of the pigments.

14. A sample of aspirin was prepared by reacting salicylic acid with ethanoic anhydride. The equation for the reaction is:

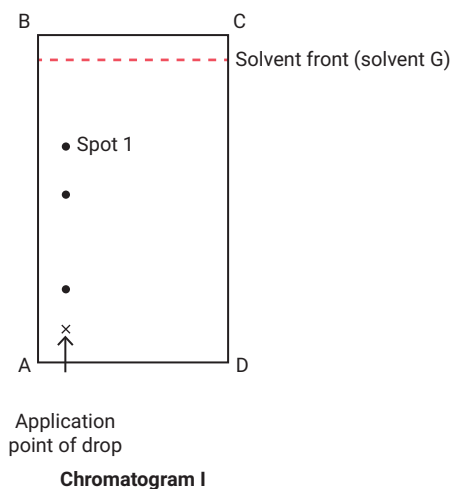


Describe how the purity of the product could be qualitatively determined by thin-layer chromatography. Both salicylic acid and aspirin can be detected with ultraviolet light. Ethyl ethanoate can be used as a suitable solvent.

DATA ANALYSIS

15. **Analyse data**

A drop that contains a mixture of four amino acids was applied to a thin-layer chromatography plate. The plate was placed in solvent G and chromatogram I was obtained.



The R_f values for each of the amino acids in solvent G are provided in **Table 1**.

TABLE 1 R_f values for amino acids in solvent G

Amino acid	R_f (solvent G)
Alanine	0.51
Arginine	0.16
Threonine	0.51
Tyrosine	0.68

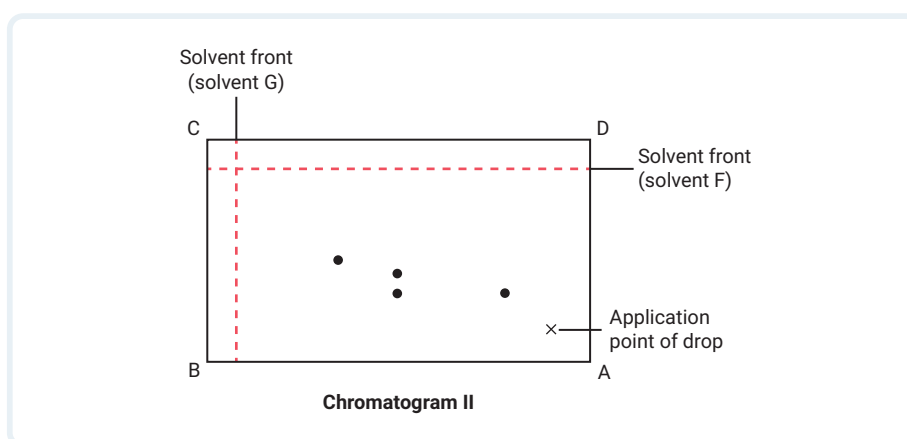
a Identify the amino acid that corresponds to spot 1.

The plate was dried, rotated 90° in an anticlockwise direction and then placed in solvent F. The R_f values for each of the amino acids in solvent F are provided in **Table 2**.

TABLE 2 R_f values for amino acids in solvent F

Amino acid	R_f (solvent F)
Alanine	0.21
Arginine	0.21
Threonine	0.34
Tyrosine	0.43

Chromatogram II was obtained.



- b** Copy the chromatogram and **identify** the spot on chromatogram II that represents alanine.
- c Explain**, in terms of the data provided, why only three spots are present in chromatogram I while four spots are present in chromatogram II.



Craig Sterken/Shutterstock.com

SYLLABUS
DOT POINTS**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 and/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 INQUIRY

Investigate:

- Boyle's law or the molar volume of a gas.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Earth's atmosphere consists almost entirely of gas. Gases are all around us and inside us. An understanding of the behaviour of gases is of crucial importance to our industry, our economy and our environment.

An appreciation of the particulate nature of gases enables us to understand the properties of gases, such as pressure, temperature and volume, which, in turn, leads to an understanding of how chemical reactions involving gases occur – knowledge that is vital in industry.

Practicals

- Estimating the molar volume of a gas

Worksheets

- Boyle's law
- Charles' law
- Ideal gas law



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ Gas particles have high kinetic energy.
- ✓ There is a fixed number of particles in a mole of particles.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ identify the conditions of standard temperature and pressure (STP)
- ✓ identify the connection between the volume of a gas, the number of moles and the molar volume at STP
- ✓ use the concept of molar volume to calculate unknown values at STP
- ✓ describe the kinetic theory of gases
- ✓ use the kinetic theory of gases to explain the relationship between volume, pressure and temperature
- ✓ use the ideal gas equation to calculate unknown values at standard temperature and pressure (STP)
- ✓ develop an understanding of the relationship between pressure, temperature and volume from data
- ✓ explain the relationship between an ideal gas and a real gas.

13.1 Gases

gas a substance (element, compound or mixture) that exists in a state with no defined volume or shape

A **gas** is a substance – an element, compound or a mixture – that exists in a state with no defined volume or shape. Because gases can move freely, they are able to expand to fill the container that they are in.

When considering the volume of gases, it is important to note the difference between the volume of the gas particles and the volume of the container in which the gas is held. The volume of the gas particles is fixed but, because the gas particles are spread out, they can be brought together (compressed) into a smaller volume. When scientists refer to the volume of a gas, they are referring to the volume of the container. Under normal conditions, the total volume of space taken up by the gas particles themselves is a tiny fraction of the total volume of the container (**Figure 13.1.1**).

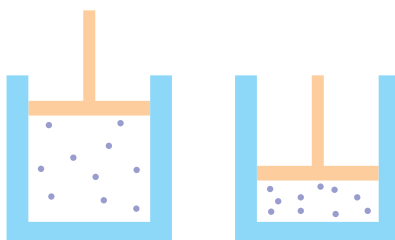


FIGURE 13.1.1 Gases take up the volume of their container. The gas in the second container (right) has been compressed so the particles occupy a smaller volume.

Volume can be measured using different units (**Table 13.1.1**). The SI unit of volume is the cubic metre (m^3). Chemists generally use the unit of litre (L).

TABLE 13.1.1 Commonly used units of volume

Name of unit	Symbol	Comparison
Cubic metre	m^3	Standard unit
Cubic decimetre	dm^3	$1000 \text{ dm}^3 = 1 \text{ m}^3$
Cubic centimetre	cm^3	$1000 \text{ cm}^3 = 1 \text{ dm}^3$
Litre	L	$1 \text{ L} = 1 \text{ dm}^3$
Millilitre	mL	$1 \text{ mL} = 1 \text{ cm}^3$

The *Formula and data book* contains helpful volume and capacity conversion.

The relationship between number of moles of gas and volume

Chemists find it more convenient to think in terms of the number of gas particles present in a container, rather than just its volume, because it more accurately represents the amount of matter or mass of that gas.

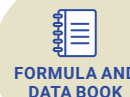
In 1811, Amedeo Avogadro attempted to show how the volume of a gas was related to the amount of gas present. He proposed that equal volumes of gas contain equal numbers of particles at the same temperature and pressure. This became known as **Avogadro's hypothesis**. Avogadro's hypothesis states that equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles. This means that, regardless of the composition of the gas, the total number of particles is the same under the same conditions. So, provided temperature and pressure are the same, 1 L of hydrogen gas has the same number of particles as 1 L of helium gas, or 1 L of oxygen gas and so on. Also, if the number of particles is doubled, then the volume doubles.

To compare changes in conditions that involve gases, scientists use a set of standard conditions for temperature and pressure. In 1982, the International Union of Pure and Applied Chemistry (IUPAC) established **standard temperature and pressure (STP)** as a temperature of 273.15 K (0°C) and an absolute pressure of 100.00 kPa (0.987 atm or 1.000 bar). The *Formula and data book* has these conditions as 273 K and 100 kPa.

Molar volume

The mole is defined as a fixed number of particles (Avogadro's number), so it follows that equal volumes of different gases, measured at the same temperature and pressure, contain the same number of moles. This is called the **molar volume of a gas**.

The units of the Avogadro number may be electrons, atoms, ions or molecules, depending on the nature of the substance. At STP (0°C and 100 kPa) the molar volume of a gas is 22.71 L mol^{-1} . In the *Formula and data book*, this is rounded to 22.7 L mol^{-1} . Under standard laboratory conditions (SLC) (25°C and 1 atm), the molar volume of a gas is 24.8 L mol^{-1} because the higher temperature causes the gas to expand and the volume to increase.



FORMULA AND
DATA BOOK

Avogadro's hypothesis
equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles



FORMULA AND
DATA BOOK

molar volume of a gas
the volume occupied by 1 mole of gas at a known temperature and pressure



Syllabus link
Chapter 9 discusses Avogadro's number.

KEY FORMULA

The relationship between number of moles and volume of gas at STP can be expressed as:

$$\text{Number of moles of gas} = \frac{\text{volume of gas at STP}}{\text{volume of 1 mole of gas at STP}}$$
$$n = \frac{V}{22.7}$$

where: n = number of moles

V = volume of gas (L) at STP.

WORKED EXAMPLE 13.1.1

- a** How many moles of carbon dioxide gas are present in 350 mL of gas at STP?
b What is the mass of carbon dioxide in part **a**?
c What volume is occupied by 160 g of helium at STP?

ANSWERS

- a 1** Write the relationship between the number of moles of a gas (n), gas volume (V , in L) and volume of 1 mole of gas at STP (22.7 L) using the formula:

$$n = \frac{V}{22.7} \text{ mol}$$

- 2** Determine the quantities and check the units.

$$n = ?, V = 350 \text{ mL} = 0.350 \text{ L}$$

- 3** Substitute to calculate moles.

$$n = \frac{0.350}{22.7}$$
$$= 0.0154 \text{ mol}$$

- b 1** Write the relationship between number of moles (n), mass (m) and molar mass (M) using the formula.

$$n = \frac{m}{M}$$

- 2** Determine the quantities of n and M .

From part **a**, $n = 0.0154 \text{ mol}$.

$$M(\text{CO}_2) = 12.01 \text{ g (C)} + (2 \times 16.00) \text{ g (O}_2\text{)}$$
$$= 44.01 \text{ g mol}^{-1}$$

- 3** Substitute to calculate mass.

$$0.0154 \text{ mol} = \frac{m}{44.01}$$

$$m = 0.0154 \text{ mol} \times 44.01 \text{ g mol}^{-1}$$
$$= 0.678 \text{ g}$$

- c 1** Write the relationship between the number of moles of a gas (n), gas volume (V , in L) and volume of 1 mole of gas at STP (22.7 L) using the formula:

$$n = \frac{V}{22.7}$$

2 Determine the quantities and check the units.

$$n = ?, V = ?, m = 160 \text{ g}$$

3 Calculate the number of moles.

$$n = \frac{m}{M}$$

$$m = 160 \text{ g}, M(\text{He}) = 4.00 \text{ g mol}^{-1}$$

$$n = \frac{160}{4.00} = 40 \text{ mol}$$

4 Substitute to calculate volume.

$$n = \frac{V}{22.7}$$

$$40 = \frac{V}{22.7}$$

$$V = 40 \times 22.7 \text{ L} = 908 \text{ L}$$

LEARNING CHECK 13.1

DESCRIBING

1 Define:

- a gas
- b Avogadro's hypothesis
- c molar volume.

2 What are the conditions at STP?

APPLYING

- 3 If the volume of 0.13 mol of a gas at room temperature is 385.9 mL and the volume of the same amount of gas at the same temperature is 35.9 mL, what condition must be different? **Explain** your answer.
- 4 a What volume container would you need to store 0.095 mol of nitrogen at STP?
- b i How many moles are present in 0.450 L of oxygen gas at STP?
ii What is the mass of oxygen in part i?
- c i What volume will 10 g of hydrogen gas occupy at STP?
ii What volume will 10 g carbon dioxide occupy at STP?
iii **Explain** why these volumes differ when the mass is the same.
- 5 A 1.5 L sample of helium and oxygen was collected in two different containers at STP. **Compare** the number particles of helium and oxygen. **Explain** your answer.

13.2 Kinetic theory

Experimental observations by chemists that different gases behave in a similar manner led to the development of mathematical relationships linking volume, mass, pressure and temperature. These relationships are based on the notion of an **ideal gas**; that is, one that obeys the relationships perfectly. Although no real gas behaves in a completely ideal manner, most gases are considered as being sufficiently close to ideal under normal conditions that these relationships can be used.

ideal gas a gas that perfectly obeys all the proposals of the kinetic theory of gases and the gas laws

kinetic theory a model used to explain the behaviour of gases based on their movement

elastic collision a collision with no net loss of energy



Weblink
Kinetic theory of gases

The **kinetic theory** of gases explains much of the physical behaviour of gases. The term 'kinetic' comes from the Greek *kinetikos*, which means 'moving', and indeed many of the properties of gases relate to their motion.

The kinetic theory of gases proposes that:

- gases consist of molecules (except the noble gases, which consist of atoms) that move in continual random straight-line motion
- the average distance between gas molecules is very large compared to the size of the molecule
- intermolecular forces between gas molecules are negligible because of the relatively large distances between molecules
- all collisions of gas molecules are perfectly **elastic collisions**, which means no net energy is lost during these collisions
- pressure is due to collisions of the molecules with the walls of the container
- temperature is a measure of the average kinetic energy of the molecules.

Note that this theory uses the model of an ideal gas. Gases such as hydrogen and helium, which have small, light particles come close to 'ideal' under conditions of low pressure and high temperature. Chemists apply kinetic theory to real gases under normal conditions.

Temperature

As mentioned in Chapter 10, temperature is a measure of average kinetic energy. When considering a gas sample, the average kinetic energy of the molecules is used; however, the individual molecules move at various speeds. Some move faster and some move slower than the average. Collisions between particles change the speeds of particles.

However, at a given temperature, all gases, no matter what size, shape or mass, have the same average kinetic energy. As the temperature increases, the kinetic energy of the particles increases, and vice versa. Although the gas particles have a range of speeds at any given temperature, the average speed of the particles of a gas increases as temperature increases (**Figure 13.2.1**).

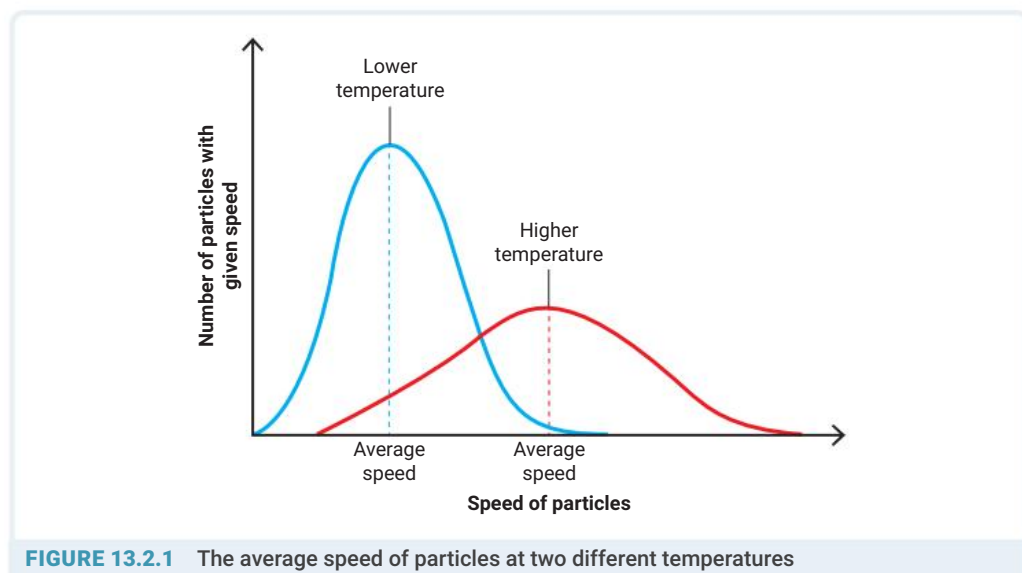


FIGURE 13.2.1 The average speed of particles at two different temperatures

This means that decreasing the temperature results in the particles slowing down. Theoretically, at low enough temperatures, the particles will eventually stop moving. The theoretical lowest temperature possible is known as **absolute zero** (0 K), which is equivalent to -273.15°C . The relationship to convert temperature from kelvin to degrees Celsius is:

$$T(\text{in K}) = T(\text{in } ^{\circ}\text{C}) + 273$$

absolute zero
precisely 0 K on the Kelvin scale, which is a thermodynamic (absolute) temperature scale and -273.15°C on the Celsius scale

Pressure

According to the kinetic theory of gases, pressure is due to collisions of the molecules with the walls of the container. Gas pressure is a measure of the force per unit area exerted by gas particles on the walls of their container.

The International System of Units (SI) unit of gas pressure is the **pascal** (Pa); 1 Pa is equivalent to a force of 1 newton per square metre. This is a very small quantity, so hectopascals and kilopascals are more commonly used for measuring pressure.

pascal the SI unit of gas pressure, equivalent to a force of 1 newton per square metre; symbol Pa

The gases in the atmosphere exert pressure on every exposed surface. The pressure due to the atmosphere is called atmospheric pressure. This pressure varies from place to place and time to time. At sea level, atmospheric pressure is like having a 1 kg tub of margarine pushing on every square centimetre of every part of the surface of your body all the time. Usually, it goes unnoticed because the external air pressure exerted by the atmosphere is balanced by the internal pressure exerted outwards from the body's tissues and blood.

To make communication easier, scientists agreed on a standard of pressure to represent the average air pressure at sea level. This unit of standard atmospheric pressure is 1.01325×10^5 Pa or 101.325 kPa (kilopascals) at sea level. Another common unit for pressure is the atmosphere (atm): $1.00 \text{ atm} = 101.3 \text{ kPa}$.

Atmospheric pressure varies with height above sea level, temperature and weather conditions, so it is not always 1.00 atm. Meteorologists generally mention high and low pressure systems in weather reports and use the hectopascal (1 hPa = 100 Pa) or millibar (100 Pa) units when discussing pressure. Chemists and physicists generally use the units pascal (or kilopascal) and atmosphere (**Table 13.2.1**).

TABLE 13.2.1 Commonly used units of gas pressure

Name of unit	Symbol	Comparison
Pascal	Pa	Standard unit, 1 newton per square metre (N m^{-2})
Atmosphere	atm	101.3 kPa
Hectopascal	hPa	100 Pa
Kilopascal	kPa	1000 Pa
Millimetres of mercury	mmHg	$760 \text{ mmHg} = 1 \text{ atm}$

Applying the kinetic theory

The kinetic theory developed from the findings of many different scientists. It can be used to explain many properties of gases (Figure 13.2.2).

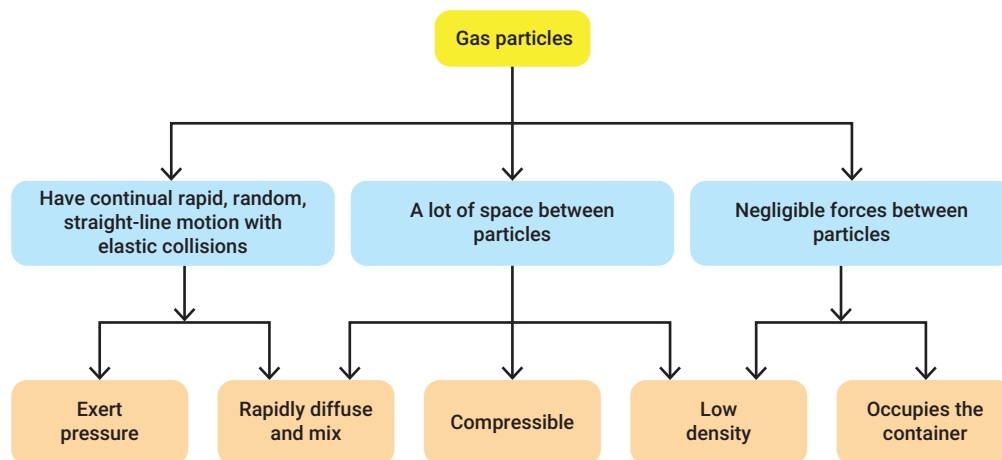


FIGURE 13.2.2 A schematic diagram of the kinetic theory of gases

The relationship between pressure, temperature and volume

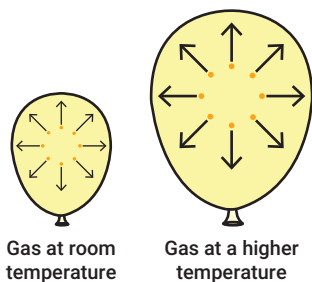


FIGURE 13.2.3 The relationship between temperature and volume for an enclosed gas where the pressure remains constant

Consider a gas placed in a piston (Figure 13.1.1). If the space inside the piston is made smaller, the gas particles occupy a smaller space. If the temperature is kept constant, then, according to the kinetic theory, the gas particles will move randomly at the same speeds. More collisions will occur per unit area of the walls of the piston. In other words, the pressure will increase.

Consider a balloon filled with gas at room temperature (Figure 13.2.3). If the temperature of the balloon is increased, then, according to the kinetic theory, the gas particles will move faster and the force per unit area exerted by the gas particles will increase. The volume of the balloon will increase.

The relationships between pressure, temperature and volume are encapsulated quantitatively in what have become known as the gas laws.

Gas laws

Boyle's law

More than 300 years ago, Robert Boyle was the first chemist to perform quantitative experiments on gases. By measuring the pressures and volumes of many different gases, he observed that, at a constant temperature, the volume of a given mass of gas is inversely proportional to the pressure (Figure 13.2.4a). This is known as Boyle's law.

In mathematical terms, Boyle's law can be expressed as $V \sim 1/P$, which means the greater the applied pressure, the smaller the volume (Figure 13.2.4b).

As can be seen in graph in Figure 13.2.4a, an increase in volume leads to a decrease in pressure. When expressed as an inverse relationship (Figure 13.2.4b), we can see that Boyle's law is still true.



Weblink
Boyle's law

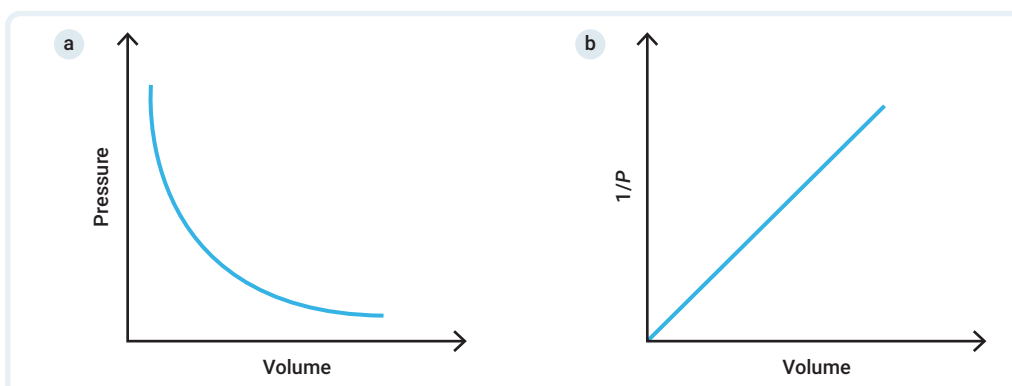


FIGURE 13.2.4 The relationship between pressure and temperature: (a) the overall relationship, (b) the graph expressed as $\frac{1}{P}$, showing a linear relationship.

Another way to state Boyle's law is $PV = k$, where k is a constant. For a particular sample of gas at a constant temperature, the product of the pressure and volume is a constant value.

A useful alternative form of Boyle's law is:

$$P_1 V_1 = P_2 V_2$$

where: P_1 = initial pressure

V_1 = initial volume

P_2 = final pressure

V_2 = final volume.

Note: Any units for pressure and volume can be used, as long as they are consistent.



Worksheet
Boyle's law

KEY LAW

Boyle's law states that, at a constant temperature, the volume of a given mass of gas is inversely proportional to the applied pressure.

$$P_1 V_1 = P_2 V_2$$

WORKED EXAMPLE 13.2.1

A sample of gas collected in a 242 mL container has a pressure of 87.6 kPa. What would the volume of this gas be at 101.3 kPa? (Assume temperature remains constant.)

ANSWER

1 Write the relationship.

$$P_1 V_1 = P_2 V_2$$

2 Determine the quantities.

$$P_1 = 87.6 \text{ kPa}, P_2 = 101.3 \text{ kPa}, V_1 = 242 \text{ mL}, V_2 = ?$$

3 Rearrange the formula.

$$V_2 = \frac{P_1 V_1}{P_2}$$

4 Substitute to calculate volume.

$$\begin{aligned} V_2 &= \frac{87.6 \times 242}{101.3} \\ &= 209 \text{ mL} \end{aligned}$$

Now that we conceptually understand this law, it's possible to analyse data to determine if the law applies.

WORKED EXAMPLE 13.2.2

Students wanted to conduct an experiment to test Boyle's law. In the experiment, gas was drawn into a syringe, the set-up was connected to a pressure gauge and the temperature was kept constant. The initial pressure and volume of gas were recorded. The volume of the gas was adjusted, and both the volume and pressure were recorded for each measurement. The results are shown below.

Volume (mL)	25	22	19
Pressure (kPa)	102	118	134

Determine if the results confirm Boyle's law

ANSWER

1 Write the relationship.

$$P_1V_1 = P_2V_2$$

2 Determine the quantities.

$$P_1 = 102 \text{ kPa}, V_1 = 25 \text{ mL}, P_2 = 118 \text{ kPa}, V_2 = ?$$

(Although we have data, we want to confirm if this is correct.)

3 Rearrange the formula and substitute quantities.

$$\begin{aligned} V_2 &= \frac{P_1V_1}{P_2} \\ &= \frac{102 \times 25}{118} \\ &= 21.6 = 22 \text{ mL} \end{aligned}$$

The results confirm Boyle's law.

For Worked example 13.2.2, to confidently confirm Boyle's law, we would ideally use the third set of data in the calculations.

Charles' law

About 100 years after Boyle investigated the relationship between pressure and volume, French physicist Jacques Charles studied the relationship between temperature and volume using gases. The results of his many experiments led him to determine that for a fixed quantity of gas at a constant pressure, the volume increases linearly with temperature. This relationship is called Charles' law. When experimental temperature and volume data are plotted, a linear relationship emerges (**Figure 13.2.5**).

Extrapolation of the data ($V = 0$) leads to a temperature of -273°C , now known as absolute zero. Charles' experiments led to the development of a new temperature scale. If the data is plotted using the absolute temperature scale, all lines pass through the origin of the graph, so temperature (K) is directly proportional to volume. Furthermore, in calculations involving gases, temperature must be given in kelvin.

In mathematical terms, Charles' law can be expressed as $V \propto T$ or:

$$\frac{V}{T} = \text{constant}$$

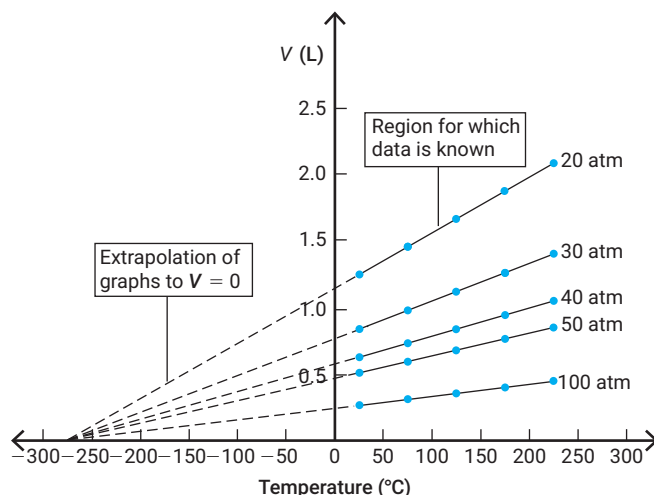


FIGURE 13.2.5 A graph showing how the volume of 1 mole of hydrogen gas varies with temperature for a number of different pressures



Weblink
Charles' law

Worksheet
Charles' law

An alternative form of this equation useful for calculations is:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

where: V_1 = initial volume

T_1 = initial temperature (K)

V_2 = final volume

T_2 = final temperature (K).

Note that any unit for volume can be used, as long as it is used consistently, but temperature must be in kelvin.

With this in mind, we can now analyse data to identify whether or not Charles' law is being obeyed.

KEY LAW

Charles' law states that at constant pressure, the volume of a fixed quantity of gas is directly proportional to its absolute (kelvin) temperature.

WORKED EXAMPLE 13.2.3

A 225 mL volume of gas is collected at 58°C. What volume would this sample occupy at 25°C?

ANSWER

1 Write the relationship.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

2 Determine the quantities and convert degrees Celsius to kelvin by adding 273.

$$V_1 = 225 \text{ mL}, V_2 = ?, T_1 = 58^\circ\text{C} = 331 \text{ K}, T_2 = 25^\circ\text{C} = 298 \text{ K}$$

3 Rearrange the formula.

$$V_2 = \frac{T_2 V_1}{T_1}$$

4 Substitute to calculate volume.

$$\begin{aligned} V_2 &= \frac{225 \times 298}{331} \\ &= 203 \text{ mL} \end{aligned}$$

WORKED EXAMPLE 13.2.4

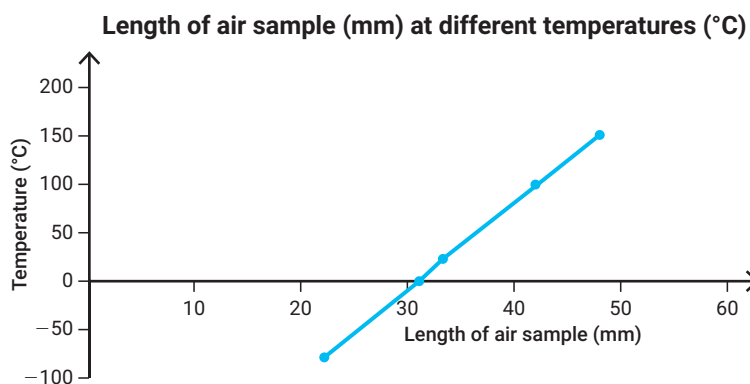
Students conducted an experiment to identify the relationship between the temperature and volume of a gas. In the experiment, they trapped a sample of gas in a capillary tube that was taped to a ruler. They then immersed the apparatus in different liquids at different temperatures and measured the length of the air sample in the capillary tube. The results are shown in the following table.

Liquid used and its temperature	Length of air sample (mm)
Methanol cooled with dry ice to -78°C	22
Ice water, 0°C	31
Room temperature water, 23°C	33
Boiling water, 100°C	42
Heated ethylene glycol, 151°C	48

- Construct a graph to display the results of this experiment.
- Using the graph, determine whether the results obey Charles' law.

ANSWER

- Use the data in the table to draw a graph.



- Identify the trend shown by the graph.**
The graph shows that as the temperature increases, the length of the air sample (and thereby volume) increases.
 - Determine whether this aligns with Charles' law.**
This is Charles' law.

Combined gas law

When dealing with gases, temperature and pressure typically change at the same time. By combining Boyle's law and Charles' law, a new relationship called the combined gas law is produced. This gives a relationship between volume, pressure and absolute temperature for a given mass of gas. It is most useful in the following form:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

where: P_1 , V_1 and T_1 are initial pressure, volume and temperature (K)

P_2 , V_2 and T_2 are final pressure, volume and temperature (K).

Temperature must be measured in kelvin and the units for volume and pressure must be consistent.

WORKED EXAMPLE 13.2.5

A sample of gas at a pressure of 105 kPa has a volume of 2.50 L at a temperature of 25°C. What pressure is needed to compress it to 1.0 L at 20°C?

ANSWER

- 1 Write the relationship.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

- 2 Determine the quantities and convert temperature to kelvin by adding 273.

$$P_1 = 105 \text{ kPa}, P_2 = ?, V_1 = 2.50 \text{ L}, V_2 = 1.0 \text{ L}, T_1 = 25^\circ\text{C} = 298 \text{ K}, T_2 = 20^\circ\text{C} = 293 \text{ K}$$

- 3 Rearrange the formula.

$$P_2 = \frac{P_1 V_1 T_2}{T_1 V_2}$$

- 4 Substitute to calculate pressure.

$$\begin{aligned} P_2 &= \frac{105 \times 2.5 \times 293}{298 \times 1.0} \\ &= 258 \text{ kPa} \end{aligned}$$

The ideal gas equation

All the gas laws studied so far have required that at least one of the conditions of pressure, temperature, volume and quantity of gas remain constant. In reality, when dealing with gases in situations such as scuba diving and using anaesthetics, all the conditions can vary. However, if Boyle's and Charles' laws are combined with Avogadro's hypothesis, the resulting relationship, called the ideal gas equation, includes all conditions as variables.

The ideal gas equation is also known as the ideal gas law and the universal gas equation. It provides a relationship between pressure, volume, temperature and the amount of gas:

$$PV = nRT$$

where: P = pressure (kPa)

V = volume (L)

n = total number of moles of gas present

R = ideal gas constant = $8.31 \text{ J mol}^{-1} \text{ K}^{-1}$

T = temperature (K).

Regardless of the chemical formula and properties of different gases, the ideal gas equation describes the physical properties of all gases. The equation is called the ideal gas equation because real gases deviate slightly from it, particularly at high pressures and low temperatures, because the intermolecular forces between real gases come into play as the particles slow down and move closer together.

It is important to use the correct value of R in the ideal gas equation. Always include the unit and the number when recording the value of a physical quantity to reduce mistakes.

The value of R , the ideal gas constant, depends on units. The value of $8.31 \text{ J mol}^{-1} \text{ K}^{-1}$ is based on an SI unit for pressure. Sometimes, it is more convenient to use pressure in atmospheres. When the unit for pressure is atmospheres, the value of R is $0.0820 \text{ L atm mol}^{-1} \text{ K}^{-1}$. **Table 13.2.2** shows the units and values for R .



Weblink
Ideal gas law simulation

Worksheet
Ideal gas law

KEY FORMULA

The ideal gas equation, $PV = nRT$, gives the relationship between pressure, temperature, volume and number of moles of a gas.



**FORMULA AND
DATA BOOK**

TABLE 13.2.2 Units and values for the ideal gas constant, R

Condition	$R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$	$R = 0.082 \text{ L atm mol}^{-1} \text{ K}^{-1}$
Pressure	kPa	atm
Volume	L	L
Temperature	K	K
Amount of gas	mole	mole

WORKED EXAMPLE 13.2.6

What is the mass of carbon dioxide in a sample that occupies 1.50 L at 20°C and 125 kPa?

ANSWER

- 1 Write the relationship.**

$$PV = nRT$$

- 2 Determine the quantities and check the units.**

$$m = ?, P = 125 \text{ kPa}, V = 1.50 \text{ L}, n = ?, R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}, T = 20^\circ\text{C} = 293 \text{ K}$$

- 3 Rearrange the formula.**

$$n = \frac{PV}{RT}$$

- 4 Substitute to calculate moles.**

$$\begin{aligned} n &= \frac{125 \times 1.5}{8.31 \times 293} \\ &= 0.077 \text{ mol} \end{aligned}$$

- 5 Calculate mass from moles.**

$$n = \frac{m}{M}$$

$$\therefore m = n \times M$$

$$\text{where: } M(\text{CO}_2) = 44$$

$$m = 0.077 \times 44 = 3.39 \text{ g}$$

WORKED EXAMPLE 13.2.7

A 50 mL sample of nitrogen gas exerts a pressure of 100 kPa at 271 K. Determine the mass of gas in the container.

ANSWER

- 1 Write the relationship.**

$$PV = nRT$$

- 2 Determine the quantities and check the units.**

$$P = 100 \text{ kPa}, V = 50 \text{ mL}, n = ?, R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}, T = 271 \text{ K}$$

3 Substitute to calculate the number of mol of gas present.

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$= \frac{100 \times 0.050}{8.31 \times 271}$$

$$= 0.0022$$

4 Calculate mass from moles.

$$m = 0.0022 \times (14.01 \times 2)$$

$$= 0.062 \text{ g}$$

Chemical reactions involving gases

In Chapter 9, we looked at mass calculations based on amounts of reactants and products. The calculations needed to use the mole relationships given by coefficients in a balanced chemical equation are used to determine the amounts of reactants and products. This relationship can now be extended to include volumes of gases (Figure 13.2.6 and Worked example 13.2.7).

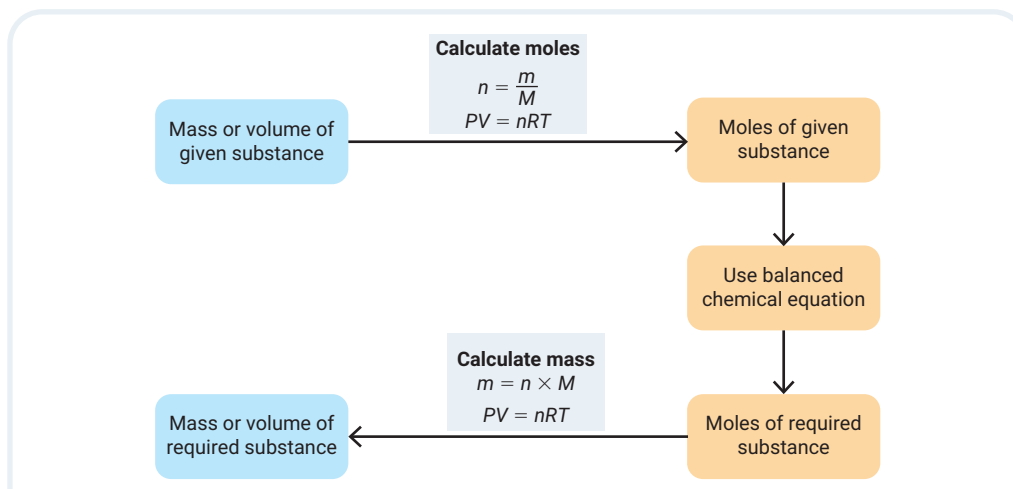


FIGURE 13.2.6 Mass–volume relationships, where m is mass in grams, n is number of moles, M is molar mass, P is pressure in kPa, V is volume in L, T is temperature in K and R is ideal gas constant

WORKED EXAMPLE 13.2.8

Consider the reaction between an excess of dilute hydrochloric acid (HCl(aq)) and 0.325 g of zinc at STP. You already know how to calculate the mass of hydrogen (H_2) produced. By adopting the procedure outlined in Figure 13.3.1, it is now possible to determine the *volume* of the gas produced.

What volume of hydrogen is produced when 0.325 g of zinc metal reacts with an excess of hydrochloric acid at STP?

ANSWER

1 Write the balanced equation.



2 Convert mass of Zn to moles.

$$\begin{aligned}\text{Moles of zinc } (n) &= \frac{0.325}{65.4} \\ &= 4.97 \times 10^{-3} \text{ moles}\end{aligned}$$

3 Use the balanced equation to determine the stoichiometric ration between Zn and H₂.

1 mole Zn : 1 mole H₂

Therefore 4.97×10^{-3} mol of Zn produces 4.97×10^{-3} mol of H₂(g).

4 Use Avogadro's hypothesis to determine the volume of H₂(g) produced.

$$\text{Number of moles of gas} = \frac{\text{volume of gas at STP}}{\text{volume of one mole of gas at STP}}$$

$$n = \frac{V}{22.7}$$

$$\begin{aligned}V &= 4.97 \times 10^{-3} \times 22.7 = 0.113 \text{ L} \\ &= 113 \text{ mL}\end{aligned}$$

WORKED EXAMPLE 13.2.9

Ammonia (NH₃) is produced by mixing nitrogen (N₂) and hydrogen (H₂). What volume of hydrogen is required to react with an excess of nitrogen to produce 3.95 L of ammonia at STP?

ANSWER

1 Write the balanced equation for the reaction.



2 Convert volume of NH₃(g) to moles.

$$\begin{aligned}n &= \frac{V}{22.7} \\ &= \frac{3.95}{22.7} \\ &= 0.174 \text{ mol}\end{aligned}$$

3 Use the balanced equation to determine the stoichiometric ratio between H₂ and NH₃.

3 mol H₂(g) : 2 mol NH₃(g)

If there are 0.174 mol of NH₃, then:

$$\begin{aligned}\frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} &= \frac{x \text{ mol H}_2}{0.174 \text{ mol NH}_3} \\ x &= \frac{3}{2} \times 0.174 \\ &= 0.261 \text{ mol H}_2(\text{g}) \text{ required}\end{aligned}$$

4 Use Avogadro's hypothesis to determine the volume of H₂(g) required.

$$\begin{aligned}0.261 &= \frac{V}{22.7} \\ V &= 0.261 \times 22.7 = 5.92 \text{ L of H}_2(\text{g}) \text{ required}\end{aligned}$$

PRACTICAL ACTIVITY 13.2.1

ESTIMATING THE MOLAR VOLUME OF A GAS

Introduction

In section 13.1, you learnt about Avogadro's hypothesis, which states that equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles. In other words, at STP 1 mol of any gas occupies 22.7 L. If a gas-producing reaction is carried out in the laboratory, it is possible to collect the gas and measure its volume. Using Avogadro's hypothesis, the molar volume of the collected gas can be estimated.

Research question

Can the molar volume of a gas be determined experimentally under standard laboratory conditions?

Aim

To estimate the molar volume of a gas

Equipment per group

- 1.50 g powdered calcium carbonate
- 150 mL of 1 mol L⁻¹ ethanoic (acetic) acid
- 200 mL measuring cylinder
- 50 mL measuring cylinder
- boiling tube
- rubber bung fitted with glass tube
- retort stand and clamp
- electronic balance
- weighing boat or filter paper
- spatula
- pneumatic trough or large plastic tub
- water
- paper towel



What are the risks in doing this activity?

How can these risks be managed?

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your risk assessment before you proceed.

Procedure

- 1 Set up the apparatus as shown in **Figure 13.2.7**.
- 2 Place 30.00 mL of ethanoic acid in the boiling tube.
- 3 Place a weighing boat (or filter paper) on a mass balance and zero the balance (press tare).
- 4 Place about 0.10 g of calcium carbonate onto the weighing boat and accurately record the mass.
- 5 Remove the bung from the boiling tube and carefully pour the calcium carbonate into the tube.

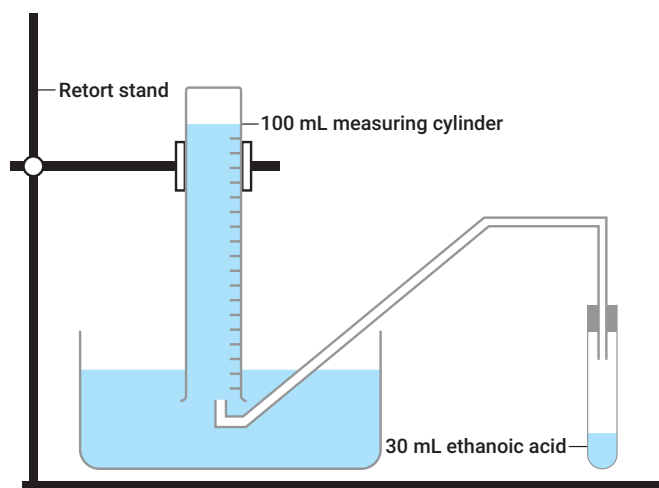


FIGURE 13.2.7 The experimental set-up for estimating the molar volume of a gas

- 6 Quickly replace the bung.
- 7 When the reaction is complete, record the volume of gas in the measuring cylinder.
- 8 Repeat the experiment four times, increasing the mass of calcium carbonate by 0.10 g each time.

Results

Copy the following table, inserting a new row for each experiment, and use it to record your results.

Experiment number	Mass of calcium carbonate (g)	Volume of carbon dioxide (mL)
1		

Analysis of results

- 1 Write a balanced equation for the reaction between calcium carbonate and ethanoic acid.
- 2 Plot a graph of the volume of carbon dioxide (y-axis) against the mass of calcium chloride.
- 3 What is the mathematical relationship of the regression line? How can the information provided by the regression line help you determine the volume of 1 mole of carbon dioxide?
- 4 Draw a line of best fit passing through the established data points.
- 5 Does your graph go through the origin? If it does, what does this suggest?
- 6 For each experiment that you carried out, determine the volume of 1 mole of carbon dioxide (the experimental molar volume).
- 7 Using the equation you recorded in step 1, determine the number of moles of carbon dioxide produced in each experiment to calculate the volume of 1 mole of carbon dioxide (the theoretical molar volume).

Interpretation

- 8 Compare the experimental molar volume of carbon dioxide produced in each experiment with the theoretical molar volume of carbon dioxide and calculate the percentage error in each case.

Evaluation

- 9 Within the experimental percentage error of your investigation, determine if your value for molar volume agrees with the theoretical value.
- 10 Identify possible sources of error and describe how these would affect the result.
- 11 For each source of error, suggest how to improve the experiment to minimise these errors.

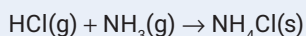
LEARNING CHECK 13.2

DESCRIBING

- 1 **List** the proposals of the kinetic theory.

APPLYING

- 2 **Explain** why an increase in temperature of a gas-filled balloon would cause it to burst.
- 3 **Explain** why the particles of hydrogen gas move faster than those of oxygen gas at the same temperature.
- 4 Hydrogen chloride gas and ammonia gas react according to the equation:



The ammonium chloride produced is present as thick white smoke.

In the experiment depicted in **Figure 13.2.8**, both gases were released at the same time. **Predict** where you think the thick white smoke will form. **Explain** your answer.

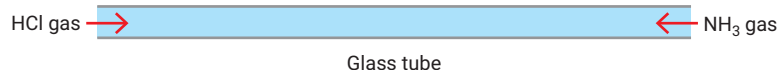


FIGURE 13.2.8 The apparatus for demonstrating the ammonia and hydrogen chloride reaction

- 5 Assuming the temperature is constant, carry out the following calculations.
- If 4.42 L of gas was collected at a pressure of 94.2 kPa, what volume would it occupy at 1 atm pressure?
 - A 210 mL sample of nitrogen gas at 1.0 atm is compressed until the final volume is 150 mL. **Determine** its final pressure.
 - When the pressure is 522 kPa, a gas has a volume of 55 mL. **Calculate** its volume at standard atmospheric pressure.
 - A sample of hydrogen gas at 73.3 kPa in a 2.5 L flask is expanded into a total volume of 7.2 L. **Calculate** the final pressure.
- 6 Assuming pressure is constant, carry out the following calculations.
- A sample of oxygen has a volume of 3.50 mL at 20°C. What volume would it occupy at 80°C?
 - If a sample of gas had a volume of 1.7 L at 25°C and 1 atm pressure, what:
 - would its volume be at 100°C?
 - temperature would it need to be cooled to have a volume of 0.80 L?
 - A gas occupies a volume of 60.0 mL at 36°C. What volume would the gas occupy at 0°C?
- 7 Assuming pressure is constant, carry out the following calculations.
- What is the volume of 0.34 mol of H₂S gas at 100°C and 75.0 kPa pressure?
 - A sample of oxygen gas is collected and found to occupy a volume of 527 mL at 1.05 atm and 20°C. What is the mass of the sample?
 - What is the pressure of 4.00 g of He in a container of volume 7.5 L at 30°C?
- 8 A mass of 4.86 g of magnesium is ignited in nitrogen dioxide. The products of the reaction are magnesium oxide and nitrogen gas. **Calculate** the volume of nitrogen gas produced at 273 K and 101.3 kPa.
- 9 The complete combustion of carbon in oxygen produces carbon dioxide. **Calculate** the volume of oxygen that would react at STP with 10.0 g of carbon, and the volume of carbon dioxide produced under the same conditions.
- 10 Hydrogen peroxide (H₂O₂) decomposes to give water and oxygen gas. How many grams of hydrogen peroxide are needed to produce 100 L of oxygen at STP?

CHAPTER SUMMARY

Molar volume

- For ideal gases, 1 mol of gas particles at STP occupies a volume of 22.7 L.
- The relationship between number of moles and volume of gas at STP can be expressed as:

$$\text{Number of moles of gas} = \frac{\text{volume of gas at STP}}{\text{volume of 1 mole of gas at STP}}$$

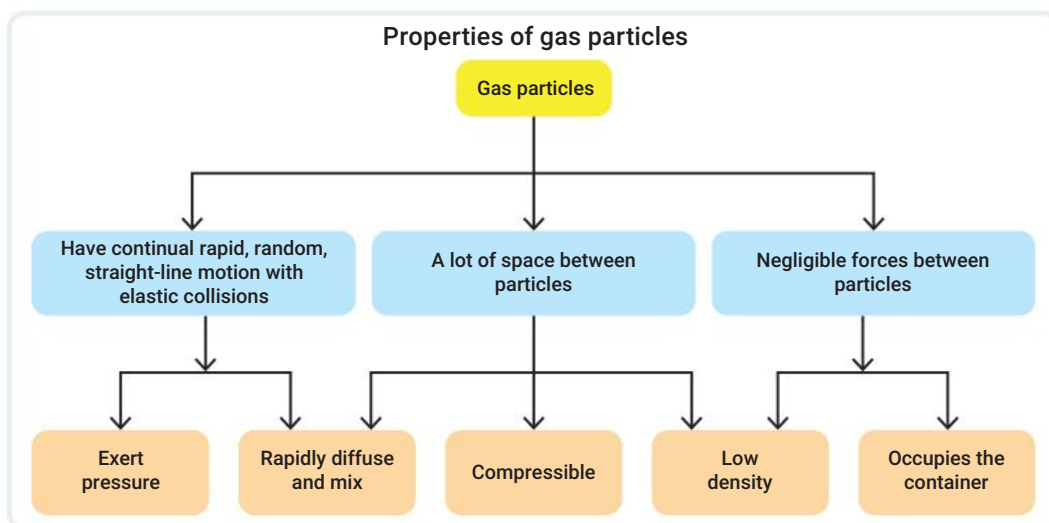
$$n = \frac{V}{22.7}$$

where: n = number of moles

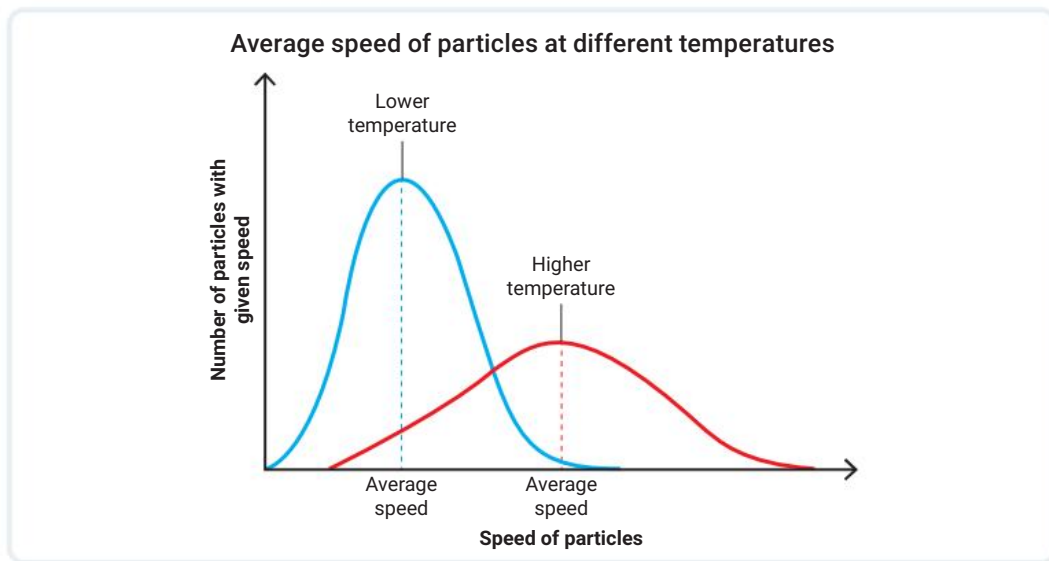
V = volume of gas (L) at STP.

Kinetic theory of gases

- The kinetic molecular theory of gases describes the behaviour of gas particles.



- As temperature increases, the kinetic energy of particles increases.

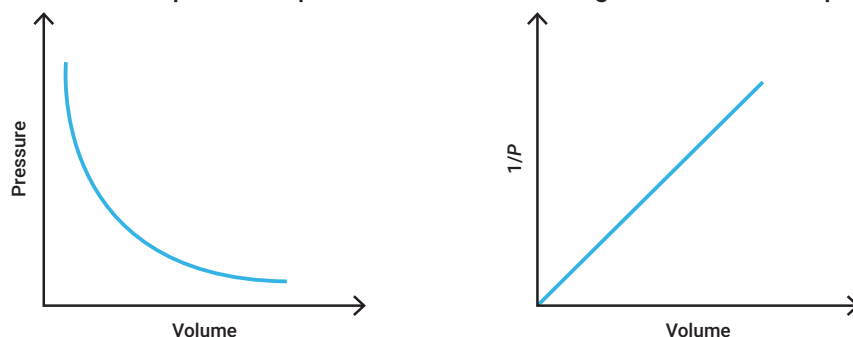


Gas laws

- According to Boyle's law, the volume of a given mass of gas is inversely proportional to the pressure
- Boyle's law can be expressed as:

$$P_1 V_1 = P_2 V_2$$

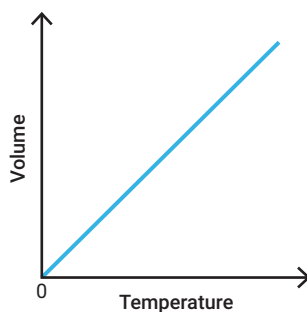
The relationship between pressure and volume of a gas at constant temperature



- According to Charles' law, for a fixed quantity of gas at a constant pressure, there is a linear relationship between volume and temperature.
- Charles' law can be expressed as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Impact of temperature on gas particles



- Combining Boyle's and Charles' laws gives us the ideal gas equation (also known as the ideal gas law or universal gas equation):

$$PV = nRT$$

where: P = pressure (kPa)

V = volume (L)

n = total number of moles of gas present

R = ideal gas constant = $8.31 \text{ J mol}^{-1} \text{ K}^{-1}$

T = temperature (K).

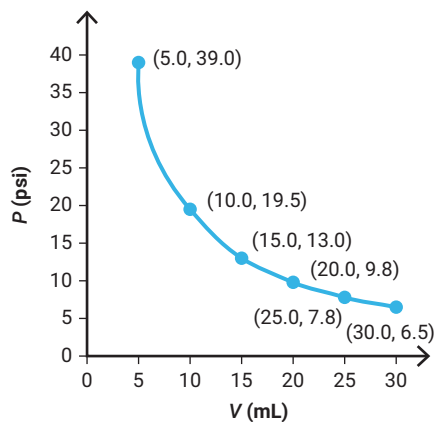
- The ideal gas equation describes the properties of all gases.

CHAPTER EXAM

MULTIPLE CHOICE

- At STP conditions:
A concentration is 177M.
B pressure is 100 kPa.
C temperature is 273°C.
D volume is 177 L.
- Which of the following is the volume of 0.10 mol of gas at STP?
A 0.227 L
B 2.3 L
C 27.0 L
D 28.2 L
- In the ideal gas equation, R represents:
A a constant.
B pressure.
C temperature.
D volume.
- Which of the following needs to be kept consistent for Charles' law to be true?
A Pressure
B Temperature
C Units
D Volume

Questions 5 and 6 refer to the following graph for a gas (1 psi = 6.89 kPa).



- The relationship demonstrated in this graph is:
A volume and pressure are equal to each other.
B an increase in pressure results in an increase in volume.
C an increase in pressure results in a decrease in volume.
D a decrease in volume results in a decrease in pressure.

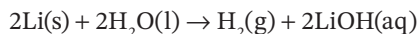
6. Which property needs to be kept constant in order for this relationship to be true?
 - A Volume
 - B Temperature
 - C Units of measurement
 - D Concentration of particles
7. A balloon with a volume of 2.0 L at 21°C was heated until its volume was 2.1 L. What was the temperature of the gas at the new volume?
 - A 36°C
 - B 39°C
 - C 41°C
 - D 45°C
8. If 60 mL of gas was cooled from 32.0°C to 5.0°C, the volume:
 - A decreased by 5.3 mL.
 - B decreased by 9.8 mL.
 - C increased by 5.1 mL.
 - D increased by 12.0 mL.
9. What is the number of moles of carbon dioxide in a 2 L container at 28°C and 142 kPa?
 - A 0.00474
 - B 0.113
 - C 1.01
 - D 1.23
10. If 0.442 mol of a gas is placed into a 1.5 L container at 17.5°C, what is the pressure of gas in kPa?
 - A 24.9
 - B 42.9
 - C 69.1
 - D 711

SHORT RESPONSE

11. A gas was placed in a container at 21°C occupying a volume of 22 mL. The gas was heated to 97°C. Assuming the pressure remains constant, **determine** the final volume of the gas.

©VCAA VCAA sample exam 2013 SAQ Q8c, modified

12. Some early lithium metal batteries exploded when exposed to water. The following reaction occurs when lithium reacts with water:

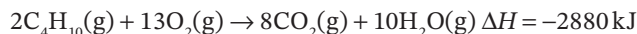


What volume, in L, of hydrogen gas at 20.0°C and 127 kPa is produced by the reaction of 1.00 g lithium with excess water?

CROSS-CHAPTER QUESTION

©VCAA VCAA VCAA 2021 MCQ Q12, modified

13. Butane (C_4H_{10}) undergoes complete combustion according to the following equation.



A sample of C_4H_{10} was used and released 3330 kJ of energy during complete combustion at standard laboratory conditions (SLC).

- a **Determine** the amount of butane used in the reaction.
- b **Calculate** the volume of carbon dioxide produced.

DATA ANALYSIS

14. Analyse data

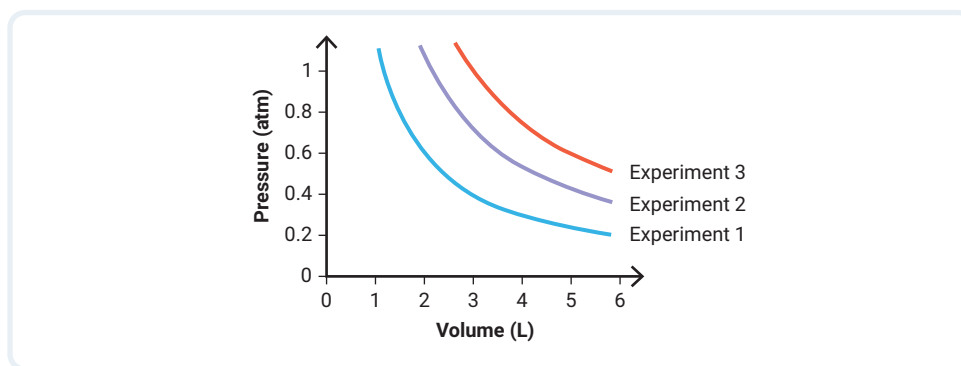
A 10 g sample of oxygen was placed in a cube container measuring $10\text{ cm} \times 10\text{ cm} \times 10\text{ cm}$ at 25°C .

Calculate the pressure of oxygen (in kPa) in the container.

15. Interpret evidence

A student conducted a series of experiments to determine the relationship between pressure and volume. The same number of particles was used in all experiments.

The following graph represents data from the experiment.



- Identify** the independent and dependent variables in this series of experiments.
- Determine** which gas law is demonstrated by the trends seen in experiment 1.
- Determine** what variable was changed between the experiments that resulted in the changes seen between the graphs.

SCIENCE AS A HUMAN ENDEAVOUR

Syllabus dot point

- Explore Jacques Cousteau and Emile Gagnan's role in the invention of SCUBA.
- Appreciate that safe scuba diving requires knowledge of the behaviour of gases.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Making sea exploration possible

Without specialised diving technology and equipment, we would not know about the beauty of coral or the existence of marine life such as sharks and dolphins, which exist on the ocean floor at shallow depths. Our bodies are not adapted to be able to dive deep towards the ocean floor. We can use our understanding of gas laws to explain why.

According to Boyle's law, there is an inverse relationship between volume and pressure. This has important implications when diving. As divers go deeper, pressure progressively increases, which can result in:

- the air spaces in the body, including the lungs, compressing – this makes it feel as though you are being squeezed and the air volume in the lungs decreases
- nitrogen narcosis – nitrogen becomes more soluble than oxygen in water (in the blood). Because air is mostly nitrogen, this results in high levels of nitrogen in the body, which can disrupt the function of the brain, causing dizziness, disorientation and impaired judgement.

Divers also need to be just as careful when they ascend as when they descend. The decrease in pressure that occurs when they resurface can result in:

- the air in the lungs and sinuses expanding – if this happens too quickly, the lungs can overinflate and cause the tissue to rupture
- decompression sickness – if divers ascend too quickly, there is insufficient time for nitrogen in the blood to be released, causing bubbling in the blood and tissues. This causes decompression sickness and can result in pain, paralysis and even sometimes death.

Using our understanding of gas laws, we have been able to develop equipment and technologies to enable us to explore the ecosystems that exist deep in the ocean.

Invention of diving equipment

Scuba diving has become a popular water activity around the world thanks to the invention of the self-contained underwater breathing apparatus (scuba). One of the most significant contributions was the invention of the Aqua-Lung.

In the 1940s, a collaboration between engineer Émile Gagnan, an expert in pressure regulators, and Jacques Cousteau, a naval officer with a passion for underwater exploration, led to a breakthrough invention.

Together, they were able to adapt available regulators for underwater use. The regulator is linked to air tanks and helps to deliver compressed air on demand,



FIGURE 1 Scuba diving on the Great Barrier Reef is now common thanks to the work of Émile Gagnan and Jacques Cousteau.

Islandjems - Jemma Craig/Shutterstock.com



taking into account the changes in pressure in a person's lung. It helps to vary the pressure of the air that is being supplied so that it is safe to breathe. This game-changing piece of innovation resulted in the development of the Aqua-Lung (Figure 2).

J. Y. Cousteau et al. (1949). Diving unit, US2485039A, US Patent.

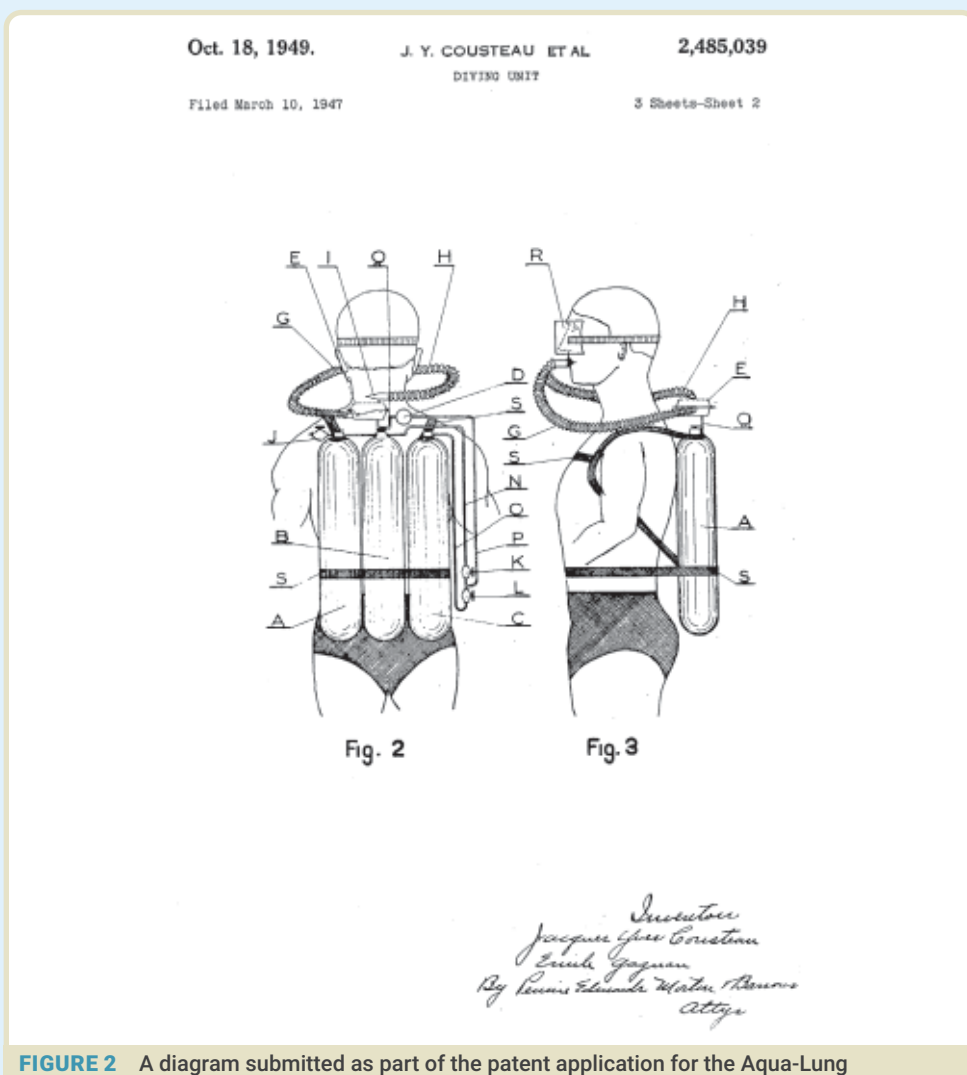


FIGURE 2 A diagram submitted as part of the patent application for the Aqua-Lung

While not the first scuba, the Aqua-Lung revolutionised underwater exploration by being the first commercially successful and widely adopted design. This paved the way for modern scuba equipment, all thanks to our understanding of gas laws.

Aqueous solutions and molarity



Krakenimages.com/Shutterstock.com

SYLLABUS DOT POINTS

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* and *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) = $\frac{\text{moles of solute } (n)}{\text{volume of solution } (V)}$)

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

The water molecule is small but mighty, having unique physical and chemical properties that are important for many processes on Earth. For example, water is the solvent for a range of influential aqueous solutions and exists in solid, liquid and gaseous forms. Unlike most solids that expand when they melt, water expands when it freezes. These unique properties arise from the strength of the intermolecular forces between the water molecules and their interactions with other substances. Understanding the chemistry of the water molecule unlocks the secret to these unique characteristics.

Practical

- Preparing and diluting solutions

Worksheets

- Different measures of concentration
- Calculating unknowns quantities
- Problems involving molarity



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Electronegativity measures an atom's ability to attract bonding electrons towards itself.
- ✓ Depending on the atoms involved, there are different types of bonds that hold molecules and compounds together.
- ✓ VSEPR theory can be used to determine the shape and polarity of molecules.
- ✓ Physical properties of substances depend on the strength of their intermolecular forces.
- ✓ The relationship between moles and mass can be represented as $n = \frac{m}{M}$.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ describe water as a bent molecule, containing polar bonds, which affects the unique properties that water exhibits
- ✓ use the molecular polarity and the associated strong intermolecular forces (hydrogen bonding) between the small water molecules to explain why water has unique properties
- ✓ identify the key differences between solutions, solutes and solvents
- ✓ identify the key differences between unsaturated, saturated and supersaturated solutions
- ✓ describe the concentration as a measure of the amount of solute dissolved in a fixed volume of solution
- ✓ calculate the molarity (M) of a solution using the formula: $c = \frac{n}{V}$.
- ✓ calculate the molarity or volume of dilutions using $c_1V_1 = c_2V_2$. A solution can be diluted by adding more solvent.

14.1 The structure and properties of water



FIGURE 14.1.1 Ice floats in water

Water is a tasteless, odourless liquid that is essential to our survival. All organisms are at least 60% water. Although water is very common, its particular properties make it an unusual material when compared to other molecular structures. These properties have enabled life to evolve on Earth and explain why water is vital for life.

Water freezes at 0°C and boils at 100°C at standard pressure. It can act as a solvent over this entire temperature range, dissolving and transporting materials across a cell or over the oceans. It is known as the 'universal solvent' for its ability to dissolve more substances than any other liquid. It is the only substance on Earth that exists naturally in its three states of matter: solid, liquid and gas. Unlike other substances, solid water is less dense than its liquid form. This explains why ice floats, while most other solids sink when mixed with their liquid form (Figure 14.1.1). Given that water is such a small molecule and common substance, why is it considered to be unique?

Water has some unique properties that result from its bonding characteristics both within the molecule and between other molecules, including:

- unusual melting and boiling points
- unusually high heat capacity
- density in solid and liquid phases
- surface tension
- an ability to act as a very good solvent.

These properties will be discussed in further detail throughout the chapter.

Boiling point, **density** in solid and liquid phases, **surface tension** and ability to act as a solvent are the unique properties of water. These properties are due to the water molecule's shape, bonding within the molecule, and the intermolecular forces it experiences between molecules. Ultimately, it is these chemical forces that make water vital in biological, chemical and physical processes on Earth.

density the mass per unit volume of a substance

surface tension the tension of the surface film of a liquid, due to the attraction between particles in the surface layer

The structure of water

Shape of the water molecule

As you know, each water molecule (H_2O) consists of two hydrogen atoms each covalently bonded to an oxygen atom, which also has two lone pairs of electrons. You should recall from Chapter 11 that the shape of the water molecule is largely determined by the valence electrons, including both bonded pairs of electrons as well as both lone pairs. Interactions between the lone pairs and the bonding electrons cause water to have a bent structure (**Figure 14.1.2**).

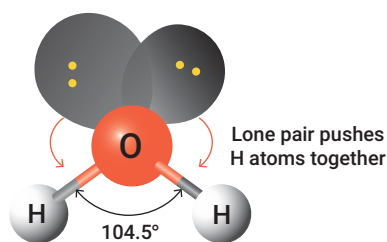


FIGURE 14.1.2 Oxygen atoms in water molecules have lone pairs that affect the shape of the molecule.

O—H bonds within the water molecule

The polarity of the bonds in water can be determined by examining the electronegativities of hydrogen (H) and oxygen (O) in the *Formula and data book*. When comparing the two atoms, oxygen has an electronegativity of 3.4 and hydrogen has an electronegativity of 2.2. This means the oxygen atom is more likely to attract the valence electrons. In an O—H bond, the electrons are drawn closer to the oxygen atom than the hydrogen atom. As a result, oxygen is said to have a partial negative charge (δ^-) and hydrogen is said to have a partial positive charge (δ^+). This results in a polar covalent bond between the oxygen and hydrogen atoms (**Figure 14.1.3**).

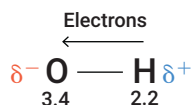


FIGURE 14.1.3 The partial charges of a polar covalent O—H bond

Polarity of the water molecule

Once the polarity of the bonds *within* the water molecule have been revealed, the overall polarity of the molecule is determined by considering its shape. Water is a bent molecule that has strong polar covalent bonds. The oxygen attracts the negative electrons, and gains a partial negative charge. The hydrogen atoms therefore attain a partial positive charge (Figure 14.1.4). The presence of these partial charges, in combination with the bent shape, gives the water molecule (H_2O) its overall molecular polarity as a positive and negative dipole are formed.

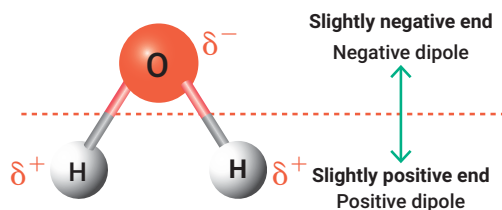


FIGURE 14.1.4 The overall polarity of a water molecule

Intermolecular forces in water

Intermolecular forces – the forces that act *between* molecules – are of great importance in explaining many of the unique physical and chemical properties of water. Strong covalent bonds exist within the water molecule. The differences in electronegativity of hydrogen and oxygen and the bent shape of the molecule result in it having a molecular dipole. Molecular polarity affects the type of forces *between* molecules. There are all three types of intermolecular forces in existence in water: dispersion forces, dipole–dipole attractions and hydrogen bonding. Of these three, hydrogen bonding is the strongest type of intermolecular interaction and accounts for many of the properties of water (Figure 14.1.5).

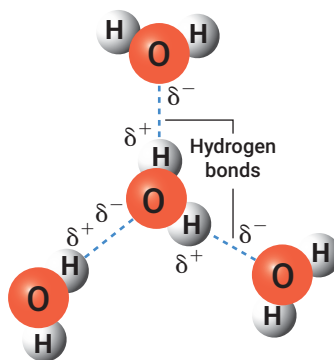


FIGURE 14.1.5 Hydrogen bonding between water molecules

Water's unique properties

Boiling point and vapour pressure

Hydrogen bonding explains all the properties that relate to temperature. A relatively large amount of energy is required to break the intermolecular forces between water molecules, especially since water molecules are so small compared to the strength of the intermolecular forces between them. This explains water's relatively high **boiling point** of 100°C at a pressure



Syllabus link

Chapter 11 covered the three types of intermolecular forces in detail.

boiling point the temperature at which bubbles of gaseous vapour are formed within the liquid.

of 1 atmosphere. Other hydrides of similar molecular structure to water, such as H_2S , H_2Se and H_2Te , have much lower boiling points (Figure 14.1.6). This is because they lack hydrogen bonding intermolecular forces holding them together in the liquid form.

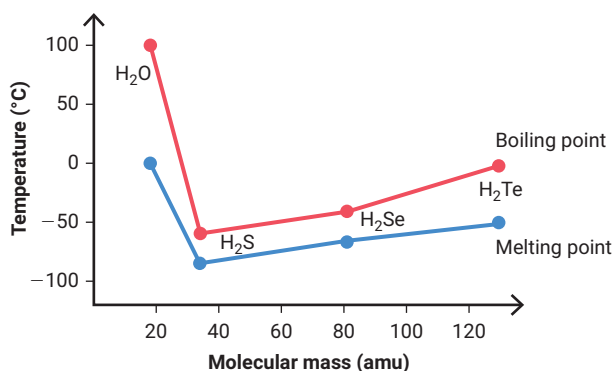


FIGURE 14.1.6 Comparative melting and boiling points of hydrides with a similar structure to water (at 1 atm pressure)

Water's strong hydrogen bonds also afford it a high latent heat of vaporisation. To increase the average temperature of water by as little as 1°C requires more heat energy than for almost any other liquid.

At room temperature, fewer molecules of water escape into the air than other molecules of similar structure. This is because the energy available at room temperature is not enough to overcome the intermolecular forces between water molecules. There is a negative correlation between intermolecular force strength and vapour pressure. Compounds with strong intermolecular forces have low vapour pressure, because they will have fewer vaporised molecules at a given temperature. Fewer vaporised water molecules means lower vapour pressure.

Density of ice

As with all substances, as water cools, its particles slow down and pack closer together, resulting in an increase in density. This explains why cooler water has a higher density than warmer water. However, water has the unique property of expanding as it freezes. When the temperature is less than 4°C , the molecules do not have as much kinetic energy and no longer move between the spaces, but form more hydrogen bonds, as shown in Figure 14.1.7. The shape of the water molecule means a regular hexagonal shape is formed when ice crystallises to form a solid, as shown in Figure 14.1.7. To form this regular shape, the water molecules must spread apart. Solid ice is about 8% less dense than water. This explains why solid ice floats on liquid water. This attribute of water helps prevent lakes from completely freezing as the lower parts are protected by the layer of ice floating above.



Weblink
Density of ice

Surface tension and capillary action

Have you ever seen insects walk on water? This is possible because of water's high surface tension, which is a result of water's strong hydrogen bonding. The water molecules beneath the surface are attracted in all directions by other water molecules. The molecules at the surface are pulled together and into the main body of water (Figure 14.1.8). This results in a tight arrangement of molecules at the surface. This property allows many small aquatic insects to walk on the surface of water, and causes water to form a spherical bead (Figure 14.1.9).

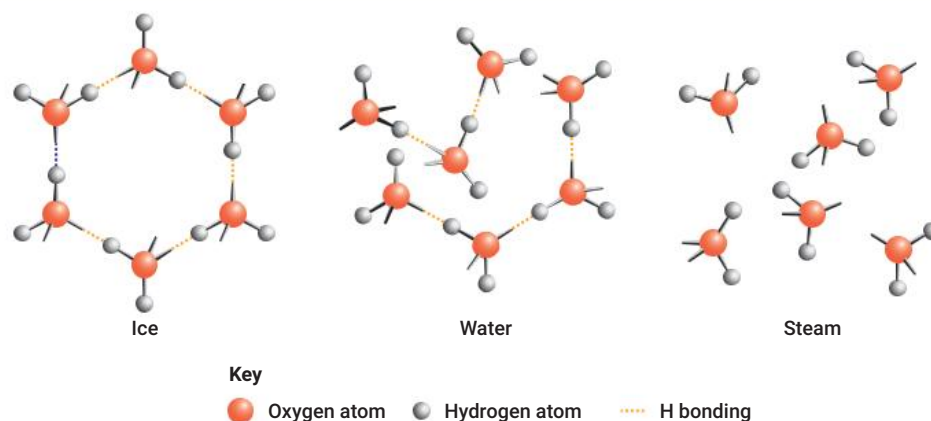


FIGURE 14.1.7 The structure of ice compared to that of liquid water and water vapour showing hydrogen bonding

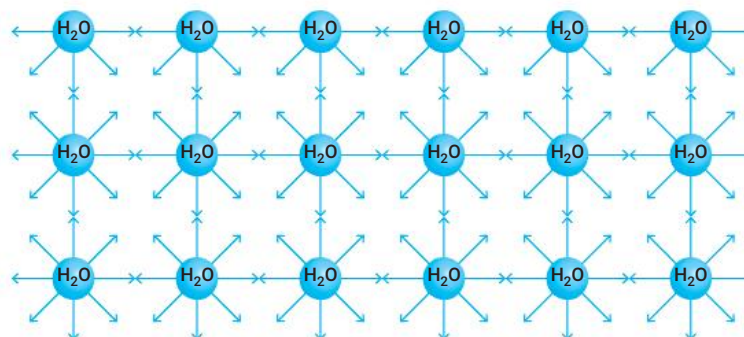


FIGURE 14.1.8 In the bulk of liquid water, each molecule has forces operating in all directions. However, on the surface, there is an imbalance – the liquid molecules are attracted to each other and exert a net force that pulls them together.

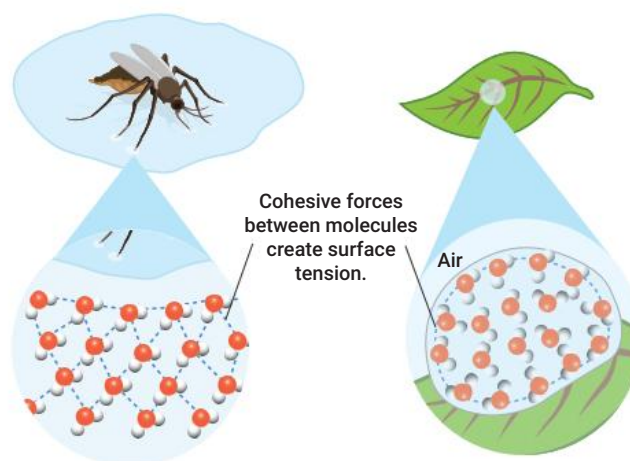


FIGURE 14.1.9 Cohesive forces between water molecules cause surface tension and water forms beads.

Capillary action is a surface tension phenomenon where a liquid can ‘climb’ up a thin tube, or capillary. It relies on the **adhesive forces** of water molecules to attract to the sides of the thin tube, and the **cohesive forces** of water to attract to other water molecules. If the tube is made of a polar substance, the water molecules have more intermolecular attraction and therefore travel higher than in tubes made of a non-polar substance.

Capillary action can be seen when part of a towel is dipped in water – the whole towel eventually gets soaked as the water ‘climbs’ up the towel fibres. Capillary action can also be seen in narrow glass vessels, where adhesive forces between water molecules and the surface of the glass and cohesive forces in water cause the surface of the water to curve to form a **meniscus** (Figure 14.1.10). The water continues to be drawn up, until gravity overcomes the adhesive force. Capillary action allows dissolved nutrients to be transported through thin tubular cells in plants to their top leaves, even in tall trees.

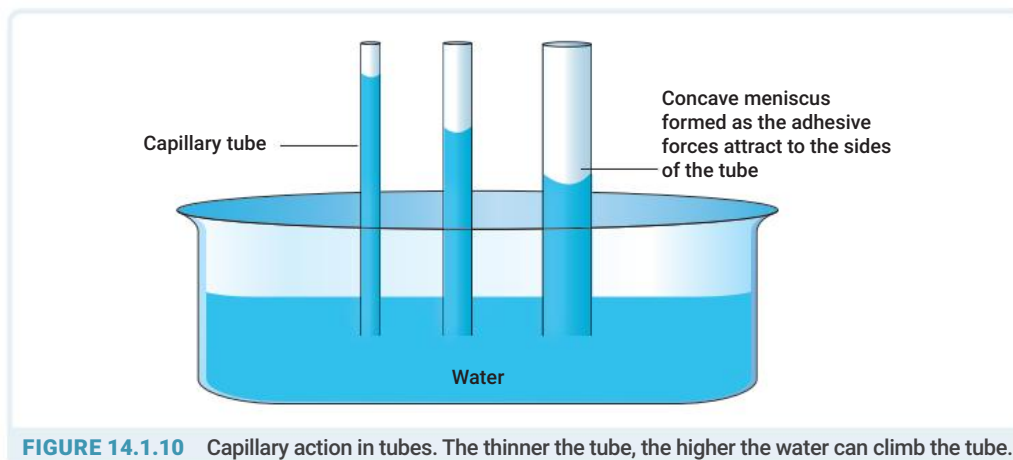


FIGURE 14.1.10 Capillary action in tubes. The thinner the tube, the higher the water can climb the tube.

Solvent ability of water

Another property of water is its ability to act as a solvent. Pure water is liquid from 0°C to 100°C. This means that it can act as a solvent over this temperature range. Water dissolves and transports a range of materials around the world and through each cell of living organisms. The polar nature of the water molecule is largely responsible for its ability to dissolve a wide range of substances. Water can dissolve ionic compounds, such as sodium chloride, and polar covalent compounds, such as sugar. Non-polar compounds, such as oil, do not dissolve in water. When considering the solubility of substances, the general rule of ‘like dissolves like’ applies. Water is polar and can dissolve other polar substances. It does this by forming hydrogen bonds or dipole–dipole bonds with polar covalent molecules or ion–dipole bonds with the ions of ionic compounds. Non-polar substances are insoluble in water because they are not attracted to the water molecules.

Water of crystallisation

Water of crystallisation occurs when water molecules are attracted to the ions of a salt. Many ionic compounds that crystallise out of an aqueous solution include water molecules in a regular array through the crystal lattice. The water molecules are in set ratio to the ionic formula unit. These compounds are called ‘**hydrated**’. For example, hydrated copper(II) sulfate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$) has five water molecules of crystallisation per formula unit, as shown in Figure 14.1.11.

The water of crystallisation can be evaporated by heating the hydrated compound. When all the water molecules have been removed, the compound is said to be ‘**anhydrous**’. To do this, a known mass of the hydrated compound is heated so that the mass of water removed can be determined. From this, the empirical formula of the hydrate can be calculated.

capillary action the ability of a liquid to flow through narrow spaces in opposition to gravity

adhesive force a force of attraction between unlike molecules

cohesive force a force of attraction between like molecules in a substance that helps to maintain a certain shape of the liquid

meniscus the curved shape formed by the surface of a liquid in contact with a cylinder or tube



Syllabus link
Chapter 16 discusses the ability of water to dissolve substances in greater depth.

water of crystallisation water molecules present in a fixed ratio that form an essential part of the crystal structure of some ionic compounds

hydrated ionic crystals with a specific number of water molecules in their chemical formulas

anhydrous containing no water, particularly of ionic solids

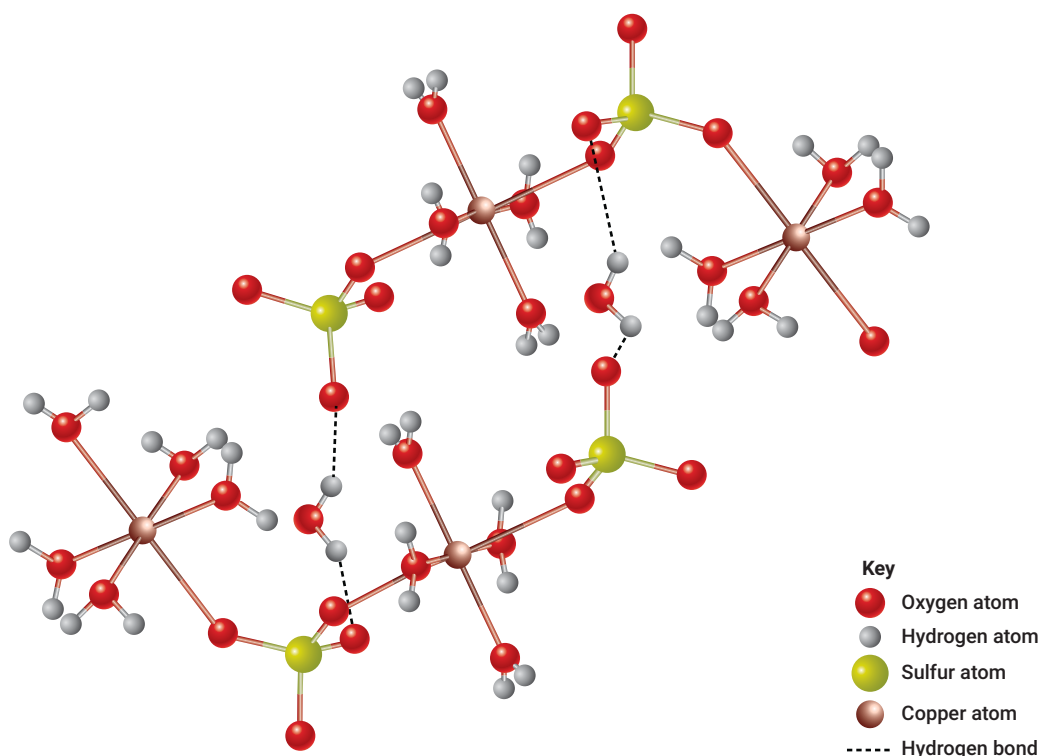


FIGURE 14.1.11 The crystals of copper(II) sulfate consist of an octahedral arrangement of water molecules and sulfate ions around the central copper ion.

LEARNING CHECK 14.1

DESCRIBING

- 1 **Define:**
 - a density
 - b surface tension
 - c vapour pressure
 - d capillary action
 - e water of crystallisation.
- 2 **Identify** three reasons why water can be considered an unusual substance, despite it being a very common molecule.
- 3 **Explain** how hydrogen bonding between water molecules can cause the phenomenon of surface tension.
- 4 **Explain** why water has a low vapour pressure at room temperature.

APPLYING

- 5 **Explain** why you should not put full water bottles in the freezer.
- 6 **Discuss**, with the aid of diagrams, how the electronegativity of atoms *within* the water molecule relates to the strength of the intermolecular forces *between* water molecules.

- 7 **Figure 14.1.12** shows the boiling points of four hydrides of group 16 elements at 1 atmosphere of pressure.

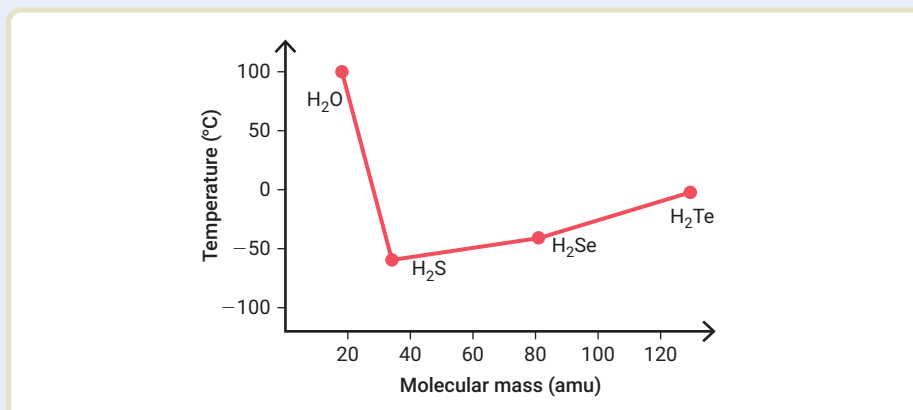


FIGURE 14.1.12 Boiling points of some group 16 hydrides

- Identify** the boiling point of H₂O.
- Classify** whether water aligns with the trend shown by the other hydrides, or if it is an anomaly.
- Explain** your answer to part **b**, using diagrams of the water molecule.

ANALYSING

- A sample of copper(II) sulfate was tested to determine whether it was anhydrous. This was done by repeatedly heating, cooling and weighing the sample. **Infer** how the results of this experiment could be used to determine an outcome.
- The tallest trees in the world are redwoods, found on the western coast of the United States. Water is transported up the trunk and stems in tubes called xylem. A particularly notable feature of the redwood trees is that the further up the tree they are, the thinner the xylem become. **Deduce** how this phenomenon might enable water to be transported up to the topmost leaves.
- A sample weighing 6.25 g of blue hydrated copper(II) sulfate, CuSO₄ · xH₂O, (x unknown) is gently heated in a crucible until the mass remaining is a constant 4.00 g. **Determine** the formula of the hydrated salt (i.e. the value of x in CuSO₄ · xH₂O).

14.2 Solutes, solvents and solutions

Water is considered the universal **solvent** and plays a crucial role as it dissolves and carries valuable minerals, nutrients and chemicals that are essential for life. Its ability to dissolve many different types of **solute**s is mainly due to its polarity. The **solution** that is formed from a solute dissolving in a solvent is homogenous as the dissolved particles are too small to be seen. For example, sugar dissolved in water still appears to be clear, but the sweet taste reveals that there are sugar molecules in solution.

When water dissolves a solute, the water molecules are attracted to the solute particles (molecules or ions) of the solute. This causes the solute particles to separate from each other and instead interact with the water molecules. Since water is polar, substances that may dissolve are those that are polar or charged, by forming ion–dipole, dipole–dipole or hydrogen bonds with water molecules. This follows the general rule of ‘like dissolves like’ where polar solvents are attracted to polar solutes and non-polar solvents are attracted to non-polar solutes. However, it is not always this simple.

solvent a substance (usually a liquid) in which another substance is dissolved

solute a substance that is dissolved in a solvent to form a solution

solution a homogeneous mixture with a solute dissolved in a solvent, where the particles of the solute cannot be observed

soluble capable of being dissolved in a solvent

solubility the amount of solute that can dissolve in 100 g water (solvent) at a particular temperature (100 g water = 100 mL water)

insoluble incapable of being dissolved or very poorly soluble in a solvent

Not all polar molecules and ionic substances are completely **soluble** in water. **Solubility** depends on the physical and chemical properties of both the solute and the solvent. The strength of the attractions between the solvent and solute particles, temperature, molecule size and pressure can play a part in the solubility of a substance. The smaller and more polar a solute is, the more likely it will dissolve in water. Solubility will be discussed in greater detail in Chapter 16.

LEARNING CHECK 14.2

DESCRIBING

- 1 **Identify** the term used to describe the maximum amount of solute that can dissolve at a particular temperature in 100 g of water.
A Solvent
B Solubility
- 2 **Explain** why the solvent properties of water are so important to the existence of life.

APPLYING

- 3 **Explain** why ionic substances such as salt dissolve in water.
- 4 **Differentiate** between the solute and the solvent in an aqueous solution of sodium chloride.
- 5 Despite being an ionic substance, barium sulfate does not actually dissolve in water. **Explain** why this might be so.

concentration a measure of the amount of solute dissolved in a given amount of solvent

molarity the number of moles of solute per litre of solution

14.3 Concentration

Aqueous solutions can contain many different dissolved substances. The amount of dissolved solute is regularly monitored in many areas of chemistry, including wine production, water quality and swimming pools. The **concentration** of a solution allows us to obtain a measurement of how much solute is dissolved in a particular volume of solvent.

There are various units used to measure concentration, all of which can be summarised as:

$$\text{concentration} = \frac{\text{amount of solute}}{\text{volume of solution}}$$

$$\text{or } c = \frac{n}{V}$$

The specific unit used depends on the context in which the measuring is carried out. One of the most common units of concentration in chemistry is **molarity**. Molarity is also referred to as molar concentration and is given the symbol M. The molarity of a solution is defined as the number of moles of solute per litre of solution. For example, 1 M represents 1 mol of solute dissolved per litre (L) of solution.

Scientists use square brackets [] when referring to the molarity of a substance. For example, $[\text{CuSO}_4] = 0.1 \text{ M}$ means 'the concentration of the copper sulfate solution is 0.1 moles per litre.

A wide array of other concentration units are also used to express concentration. For example, m/v (mass of solute to volume of solution), m/m (mass of solute to mass of solution), or v/v (volume to volume) may be used. Lower concentrations are often measured in parts per

KEY FORMULA

The molarity of a particular solution is given by the formula:

$$c = \frac{n}{V}$$

where: c = concentration (M)
 n = number of moles
 V = volume (L).



Worksheet

Different measures of concentration

million (ppm) or parts per billion (ppb). **Table 14.3.1** lists some common units of concentration with some examples of their context.

TABLE 14.3.1 Some commonly used units of concentration

Unit of concentration	Definition	Context
molarity M mol/L mol L ⁻¹	Number of moles of solute per litre of solution	The concentration of a solution of hydrochloric acid was 0.1 M.
g/L g L ⁻¹	Number of grams of solute per litre of solution	The concentration of a solution of sugar syrup was 855 g L ⁻¹ .
mol/dm ³ mol dm ⁻³	Number of moles per cubic decimetre (1 dm ³ is equivalent to 1 L)	The concentration of a solution of sulfuric acid was 5.0 mol dm ⁻³ of solution.
g/100 g	Number of grams of solute per 100 g of solvent	The maximum solubility of sodium chloride at 20°C is 25 g of NaCl per 100 g of water.
% v/v	Percentage by volume	The alcohol content of a bottle of wine might be 12.3% v/v, i.e. there is 12.3 mL of alcohol in every 100 mL of wine.
% w/w	Percentage by mass	The concentration of sugar in a solution is 25% w/w, i.e. there is 25 g of sugar in every 100 g of solution.
% w/v	Percentage of mass by volume of solution	The concentration of potassium chloride is 5% w/v, i.e. there is 5 g of potassium chloride in every 100 mL of solution.
ppm	Parts of solute per one million parts of the solution For example, 1 ppm can be: 1 mg L ⁻¹ 1 mg kg ⁻¹	The amount of fluoride in a city's water supply was 80 ppm.

There are some useful conversions to remember when working with concentrations. For example, pure water has a density of 1 g mL⁻¹, so the mass of water in grams is always numerically equal to the volume in millilitres. For example, 100 mL of pure water has a mass of 100 g. **Table 14.3.2** shows some other useful conversions relating to water.

TABLE 14.3.2 Some useful conversions relating to water

Unit of water	Volume conversion
1 gram	1 mL
1 cubic centimetre (cm ³)	1 mL
1 kilogram (kg)	1 L
1 decimetre (dm ³)	1 L



Weblink
Concentration

WORKED EXAMPLE 14.3.1

- a A hydrochloric solution has a molarity of 2 M. Determine how many moles of hydrochloric acid are in 1 L of solution.
- b Red wine has an alcohol content of approximately 13% v/v. In a 750 mL bottle, what volume of alcohol is present?
- c A hair sample was found to have a copper level of 41 ppm. It should be less than 2.5 mg % w/w. Is the amount in the hair an acceptable level?

ANSWERS

- a **1 Rearrange the equation $c = \frac{n}{V}$ to make n the subject.**
 $n = c \times V$
- 2 Substitute the values to determine the amount of solute dissolved.**
 $n = c \times V$
 $= 2 \text{ M} \times 1 \text{ L}$
 $= 2 \text{ mol of HCl are dissolved}$
In a 2 M solution, there are 2 mol of hydrochloric acid in 1 L of solution.
- b **1 Identify the process involved in calculating total alcohol content.**
Actual volume of alcohol = total volume $\times$ percentage alcohol
- 2 Substitute the values to determine the volume of alcohol present.**
 $V = 750 \text{ mL} \times \frac{13 \text{ mL}}{100 \text{ mL}}$
 $= 97.5 \text{ mL}$
750 mL of red wine contains 97.5 mL of alcohol.
- c **1 Convert the copper level in hair from parts per million to an amount in milligrams.**
According to the conversion, 1 ppm = 1 mg kg⁻¹
41 ppm = 41 mg per kg
- 2 Calculate the % w/w.**
The legal limit has the % w/w; therefore, we need to find the mass per 100 g.
1 kg = 1000 g
 $= 41 \text{ mg}/1000 \text{ g}$
 $= 4.1 \text{ mg}/100 \text{ g}$
Therefore there is 4.1 mg %w/w.
The amount of copper in the hair is over the accepted value.

We can describe solutions as being concentrated or dilute. This depends on the amount of solute that is dissolved in a unit of volume of solvent/solution (**Figure 14.3.1**). For example, a

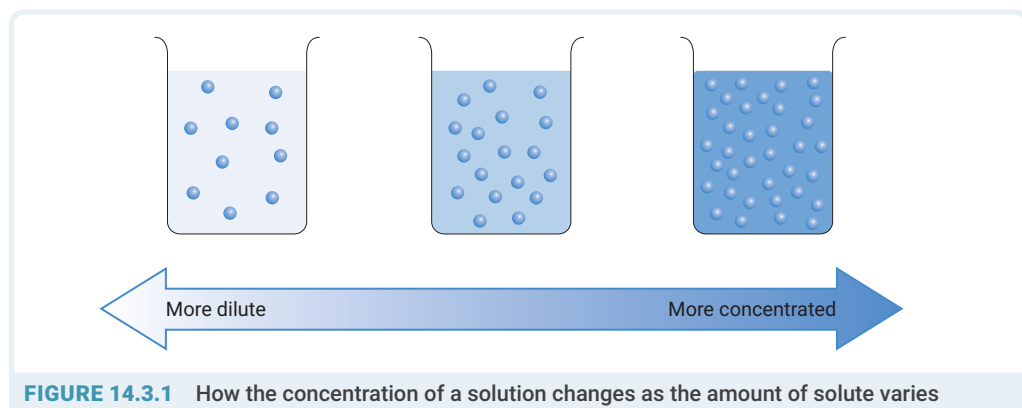


FIGURE 14.3.1 How the concentration of a solution changes as the amount of solute varies

0.1 M solution of calcium carbonate has 0.1 mol of calcium carbonate solute per litre of solution, whereas a 1.0 M solution of calcium carbonate has 1.0 mol of calcium carbonate solute per litre of solution. Dilution will be discussed further in the next section.

It is important not to confuse the concept of concentration and strength. While concentration refers to the amount of solute dissolved, strength describes how much of an acid has ionised (separated into ions) or how much of a base dissociated in the solution. For example, in an acetic acid solution (CH_3COOH), only a few of the acid molecules ionise into H^+ and CH_3COO^- ions in the solution, whereas in a hydrochloric acid solution (HCl), all of the acid molecules ionise into H^+ and Cl^- ions in the solution. In this case, acetic acid is described as a weak acid whereas hydrochloric acid is described as a strong acid even though their concentrations are the same (**Figure 14.3.2**). You will learn more about strengths of acids and bases in Chapter 18.

Since concentration and strength describe two different (but related) concepts, solutions can be described by both terms. For example, you can have a dilute solution of a strong acid and a concentrated solution of a weak acid.



Syllabus link
Chapter 18 describes the strengths of acids and bases in greater detail.

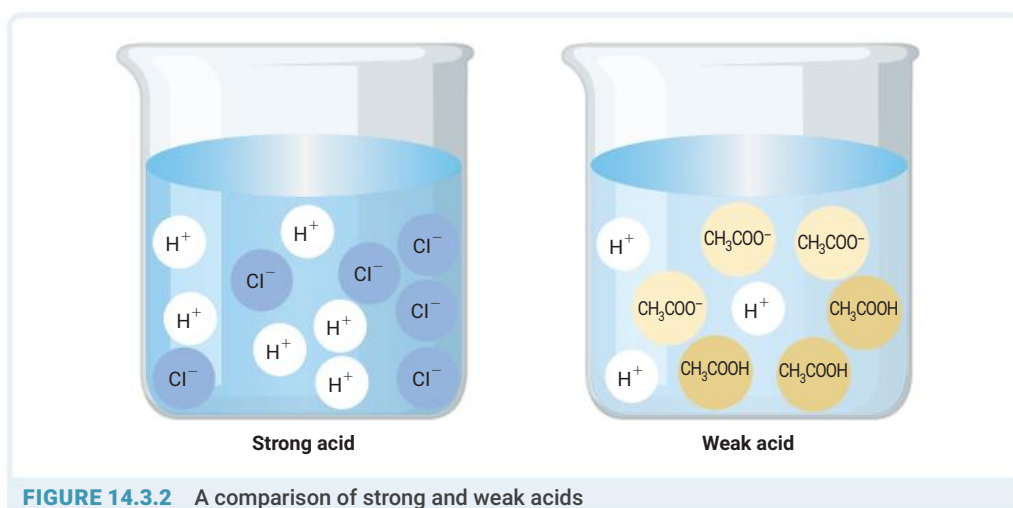


FIGURE 14.3.2 A comparison of strong and weak acids

LEARNING CHECK 14.3

DESCRIBING

- 1 A student stated the following:
'This 0.15 M solution of sodium chloride solution is more concentrated than the 0.20 M potassium chloride solution.'
Identify the error in this statement.
- 2 **Compare** the meanings and use of the terms 'molarity' and 'concentration'.
- 3 **Describe** the difference between concentrated and strong when applied to solutions.

APPLYING

- 4 **Calculate** the following concentrations.
 - a 3 mol of HCl in 1.5 L of water
 - b 2.6 mol of copper sulfate in 500 mL water
 - c 5.3×10^{-2} mol of sodium ions in 250 mL water
 - d 2.7×10^{-3} mol of cadmium ions in 5 L of water



5 Calculate the concentration of the following solutions in:

- i mol L⁻¹
- ii g L⁻¹.

- a 2.00 L of solution containing 1.50 mol of potassium bromide
- b 250 mL of solution containing 0.0025 mol of sodium iodide

6 Calculate the mass of sugar present in 200 mL of a sports drink, if it contains 6% w/v of sugar.

7 A student was asked to compare the concentrations of a 250 mL solution containing 2.1 mol of acetic acid and a solution in which 1.42 g of potassium chloride dissolved in 500 mL. **Determine** which is the more concentrated solution. Include calculations in your answer.



Syllabus link
Chapter 9 explores the relationship between moles and mass.



Weblink
Molarity

14.4 Dilutions and reacting quantities

Concentration is not the only quantity that needs to be calculated for a solution. The formula used to calculate molarity can also be rearranged appropriately to find n (given c and V) or V (given n and c):

$$n = cV$$

$$V = \frac{n}{c}$$

Sometimes, chemists may need to know the mass of solute in a solution so units of g L⁻¹ may also be used. Recall from Chapter 9 that there is a relationship between moles and mass:

$$n = \frac{m}{M}$$

Worked example 14.4.1 shows how moles, concentration and mass are calculated.

WORKED EXAMPLE 14.4.1

- a Calculate the concentration of copper(II) sulfate in a solution if 250 mL of the solution contains 0.010 mol of copper(II) sulfate.
- b Calculate the concentration of the copper(II) sulfate solution in part a in g L⁻¹.
- c Calculate the concentration of a solution prepared by dissolving 13.5 g of sodium hydroxide in a 100 mL of solution.
- d Determine the mass of silver nitrate needed to prepare 250 mL of a 0.120 mol L⁻¹ solution of silver nitrate.

ANSWERS

- a **1 Write the relationship.**

$$c = \frac{n}{V}$$

- 2 Identify the quantities and check units.**

$$c = ?, n = 0.010 \text{ mol}, V = 250 \text{ mL} = 0.25 \text{ L}$$

- 3 Substitute to find c .**

$$c = \frac{0.010}{0.25}$$

$$= 0.040 \text{ mol L}^{-1}$$

Concentration is 0.040 mol L⁻¹.



- b 1 Convert moles to grams using $n = \frac{m}{M}$.**

$$n = 0.040 \text{ mol}, M(\text{CuSO}_4) = 63.55 + 32.06 + 4 \times 16.0 = 159.6 \text{ g mol}^{-1}$$

$$0.040 = \frac{m}{159.6}$$

$$m = 0.040 \times 159.6 = 6.4 \text{ g}$$

- 2 Express the answer in g L^{-1} .**

$$V = 250 \text{ mL} = 0.250 \text{ L}$$

$$c = \frac{6.4}{0.250}$$

$$= 25.6 \text{ g L}^{-1}$$

Concentration is 25.6 g L^{-1} .

- c 1 Write the relationships.**

$$c = \frac{n}{V} \quad n = \frac{m}{M}$$

- 2 Identify the quantities and check units.**

$$c = ?, n = ?, V = 100 \text{ mL} = 0.100 \text{ L}, m = 13.5 \text{ g}$$

Calculate M and substitute to find n :

$$M(\text{NaOH}) = 40.0 \text{ g mol}^{-1}$$

$$n = \frac{13.5}{40.0}$$

$$= 0.3375 \text{ mol}$$

- 3 Substitute to find c .**

$$c(\text{NaOH}) = \frac{0.3375}{0.100}$$

$$= 3.38 \text{ mol L}^{-1}$$

Concentration is 3.38 mol L^{-1} .

- d 1 Write the relationships.**

$$c = \frac{n}{V} \quad n = \frac{m}{M}$$

- 2 Identify the quantities and check units.**

$$c = 0.120 \text{ mol L}^{-1}, n = ?, V = 250 \text{ mL} = 0.25 \text{ L}, m = ?$$

- 3 Substitute to find n .**

$$0.120 = \frac{n}{0.25}$$

$$n = 0.120 \times 0.25$$

$$= 0.030 \text{ mol}$$

- 4 Calculate M and substitute to find m .**

$$M(\text{AgNO}_3) = 169.9 \text{ g mol}^{-1}$$

$$0.030 = \frac{m}{169.9}$$

$$m = 0.030 \times 169.9$$

$$= 5.097 \text{ g}$$

Mass is 5.10 g .

Dilutions

To reduce transport and packaging costs, many products such as cordials, detergents, laboratory chemicals and pesticides are sold in a concentrated form. Water is added when they are ready for use. The amount of water added to cordials depends on personal preference, but for certain chemicals such as pesticides there are instructions on the packaging for making solutions of specific concentrations. This ensures the final mixture is suitable for its intended purpose. The process of adding water to a solution to make it less concentrated is called **dilution**.

Dilution is an important laboratory technique used to produce solutions of required concentration for testing and analysis procedures. The dilution process must be as accurate as possible, so the concentration of the required diluted solution is correct.

The usual equipment to ensure accurate measures of volume are pipettes and volumetric flasks similar to those shown in **Figure 14.4.1**.

dilution adding water to a solution to reduce the concentration

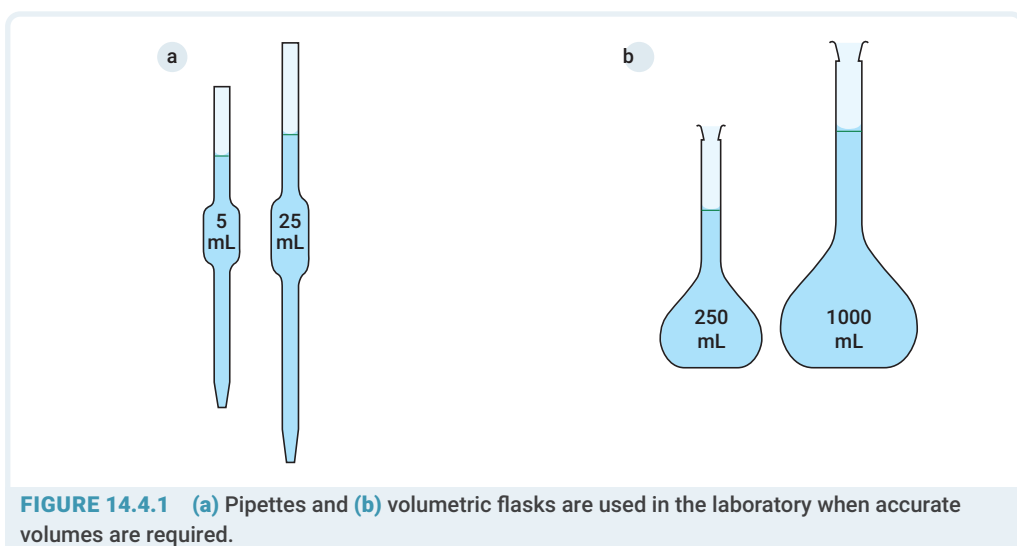


FIGURE 14.4.1 (a) Pipettes and (b) volumetric flasks are used in the laboratory when accurate volumes are required.



Weblink
Virtual lab:
Acid dilution problem

The dilution procedure involves:

- using a pipette to measure a definite volume of a solution of known concentration
- transferring this volume to a volumetric flask
- adding sufficient water to the volumetric flask to make the solution up to the calibration mark.

When calculating the concentration of the diluted solution, scientists make use of the fact that the number of moles (particles) in the diluted solution is the same as in the original volume of solution (**Figure 14.4.2**). The number of moles of solute remains constant, so scientists can apply the following formula to relate concentration and volume of the original solution to that of the diluted solution:

KEY FORMULA

The dilution formula:

$$c_1V_1 = c_2V_2$$

$$C_1V_1 = C_2V_2$$

where: c_1 and c_2 are initial and final concentrations of the solution (in the same units)
 V_1 and V_2 are initial and final volumes of the solution (in the same units).

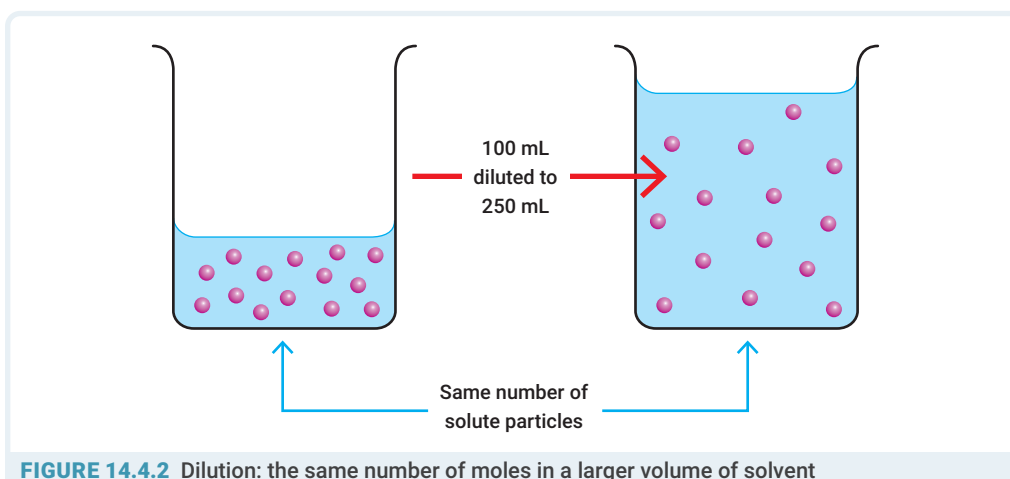


FIGURE 14.4.2 Dilution: the same number of moles in a larger volume of solvent

WORKED EXAMPLE 14.4.2

- a Calculate the volume of 12 mol L^{-1} hydrochloric acid needed to prepare 500 mL of a 0.25 mol L^{-1} solution.
- b Calculate the concentration of a solution made by diluting 20 mL of a 5.0 mol L^{-1} solution to 1.00 L .

ANSWERS

- a 1 Write the relationship.**

$$c_1 V_1 = c_2 V_2$$

- 2 Identify the quantities and check units for consistency.**

$$c_1 = 12 \text{ mol L}^{-1}, V_1 = ?, c_2 = 0.25 \text{ mol L}^{-1}, V_2 = 0.500 \text{ L}$$

- 3 Substitute values to calculate volume.**

$$12 \times V_1 = 0.500 \times 0.25$$

$$V_1 = \frac{0.500 \times 0.25}{12}$$

$$= 0.0104 \text{ L}$$

Volume is 10.4 mL .

- b 1 Write the relationship.**

$$c_1 V_1 = c_2 V_2$$

- 2 Identify the quantities and check units for consistency.**

$$c_1 = 5.0 \text{ mol L}^{-1}, V_1 = 0.020 \text{ L}, c_2 = ?, V_2 = 1.00 \text{ L}$$

- 3 Substitute values to calculate concentration.**

$$5.0 \times 0.020 = c_2 \times 1.00$$

$$c_2 = \frac{5.0 \times 0.020}{1.00}$$

$$= 0.10 \text{ mol L}^{-1}$$

Concentration is 0.10 mol L^{-1} .

PRACTICAL ACTIVITY 14.4.1

PREPARING AND DILUTING SOLUTIONS

Research question

Can solutions of known concentration be prepared and used to determine the unknown concentration of a different solution?

Aim

To prepare solutions of known concentration and use these to determine the unknown concentration of a solution

Materials

- approx. 0.1 g potassium permanganate (KMnO_4)
- solution of unknown concentration
- water
- 250 mL beaker
- 25 mL pipette
- 4 × 250 mL volumetric flasks with stoppers
- stirring rod
- filter funnel
- paper towel
- electronic balance



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Potassium permanganate can stain skin and clothing.	Wear protective clothing and avoid contact.
Potassium permanganate can be irritating to the skin and eyes.	Wear safety glasses and avoid contact with the chemical.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. In particular, expand on the two risks listed to identify specific risks involved. Ask your teacher to check your risk assessment before you proceed.

Procedure

- 1 Weigh approximately 0.1 g of potassium permanganate in one of the beakers. (Make sure you record the mass you are using accurately.)
- 2 Add just enough water to dissolve the crystals of potassium permanganate. Place the filter funnel in the top of a 250 mL volumetric flask and transfer the solution into the flask.
- 3 Remove the filter funnel and fill the flask to the 250 mL mark with water and mix thoroughly.
- 4 Place the stopper on the flask. Holding the flask with one hand on the stopper and the other on the base, mix the contents thoroughly by inverting and righting the flask several times.
- 5 Draw up a table like the one in the results section and complete it for this first solution.
- 6 Pipette 25 mL of the first solution into a clean volumetric flask and repeat steps 3 and 4.
- 7 Complete the table for this second solution.
- 8 Repeat steps 6 and 7 using the second and third solutions so you have successive dilutions.
- 9 Compare the colour of your solution of unknown concentration to the colours of your diluted solutions.

Results

Copy and complete the results table.

Solution	Relative colour intensity	Initial volume	Final volume	Mass KMnO_4	Moles KMnO_4	Concentration (mol L^{-1})
1						
2						
3						
4						

Analysis of results

- Discuss the effect dilution has on:
 - colour intensity
 - concentration of the solution
 - mass of the solute in the diluted solution compared to the original mass.
- Determine the concentration of the unknown solution.

Evaluation

- Evaluate how easy or difficult it was to determine the concentration of the unknown solution.
- Develop a process that could more accurately determine the unknown concentration.
- Devise a method for preparing a solution twice as concentrated as solution 2.

Calculating reacting quantities

Now that we have developed a relationship between moles and concentration, we can extend our calculations to include the stoichiometry of chemical reactions that involve solutions. The number of moles is a key link between reacting quantities and molarity of solutions (Figure 14.4.3).

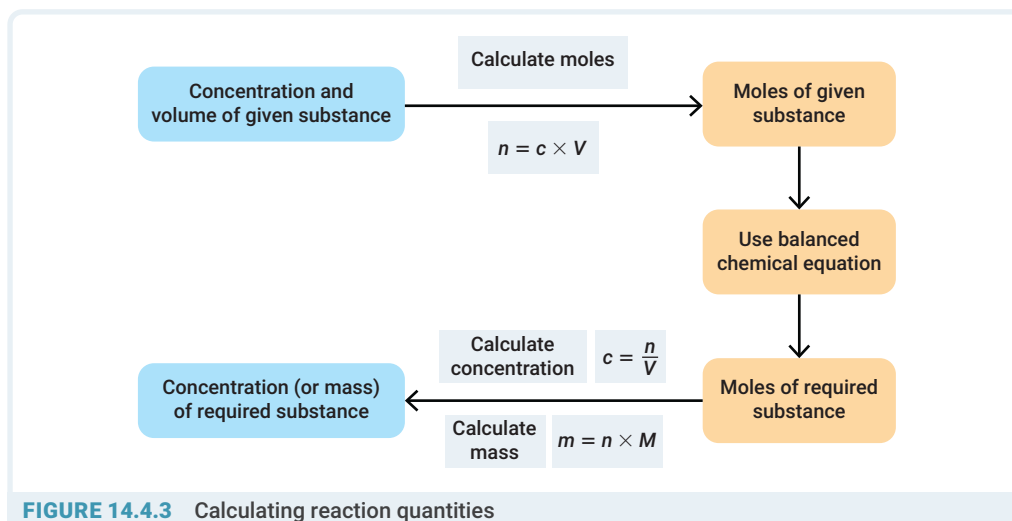


FIGURE 14.4.3 Calculating reaction quantities



Worksheets
Calculating unknowns
quantities

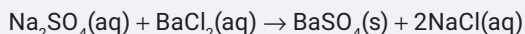
Problems
involving molarity

WORKED EXAMPLE 14.4.3

Calculate the mass of barium sulfate solid that will form when excess sodium sulfate solution is added to 50 mL of a 0.30 mol L⁻¹ barium chloride solution.

ANSWER

1 Write a balanced equation.



2 Calculate the number of moles of BaCl₂.

$$c = \frac{n}{V}$$

$$\therefore n = c \times V$$

$$c = 0.30 \text{ mol L}^{-1}, V = 50 \text{ mL} = 0.050 \text{ L}$$

$$\begin{aligned} n &= 0.30 \times 0.05 \\ &= 0.015 \text{ mol} \end{aligned}$$

3 Use the balanced equation to calculate the number of moles of BaSO₄.

$$n(\text{BaSO}_4) = n(\text{BaCl}_2) = 0.015 \text{ mol}$$

4 Calculate the mass of BaSO₄.

$$M(\text{BaSO}_4) = 233 \text{ g mol}^{-1}$$

$$n = \frac{m}{M}$$

$$\begin{aligned} m &= n \times M \\ &= 0.015 \times 233 \\ &= 3.5 \text{ g} \end{aligned}$$

Mass is 3.5 g.

LEARNING CHECK 14.4

DESCRIBING

1 Recall the formula for calculating dilutions.

APPLYING

2 Calculate the number of moles of solute needed to make the volume of the solution in:

- a 100 mL of a 0.40 M solution of barium nitrate
- b 500 mL of a 0.15 M solution of sodium sulfate
- c 2.00 L of a 0.032 M solution of copper(II) chloride.

3 Calculate the mass of solute required for each of Questions 2a–c.

4 Calculate the molar concentration of:

- a 199 g NiBr₂ dissolved in 5.00 L of solution
- b 0.059 g KF dissolved in 227 mL of solution
- c 23 g ethanol (C₂H₅OH) in 750 mL of solution.

5 Calculate the mass of solute needed to make:

- a 25.0 mL of a 0.02 M solution of sodium carbonate (Na₂CO₃)
- b 32.0 mL of a 0.10 M solution of copper(II) sulfate (CuSO₄).

6 Calculate the volume, in mL, of a 0.025 M solution that contains 1.0×10^{-3} mol of magnesium nitrate.

- 7 **Determine** the mass of copper(II) hydroxide formed when excess copper(II) sulfate is added to 100 mL of a 0.450 mol L^{-1} sodium hydroxide solution, given the balanced equation is:
- $$\text{CuSO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + \text{Na}_2\text{SO}_4(\text{aq})$$
- 8 If excess NaOH is added to 10 mL of a 0.500 mol L^{-1} solution of zinc chloride (ZnCl_2), **calculate** the mass of $\text{Zn}(\text{OH})_2$ produced. (Hint: Write a balanced chemical equation first – NaCl is also formed.)
- 9 **Determine** the number of moles of each of the following in 0.4 mL of 0.117 M magnesium chloride (MgCl_2) solution.
- Magnesium chloride
 - Magnesium ions
 - Chloride ions
- 10 **Calculate** the new concentration when:
- 1.0 L of 12 M stock solution of HCl is diluted with 6.0 L of water
 - 0.025 L of water is added to 125 mL of 0.15 M NaOH solution
 - water is added to 80 mL of 0.45 M NaOH solution to make it up to 150 mL
 - 345 mL of 1.5 M NaCl solution is boiled until the volume reduces to 250 mL
 - 500 mL of a 2.4 M KCl solution is diluted with 620 mL water.
- 11 **Determine** the mass of silver iodide formed when excess sodium iodide is added to 25 mL of a 0.15 M solution of silver nitrate, according to the following equation:
- $$\text{NaI}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{AgI}(\text{s})$$

14.5 Saturated solutions

When adding sugar to a cup of water, there reaches a point at which no more sugar can dissolve at a particular temperature. Any extra sugar added falls to the bottom of the cup. At this point, the solution formed is known as a **saturated solution**. The solution is said to be saturated with sugar. Before this point was reached, it was an **unsaturated solution**, because more sugar could be dissolved. The saturation point is temperature dependent and differs for different solute–solvent combinations (Figure 14.5.1).

saturated solution a solution in which no more solute can dissolve in a given quantity of solvent at a given temperature

unsaturated solution a solution with less than the maximum amount of solute dissolved in a given quantity of solvent at a given temperature

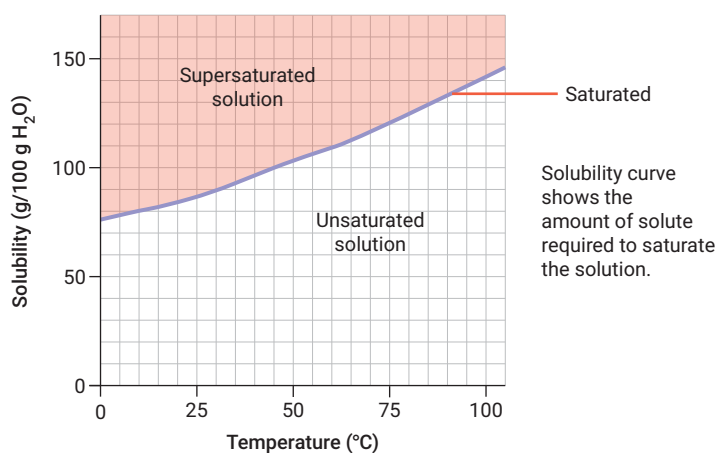


FIGURE 14.5.1 The amount of solute that can dissolve in 100 g of water to produce a saturated solution at different temperatures

**Syllabus link**

Chapter 16 discusses this concept in further detail.

supersaturated solution

an unstable solution that has a higher amount of solute than the saturated solution at a given temperature

If the water is heated, then more sugar will dissolve. If this heated solution is allowed to cool, crystals will start to form again. The temperature at which this happens is the temperature at which the solution is saturated. Heating and cooling a solution carefully can allow it to become a **supersaturated solution**. This means that there is more solute dissolved than in a saturated solution at the same temperature. Supersaturated solutions are unstable. If the supersaturated sugar solution is bumped, if a sugar crystal is added, or if the side of the glass is scratched, then the extra sugar will precipitate out again.

LEARNING CHECK 14.5

DESCRIBING

- 1 Identify** the words from the options provided that make the sentences correct.

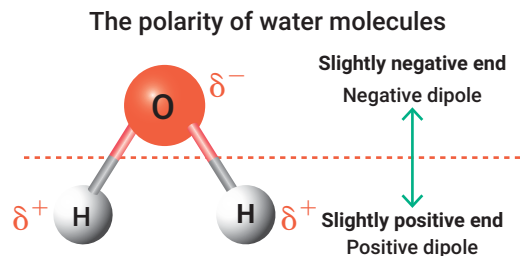
When a solvent holds as much of a solute as it normally can at a given temperature, the solution is *saturated* / *unsaturated* / *supersaturated*. When a solvent holds less dissolved solute than it normally can at a given temperature, the solution is *saturated* / *unsaturated* / *supersaturated*. When a solvent holds more dissolved solute than it normally can at a given temperature, the solution is *saturated* / *unsaturated* / *supersaturated*.

APPLYING

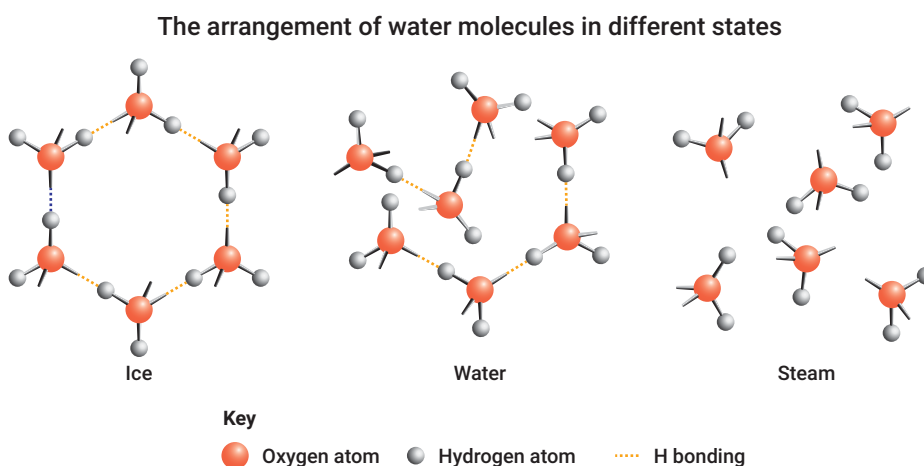
- 2** Sodium thiosulfate readily dissolves in water. At 25°C, a saturated solution has 120 g of sodium thiosulfate for each 100 g of water.
- a** If you added a 1 g crystal of sodium thiosulfate to 2 g of water, this would be the same as having 50 g of crystal in 100 g of water. **Identify** if this solution is likely to be saturated.
 - b** If you add crystals to 2 mL of water until no more dissolves, **identify** the term that describes this solution formed.
 - c** If you add 2 g of crystals to only a few drops of water, **explain** why you would not expect all the crystals to dissolve.
 - d** You heat the solution until it turns clear and then cool it back to room temperature. **Predict** what would happen if a crystal were now added to the clear solution in this test tube.

Unique properties of water

- The polar nature of water makes it a very good solvent.



- Water has strong intermolecular bonds (hydrogen bonding), which means it has relatively high melting and boiling points.
- Because of the arrangement of water molecules, ice is less dense than liquid water.



Calculating concentration

- The molarity of a particular solution is given by the formula:

$$c = \frac{n}{V}$$

where: c = concentration (M)

n = number of moles

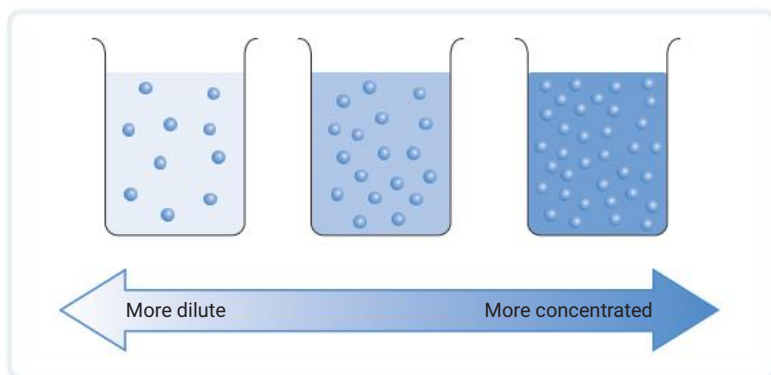
V = volume (L).

- The concentration (in molarity) of a solution is the amount of solute in a given volume of solution, and can be calculated by the above formula.
- Reacting quantities can be calculated by manipulating the $c = \frac{n}{V}$ formula.
- Since the number of moles of substance is the same before and after dilution, diluted solutions can be calculated using the formula:

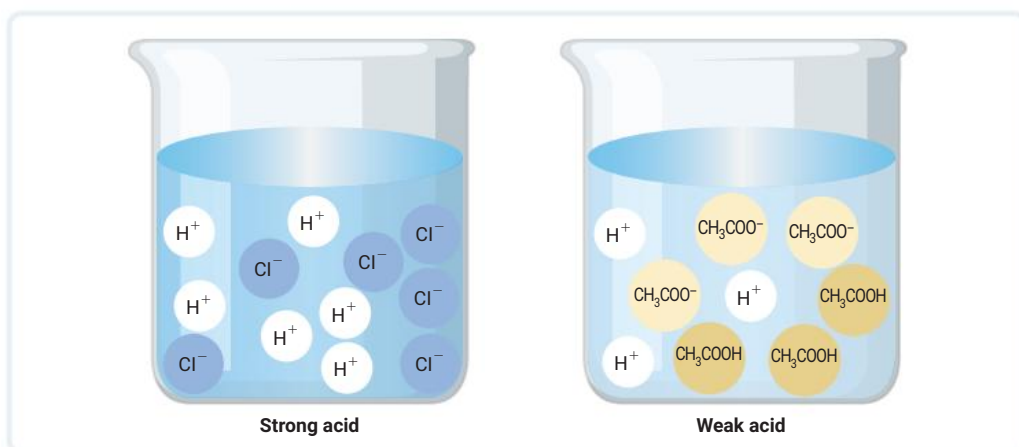
$$c_1V_1 = c_2V_2$$

Concentration vs strength of solutions

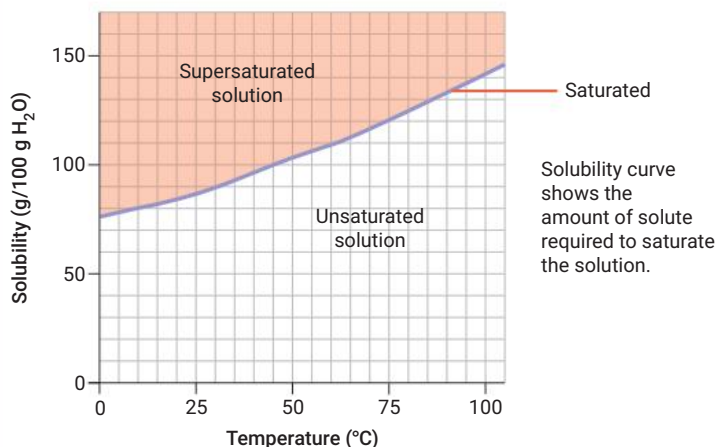
- The concentration of a solution depends on the amount of solute dissolved in a unit of solvent.
- Concentrated solutions have a relatively high amount of solute particles dissolved compared to dilute solutions.



- The strength of a solution, particularly for acids and bases, describes how much of the acid ionises or base dissociates in the solution.
- Concentration and strength should not be confused – solutions can be concentrated or dilute depending on the quantity of solute dissolved and strong or weak depending on how well the substance ionises/dissociates.



A general solubility curve of a solution



Saturated solutions

- Solutions can be described as saturated, unsaturated or supersaturated depending on the amount of solute dissolved at a specified temperature.
- When a solution is saturated, no more solute can be dissolved in the given quantity of solvent at that temperature.
- Solubility of a substance is measured as the number of grams of solute that can be dissolved in 100 g of water at a given temperature.
- The degree of saturation depends on the amount of solute dissolved in a volume of solvent at a given temperature.

MULTIPLE CHOICE

- Identify** the concentration of a solution with 5.60 mol of a solute dissolved into 2 L of solvent.
 A 1.4 M
 B 1.8 M
 C 2.8 M
 D 11.2 M
- What volume of water (mL) must be added to 7.0 mL of 10 M nitric acid to produce a solution of 3.5 M HNO_3 ?
 A 7
 B 13
 C 20
 D 27
- Which of the following is not a property of water?
 A Relatively high boiling point
 B Polar nature enabling a wide range of solutes to dissolve
 C Relatively low melting point
 D High latent heat of vaporisation
- A student adds solid sodium chloride to a beaker of water at room temperature. After stirring the mixture, the student noticed that no solid remained at the bottom of the beaker and more solid could dissolve. At this point, the solution was:
 A saturated.
 B concentrated.
 C unsaturated.
 D supersaturated.
- A 7.0 L sample of 2.0 M of nitric acid (HNO_3) has an extra 15 L of water added to it. What is the new concentration of the HNO_3 solution?
 A 0.64 M
 B 0.875 M
 C 4 M
 D 14 M
- Which property of water allows small insects, such as water striders, to walk on the surface of water?
 A Low cohesion forces
 B Low surface tension
 C High surface tension
 D High adhesion forces
- An aqueous solution of sodium hydroxide (NaOH) is prepared by dissolving 15.864 g of solid NaOH in 500.0 mL of deionised water. In this solution, the concentration of the sodium ion would be closest to:
 A 0.06 M.
 B 0.21 M.
 C 0.24 M.
 D 0.79 M.

8. **Identify** which of the following is the strongest intermolecular force present in water.
- A Covalent bonding
 - B Dipole–dipole attractions
 - C Dispersion forces
 - D Hydrogen bonding
9. The substance that dissolves in a solvent in order to form a solution is the:
- A molecule.
 - B saturate.
 - C solubility.
 - D solute.
10. A solution that has been heated to dissolve a solute and then cooled to the point that it is unstable and solute could crystallise out is called:
- A a saturated solution.
 - B an unsaturated solution.
 - C an undissolved solution.
 - D a supersaturated solution.

SHORT RESPONSE

11. **Calculate** the final molarity of the solutions after the following dilutions.
- a 50 mL of 0.242 M hydrochloric acid is diluted to 500 mL.
 - b 25 mL of 0.152 M sulfuric acid is diluted to 2.0 L.
 - c 10 mL of 0.114 M sodium hydroxide is diluted to 250 mL.
12. **Calculate** the number of moles of solute required to make the following solutions.
- a 2.00 L of a 0.200 mol L^{-1} solution of sodium iodide
 - b 250 mL of a 0.01 mol L^{-1} solution of potassium phosphate
 - c 500 mL of a $0.0075 \text{ mol L}^{-1}$ solution of barium nitrate

CROSS-CHAPTER QUESTION

13. Excess sodium carbonate is added to 75 mL of a 0.08 mol L^{-1} magnesium chloride solution. Solid magnesium carbonate and aqueous sodium chloride are formed.
- a **Determine** a balanced equation for this reaction.
 - b **Calculate** the number of moles of magnesium chloride in the reaction solution.
 - c **Determine** the number of moles of magnesium carbonate formed.
 - d **Calculate** the mass of magnesium carbonate formed.

DATA ANALYSIS

14. Analyse data

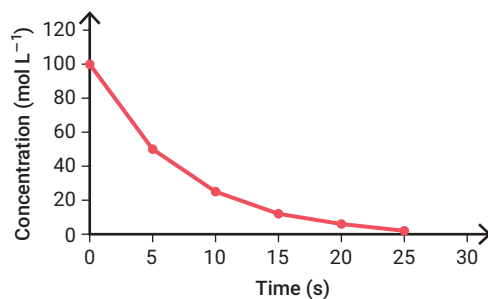
The following table of data shows the solubility of copper(II) sulfate in water from 0°C to 50°C.

Temperature (°C)	Solubility (g/100 g water)
0	14.3
10	17.4
20	20.7
30	25.0
40	28.5
50	33.3

- Identify** the trend in solubility shown by this data.
- Identify** the solubility of copper(II) sulfate in water at 40°C.
- Calculate** the concentration of the copper(II) sulfate solution at 40°C.
(Hint: 100 g water = 100 mL water)

15. Analyse data

A student was conducting an experiment involving a solution of HCl. They measured the change in HCl concentration as the reaction proceeded.



- Based on the trend shown, **identify** whether HCl is a reactant or product in the reaction. **Explain** your answer.
- If the solution was 100 mL, **calculate** the number of moles of HCl present at 5 seconds.

CHAPTER 15

Identifying ions in solution



Natalie/me/Shutterstock.com

SYLLABUS DOT POINTS

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.

SCIENCE INQUIRY

Investigate:

- precipitation reactions to identify cations and anions.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Precipitates are insoluble compounds that form when aqueous solutions are combined. This chapter examines precipitation reactions and shows how to use solubility rules to predict the products of these reactions. Precipitation reactions can be applied practically in a qualitative analysis to determine the identity of unknown cations and anions in solution.

Practicals

- Investigating precipitation reactions
- Testing for carbonate ions using limewater
- Identifying ionic compounds (online-only resource)

Worksheets

- Precipitation challenge
- Modelling precipitation



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ When writing ionic formulas, the cation is written first, followed by the anion.
- ✓ In the names of ionic compounds, the cation is stated first, followed by the anion
- ✓ A balanced equation has an equal number of atoms of each element on the reactant and product sides.
- ✓ Balanced equations include states.
- ✓ Ionic substances are formed from cations and anions.
- ✓ The terms 'solute', 'solvent' and 'solution' have different meanings.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ understand how a precipitate forms when two aqueous solutions are mixed
- ✓ apply the solubility rules to determine if a substance is soluble or insoluble, and therefore predict if a precipitate will form
- ✓ write and balance complete ionic equations and derive net ionic equations
- ✓ identify the spectator ions in a complete ionic equation
- ✓ analyse data in order to determine which specific ions are present in a solution through examining precipitate formation in experimental situations.

15.1 Solubility rules and writing precipitation equations

precipitate an insoluble solid produced when two different solutions react

precipitation reaction a reaction between two solutions that produces a solid

suspension a mixture of a solution and an insoluble solid; the solid distributes evenly throughout the solution

When two solutions of ionic compounds are mixed, typically the ionic compounds exchange partners to form new ionic compounds. This is known as a double-displacement reaction – a concept introduced in Chapter 8. Depending on the solubility of the products, a solid known as a **precipitate** may form. This type of reaction is called a **precipitation reaction (Figure 15.1.1)**. Often, when a precipitate forms, it does not immediately sink to the bottom of the container; rather, it distributes evenly throughout the liquid, forming a cloudy-looking **suspension**, which may settle out over time.

For example, when sodium chloride (NaCl) solution is added to silver nitrate (AgNO_3) solution, a white solid of silver chloride (AgCl) forms. Sodium chloride solution is composed of sodium ions (Na^+) and chloride ions (Cl^-) dissolved in water. Silver nitrate solution is composed of silver ions (Ag^+) and nitrate ions (NO_3^-) dissolved in water. When the solution of Na^+ and Cl^- ions is mixed with the solution of Ag^+ and NO_3^- ions, the Ag^+ ions become attracted to the Cl^- ions and the Na^+ become attracted to the NO_3^- ions. Ag^+ and Cl^- ions are more strongly attracted to each other than to the water, so they cluster together in groups and form a solid, which

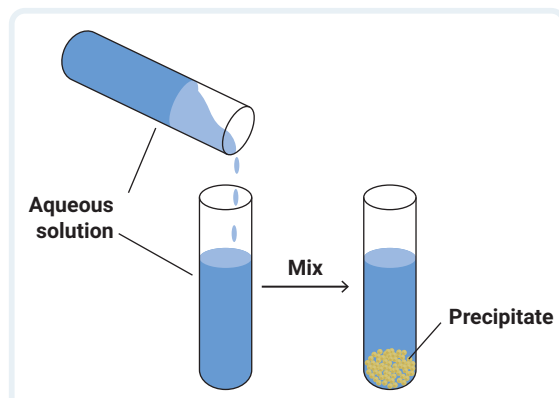
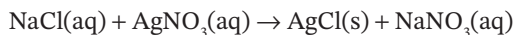


FIGURE 15.1.1 A solid precipitate can form when two aqueous solutions are mixed.

precipitates out of solution; that is, a precipitate of solid AgCl forms, as shown in **Figures 15.1.2** and **15.1.3**. This is represented by the following equation:



Sodium nitrate is very soluble in water because the sodium and nitrate ions are more strongly attracted to the water molecules than to each other, so the sodium and nitrate ions stay independent and continue to move randomly through the solution. They are not involved in the reaction and they remain in the solution as ions. Ions that are not involved in a reaction are called **spectator ions**. These spectator ions remain in solution in the **supernate**.

By systematically mixing solutions that contain known cations and anions, chemists have experimentally identified which compounds are soluble and which are insoluble, forming a precipitate.

Table 15.1.1 summarises the solubility of common ionic compounds. The table applies to some common cations and is organised by the solubility of various anions. In general, all salts containing group 1, ammonium or nitrate ions are soluble. You may also use the solubility chart found in your *Formula and Data Booklet* (**Table 15.1.2**) to determine if a substance is soluble (s), partially soluble (p) or insoluble (i) in water.



FIGURE 15.1.2 A silver chloride precipitate forms when solutions of sodium chloride and silver nitrate are mixed.

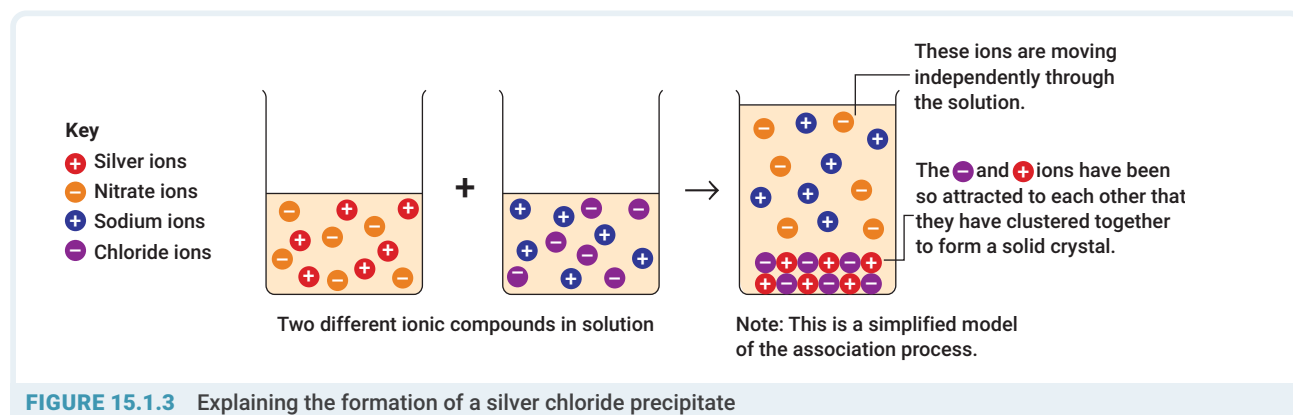


FIGURE 15.1.3 Explaining the formation of a silver chloride precipitate

TABLE 15.1.1 Solubility rules for some common ions

Soluble anions	Exceptions
Nitrates (NO_3^-)	None
Chlorides (Cl^-)	Ag^+ insoluble; Pb^{2+} slightly soluble
Bromides (Br^-)	Ag^+ insoluble; Pb^{2+} slightly soluble
Iodides (I^-)	Ag^+ , Pb^{2+} insoluble
Sulfates (SO_4^{2-})	Ba^{2+} , Pb^{2+} , Sr^{2+} insoluble; Ag^+ , Ca^{2+} slightly soluble
Insoluble anions	Exceptions
Hydroxides (OH^-)	Group 1, NH_4^+ , Ba^{2+} , Sr^{2+} soluble; Ca^{2+} slightly soluble
Oxides (O^{2-})	Group 1, NH_4^+ , Ba^{2+} , Sr^{2+} , Ca^{2+} soluble
Sulfides (S^{2-})	Groups 1 and 2, NH_4^+ soluble
Carbonates (CO_3^{2-})	Group 1, NH_4^+ soluble
Phosphates (PO_4^{3-})	Group 1, NH_4^+ soluble

spectator ion an ion that remains in solution before and after a chemical reaction has taken place; it is not involved in the reaction

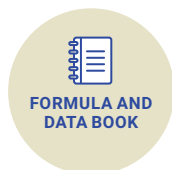
supernate the solution above a precipitate; composed of spectator ions in aqueous solution



Worksheet
Precipitation challenge

TABLE 15.1.2 Solubility of selected compounds at 298 K

	Bromide	Carbonate	Chloride	Hydroxide	Iodide	Nitrate	Oxide	Phosphate	Sulfate
Aluminium	s		s	i	s	s	i	i	s
Ammonium	s	s	s	s	s	s		s	s
Barium	s	i	s	s	s	s	s	i	i
Calcium	s	i	s	p	s	s	p	i	p
Cobalt(II)	s	i	s	i	s	s	i	i	s
Copper(II)	s		s	i	i	s	i	i	s
Iron(II)	s	i	s	i	s	s	i	i	s
Iron(III)	s		s	i	s	s	i	i	s
Lead(II)	p	i	s	i	i	s	i	i	i
Lithium	s	s	s	s	s	s	s		s
Magnesium	s	i	s	i	s	s	i	p	s
Manganese(II)	s	i	s	i	s	s	i	p	s
Potassium	s	s	s	s	s	s	s	s	s
Silver	i	i	i	i	i	s	i	i	p
Sodium	s	s	s	s	s	s	s	s	s
Zinc	s	i	s	i	s	s	i	i	s



neutral species equation

a balanced chemical equation that gives the complete neutral chemical formulas of reactants and products without breaking them down into individual ions

complete ionic equation

shows all of the soluble ions (cations and anions) present in a solution as well as the formula of any precipitate formed

net ionic equation

represents only the chemical species that are involved in the precipitate reaction, with spectator ions eliminated

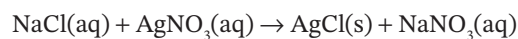
Key

Abbreviation	Explanation
s	Soluble in water (solubility greater than 10 g L^{-1})
p	Partially soluble in water (solubility between 1 and 10 g L^{-1})
i	Insoluble in water (solubility less than 1 g L^{-1})
	No data

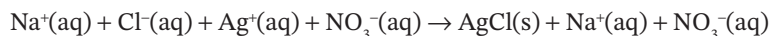
Formula and data booklet, p. 7, QCAA

To use this table, find the two ions involved in the ionic compound. Align them on the table and identify the letter in the overlapping cell. Use the key to identify whether the ionic compound is soluble, partially soluble or insoluble in water. For example, aligning the silver and chloride, we can see that there is an 'i' in the overlapping cell. This indicates that silver chloride is insoluble and therefore forms a precipitate.

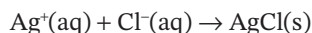
This reaction can be represented in different ways. A **neutral species equation** shows the balanced chemical equation. It can be used to identify the products of the reaction, and then the precipitate may be identified by using the solubility rules (Table 15.1.1) or the solubility chart (Table 15.1.2). The **complete ionic equation** is written to show how the soluble ions react with one another to form the precipitate. The solid precipitate is not written in ionic form because it is insoluble. The **net ionic equation** remains once the spectator ions are identified and eliminated from the complete ionic equation. The following example illustrates the difference between the neutral species, complete ionic and net ionic equations.

Neutral species (balanced chemical) equation:

AgCl is the insoluble precipitate.

Complete ionic equation:

In the complete ionic equation, the number of ions are balanced and AgCl(s) remains because it is insoluble.

Net ionic equation:

In the net ionic equation, the spectator ions Na^+ and NO_3^- have been removed.

The net ionic equation represents the reaction most accurately of the three types of equation because it focuses on the newly formed substance. The net ionic equation is most commonly used to represent precipitation reactions. Note that all three ways of representing these reactions are balanced.



Worksheet
Modelling precipitation



Weblinks
Molecular, complete ionic
and net ionic equations
Ionic reaction simulator

PRACTICAL ACTIVITY 15.1.1**INVESTIGATING PRECIPITATION REACTIONS****Introduction**

Solubility rules can be formulated from experimental results. In principle, solubility is predictable but the calculations are complex and require highly accurate estimates of intermolecular interactions, a subject that remains an active area of chemical research.

In this experiment, you will determine which mixtures of solutions form precipitates and use these results to construct solubility rules. These can then be compared with existing experimental data from the chemical literature.

Research question

How accurately can experimental precipitation reactions be used to identify the solubility rules that currently exist in the literature?

Aim

To compare the solubility of a range of ionic substances through precipitation reactions

Equipment per group

- cations (all 0.1 mol L⁻¹) in dropper bottles: sodium nitrate (NaNO₃), silver nitrate (AgNO₃), ammonium nitrate (NH₄NO₃), magnesium sulfate (MgSO₄), iron(II) sulfate (FeSO₄), copper(II) sulfate (CuSO₄), zinc sulfate (ZnSO₄), aluminium sulfate (Al₂(SO₄)₃), barium chloride (BaCl₂), calcium chloride (CaCl₂), iron(III) chloride (FeCl₃)
- anions (all 0.1 mol L⁻¹) in dropper bottles: potassium nitrate (KNO₃), potassium chloride (KCl), sodium sulfate (Na₂SO₄), sodium carbonate (Na₂CO₃), sodium acetate (NaCH₃COO), sodium hydroxide (NaOH)
- distilled water
- 6 test tubes and test-tube rack
- paper towel
- test-tube brush

**What are the risks in doing this experiment?**

Barium salts are poisonous.

Barium, silver and copper salts are harmful to the environment.

How can you manage these risks to stay safe?

Wear safety glasses and avoid contact with skin. If contact occurs, wash thoroughly with soap and water.

Do not pour solutions containing these salts down the drain. Dispose of as directed by your teacher.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the two listed risks to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- 1 Put a few drops of each anion into each test tube, as shown in **Figure 15.1.4**. Take care not to mix droppers.

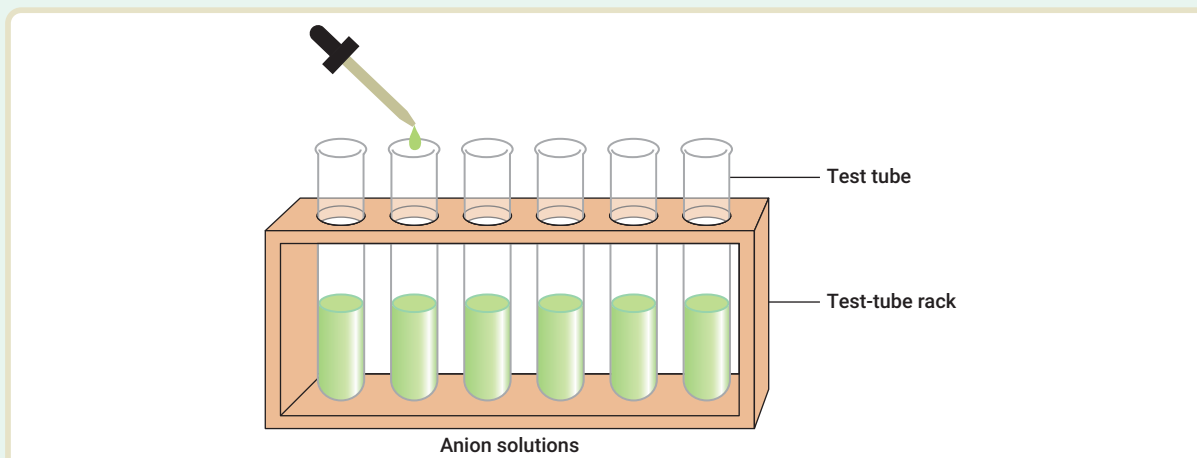


FIGURE 15.1.4 The experimental set-up

- 2 Add a few drops of the first cation to each of the six test tubes. Record your observations.
- 3 Clean all test tubes thoroughly because incorrect results can arise from contamination of solutions.
- 4 Repeat steps 2 and 3 for the next cation. Continue until all cations have been tested.

Results

Draw up a table similar to the one below to record your observations. For no reaction, record 'NR'; for a precipitate, record 'P' and the colour of the precipitate.

Cation	Anion					
	NO_3^- nitrate	Cl^- chloride	SO_4^{2-} sulfate	CO_3^{2-} carbonate	CH_3COO^- acetate	OH^- hydroxide
Na^+						
Ag^+						
NH_4^+						
Mg^{2+}						
Fe^{2+}						
Cu^{2+}						
Zn^{2+}						
Al^{3+}						
Ba^{2+}						
Ca^{2+}						
Fe^{3+}						

Analysis of results

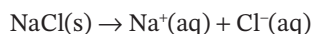
- 1 Summarise your results using statements such as 'All nitrates are ...'
- 2 Write a balanced net ionic equation for each of the reactions in which a precipitate was obtained.

Interpretation

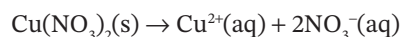
- 3
 - a Compare the results of the experiment with the solubility rules in Table 15.1.1.
 - b What, if any, were the discrepancies between your results and Table 15.1.1? Suggest a reason for these discrepancies and a way of resolving them.

Solubility equations

Scientists can determine a great deal about what is happening in a reaction from the states in a chemical equation. When salts dissolve, chemists generally do not write the ions separately. For example, dissolved sodium chloride is written as NaCl(aq) . The state aqueous (aq) indicates that the ions have **dissociated**. The Na^+ and Cl^- ions are separated and surrounded by water molecules. They are charged ionic particles, so the solution can conduct electricity; it is an **electrolytic solution**. Two examples for the dissociation of salts are:

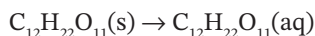


Sodium chloride $\rightarrow$ sodium ions + chloride ions



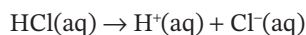
Copper(II) nitrate $\rightarrow$ copper ions + nitrate ions

When molecular substances such as carbon dioxide, hydrochloric acid (HCl) or glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) dissolve, the state is aqueous.



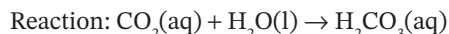
Solid sucrose $\rightarrow$ aqueous sucrose

When acids such as HCl ionise in water, they are written as HCl(aq) . This means that the solution contains separate H^+ and Cl^- ions.



Hydrochloric acid $\rightarrow$ hydrogen ions + chloride ions

After a substance becomes aqueous, further reactions may take place. For example, carbon dioxide has to dissolve in water before it can react with the water. The reaction with water is written separately:



Carbon dioxide + water $\rightarrow$ carbonic acid

Many precipitation reactions involve the exchange of ions between compounds in aqueous ionic solutions. This will be the focus for the rest of this chapter.

dissociated when the ions from an ionic compound separate from each other and form a solution in water

electrolytic solution a solution that contains ions, so it will conduct electricity

WORKED EXAMPLE 15.1.1

- a** What precipitate, if any, forms when you mix aqueous solutions of magnesium iodide and silver nitrate? If a precipitate forms, write the neutral species, complete ionic and net ionic equations for the reaction.
- b** Select a reagent that could be used to precipitate the cation in BaBr_2 . Explain your selection.

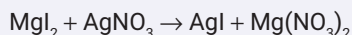
ANSWERS

- a 1 Write the formulas of the compounds present.**

Magnesium iodide (MgI_2)

Silver nitrate (AgNO_3)

- 2 Predict the products by swapping the partners of the reacting compounds and check that the formulas are correct.**

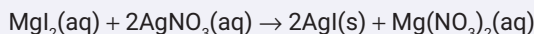


- 3 Use Tables 15.1.1 and 15.1.2 to confirm if either product is insoluble.**

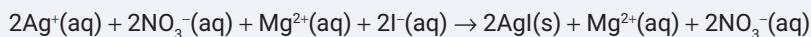
All nitrates are soluble and all iodides are soluble except Ag^+ and Pb^{2+} , so AgI will precipitate.

- 4 Write the required equations, checking they are balanced.**

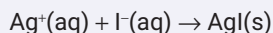
Neutral species (balanced chemical) equation:



Complete ionic equation showing all ions (except the precipitate):



Net ionic equation:



- b 1 Identify the ions present and one to be precipitated.**

Ba^{2+} and Br^- are the ions present in solution. Ba^{2+} is the cation to be precipitated.

- 2 Use Table 15.1.1 or 15.1.2 to identify an anion that could be used to precipitate the identified cation.**

Anions that will produce an insoluble compound with Ba^{2+} , CO_3^{2-} , SO_4^{2-} and PO_4^{3-} .

- 3 Choose one anion and identify a cation to form a compound. Check that the cation does not react with the anion from the original compound.**

Generally, it is best to choose a cation from group 1 or NH_4^+ because compounds of these are soluble. It would be incorrect to choose Ag^+ and Pb^{2+} because they both precipitate with Br^- .

- 4 Identify the reagent and write the formula.**

Na_2CO_3 can be used to precipitate Ba^{2+} from a solution of BaBr_2 .

LEARNING CHECK 15.1

DESCRIBING

- 1 Define:**

- a** precipitate
- b** spectator ion
- c** dissociation
- d** electrolytic solution
- e** complete ionic equation
- f** neutral species equation
- g** net ionic equation
- h** supernate.

APPLYING

- 2 **Classify** the following substances as soluble or insoluble in water.
- Sodium hydroxide
 - Aluminium oxide
 - Copper(II) sulfate
 - Silver carbonate
 - Lithium chloride
- 3 **Determine** neutral species (overall) equations, complete ionic equations and net ionic equations for the reactions (if any) that occur when solutions of the following pairs of substances are mixed. If there is no reaction, write 'NR'.
- Sodium chloride and silver nitrate
 - Copper(II) sulfate and potassium hydroxide
 - Sodium carbonate and iron(II) sulfate
 - Zinc nitrate and ammonium hydroxide
 - Potassium carbonate and calcium chloride
- 4 **Identify** an aqueous solution that you could substitute for the identified solution in the pair that mixes to produce an identical precipitate to the one formed when:
- potassium carbonate and zinc chloride are mixed (substitute potassium carbonate)
 - ammonium sulfate and lead(II) chloride are mixed (substitute ammonium sulfate)
 - barium nitrate and copper(II) sulfate are mixed (substitute barium nitrate).
- 5 **Determine** which aqueous solutions you would mix to form the following solid precipitates. Write neutral species equations for your reactions.
- Magnesium carbonate
 - Lead(II) sulfate
 - Silver bromide
- 6 **Determine** which salt, if any, will precipitate when solutions of zinc sulfate and potassium hydroxide are mixed. If a precipitate does form, write the neutral species equation and net ionic equation.
- 7 **Determine** the name and formula of the precipitate (if any) that forms when each of the following pairs of substances is mixed.
- Copper(II) chloride and potassium hydroxide
 - Sodium carbonate and ammonium sulfate
 - Lead(II) nitrate and copper(II) sulfate

ANALYSING

- 8 Separate quantities of an unknown solution form precipitates when mixed with potassium carbonate and sodium hydroxide solutions, but do not form precipitates when mixed with solutions of sodium chloride or ammonium sulfate.

Identify the possible contents of the unknown solution.

- Pb^{2+} , Zn^{2+} , Cu^{2+}
- Mg^{2+} , Al^{3+} , Zn^{2+}
- Mg^{2+} , Ba^{2+} , Al^{3+}
- Cu^{2+} , Zn^{2+} , Ag^{+}

15.2 Determination of ions in solution

Several common tests can be used to identify particular ions in a solution. The experimental results can be compared with known test results, allowing ions to be identified by the formation of precipitates.

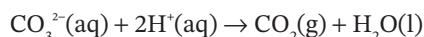
Limewater test

limewater an aqueous solution of calcium hydroxide ($\text{Ca}(\text{OH})_2$)

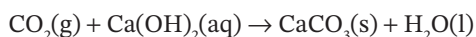
One of the most widely used tests involving precipitates is the test for carbonates. It involves using an acid and **limewater** to determine the presence of carbonate ions through formation of a precipitate.

To identify the presence of carbonates in a sample, scientists follow these steps:

1. Add dilute acid to the sample. Typically, dilute hydrochloric acid is used, but any acid will work. If there are carbonates present in the sample, the acid will react with the carbonate to form carbon dioxide gas. This will result in effervescence (fizzing) as the gas is emitted.



2. The gas given off is collected and bubbled through limewater (aqueous solution of $\text{Ca}(\text{OH})_2$). The carbon dioxide gas reacts with limewater in the following reaction:



Calcium carbonate (CaCO_3) is a precipitate. When it is formed, the precipitate is in suspension, giving the solution a white, cloudy appearance. Thus, the limewater turning cloudy indicates that the gas is carbon dioxide, and that carbonates were present in the original reaction with acid.

Practical activity 15.2.1 is an example of the carbonate test using acid and limewater.

PRACTICAL ACTIVITY 15.2.1

TESTING FOR CARBONATE IONS USING LIMEWATER

Introduction

In this experiment, you will conduct an experiment using acid and limewater to determine the presence of carbonate ions in a substance.

Research question

How can limewater be used to identify carbonates in a solution?

Aim

To determine the presence of carbonate anions using an acid-carbonate reaction and a precipitation reaction involving limewater

Materials

- calcium carbonate (CaCO_3) chips
- sodium hydrogen carbonate (NaHCO_3)
- copper(II) carbonate (CuCO_3)
- 1.0 M hydrochloric acid
- limewater solution
- 4 test tubes
- stopper with a delivery tube
- test-tube rack
- retort stand and clamp
- 10 mL measuring cylinder

Procedure

- 1 Half-fill one of the test tubes with prepared limewater solution and place it in test-tube rack.
- 2 Put a small number of calcium carbonate chips into the other test tube. Position this test tube as shown in **Figure 15.2.1**.
- 3 Measure 5 mL of 1.0 M HCl and add this to the calcium carbonate chips. Quickly place the stopper on the test tube and allow the gas evolved to bubble into the limewater solution.

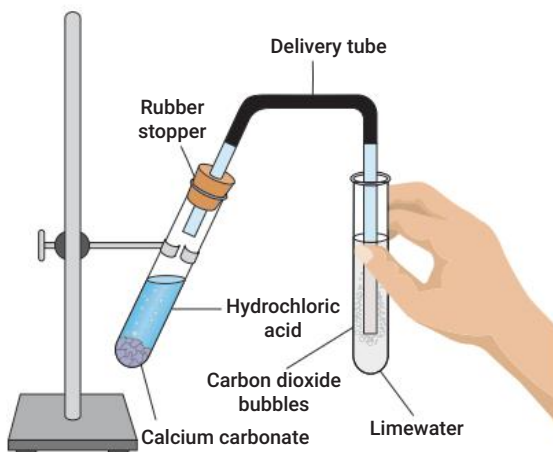


FIGURE 15.2.1 The experimental set-up

- 4 Record your observations in your results table.
- 5 Repeat steps 1–4 for the other carbonate and the hydrogen carbonate.

Results

Copy and complete the following table.

Substance	Gas formed? Y/N	Limewater observations
Calcium carbonate		
Sodium hydrogen carbonate		
Copper(II) carbonate		

Analysis of results

- 1 Determine if a precipitate formed in the limewater for each substance and complete the results table accordingly.
- 2 Explain what caused the limewater to turn milky.
- 3 Explain why the acid was added to the substances containing hydrogen carbonate or carbonate ions.
- 4 Determine the overall chemical formula for the formation of the precipitate.

Aside from testing for carbonates, there are many other tests to identify common anions and cations. These are outlined in **Tables 15.2.1, 15.2.2** and **15.2.3**.

TABLE 15.2.1 Tests used to identify common anions

Anion	Tests
Carbonate (CO_3^{2-})	- Solution has a pH between 8 and 11 (pH test strip suffices). - Addition of dilute HNO_3 evolves a colourless gas (CO_2), which is bubbled through limewater to produce a white precipitate.^A
Sulfate (SO_4^{2-})	Acidification and addition of Ba^{2+} or Pb^{2+} ions produces a white precipitate.
Phosphate (PO_4^{3-})	Addition of ammonia and Ba^{2+} or ammonia/ammonium nitrate buffer solution with Mg^{2+} produces a white precipitate.
Chloride (Cl^-)	Addition of AgNO_3 to an acidified sample produces a white precipitate,^B which dissolves in ammonia solution and darkens in sunlight.

^A Any strong acid would do, but for analysing mixtures scientists do not want to introduce any Cl^- or SO_4^{2-} .

^B In non-acidic solutions, silver nitrate also produces precipitates with carbonate and phosphate (and with sulfate at all pH values if sulfate concentration is moderately high), so this test alone does not prove the presence of chloride; it is also necessary to prove the absence of sulfate.



Weblink

Testing for negative ions

TABLE 15.2.2 Precipitates formed between cations and anion

Cation ^A	Anion					
	OH^-	Cl^-	SO_4^{2-}	CO_3^{2-} in alkaline solution ^B	PO_4^{3-}	
					Solution pH 2	Solution pH 6
Ba^{2+}	No	No	Yes	Yes	No	Yes
Pb^{2+}	Yes	Yes ^C	Yes	Yes	No	Yes
Ag^+	Yes ^D	Yes	Yes ^C	Yes	No	Yes
Cu^{2+}	Yes	No	No	Yes	No	Yes

^A In acidic or alkaline solution.

^B Cannot have carbonate in acid solution: it decomposes to $\text{CO}_2(\text{g})$.

^C Provided concentration of Cl^- or SO_4^{2-} is not too low, say $> 0.05 \text{ mol L}^{-1}$.

^D Precipitate is dark brown or black Ag_2O .

TABLE 15.2.3 Tests used to identify common cations

Cation	Tests
Ag^+	With Cl^- forms a white precipitate that dissolves in ammonia solution
Pb^{2+}	- With OH^- forms a white precipitate - With I^- forms a yellow precipitate
Ba^{2+}	- With SO_4^{2-} forms a white precipitate - Gives a pale green flame colour - No precipitate with OH^- or F^- (compare Ca^{2+})
Ca^{2+}	- With SO_4^{2-} forms a white precipitate (however, not if the test solution is dilute) - With F^- forms a white precipitate - Gives a brick-red flame colour
Zn^{2+}	- With hydroxide forms a white precipitate - This precipitate dissolves in both excess OH^- solution and ammonia solution
Cu^{2+}	- With OH^- forms a blue precipitate - This precipitate dissolves in NH_3 to form a deep blue solution - Gives a blue-green flame colour
Fe^{2+}	- With OH^- forms a green or white precipitate, which may turn brown - Decolourises acidified dilute potassium permanganate solution
Fe^{3+}	- With OH^- forms a brown precipitate - With thiocyanate (SCN^-) forms a deep-red solution



Practical activity
Identifying ionic compounds



Syllabus link
Chapter 4 discussed how the flame test can be used to identify metal ions.

In addition, cations may be identified with flame tests, by noting the distinctive colour they emit when placed in a flame.

LEARNING CHECK 15.2

DESCRIBING

- Define:**
 - suspension
 - limewater.
- Explain** how the limewater test works and which ion it identifies.

APPLYING

- Determine** if a precipitation reaction occurs when the following reactants are mixed by using a solubility table. Where a reaction occurs, write the neutral species equation and net ionic equation for the reaction.
 - Potassium chloride and zinc nitrate
 - Ammonium bromide and lead(II) nitrate
 - Ammonium sulfide and magnesium acetate



- 4 **Identify** solutions that would be mixed to produce each of the following precipitates.
- Lead(II) sulfate
 - Iron(II) carbonate
 - Magnesium hydroxide
- 5 A solution was thought to contain either lead(II) nitrate or calcium nitrate. **Describe** the tests that could be carried out to identify the substance correctly.
- 6 **Describe** a test that could be used to determine whether a solution contained:
- lead(II) nitrate or barium nitrate
 - copper(II) sulfate or iron(II) sulfate.

ANALYSING

- 7 A solution is believed to contain either barium chloride or calcium chloride. Samples of the solution give a white precipitate with Na_2SO_4 and a pale-green flame colour. **Determine** which of the ions is present. Give reasons for your response.
- 8 A student is given four concentrated solutions. Each contains one of the following cations: Pb^{2+} , Cu^{2+} , Ba^{2+} , Ca^{2+} , Fe^{2+} or Fe^{3+} . The results of a series of tests that were conducted are given in the following table; NP stands for no precipitate and ppt stands for precipitate.

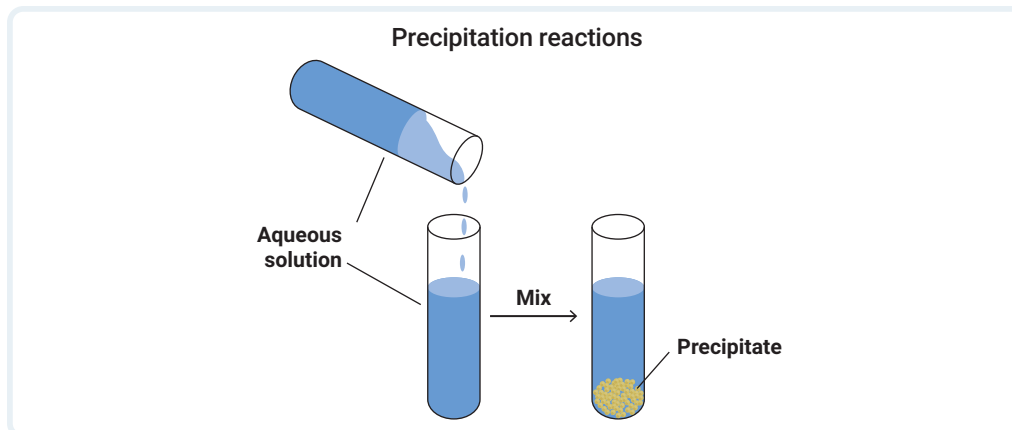
Reagent added	A	B	C	D
KI	Yellow ppt	NP	NP	NP
H_2SO_4	ppt	NP	ppt	ppt
NaOH	ppt	Brown ppt	NP	ppt

Identify possible cations for A–D using the data above. If there is not enough information, suggest other tests and their expected results that could be used to identify any still unknown solutions.

- 9 A chain of coffee shops started giving away brightly coloured drinking mugs. It was thought that lead compounds had been used in the paint and that this dissolved in acidic drinks such as soft drinks and fruit juices.
- Explain** how an experiment could be used to determine whether lead is present in the paint.

Precipitation reactions

- Insoluble solid precipitates can form when two aqueous solutions containing ions are mixed.



Solubility rules

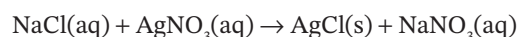
- Solubility rules or a solubility table can be used to determine if a substance is soluble or insoluble.

Soluble anions	Exceptions
Nitrates (NO_3^-)	None
Chlorides (Cl^-)	Ag^+ insoluble; Pb^{2+} slightly soluble
Bromides (Br^-)	Ag^+ insoluble; Pb^{2+} slightly soluble
Iodides (I^-)	Ag^+ , Pb^{2+} insoluble
Sulfates (SO_4^{2-})	Ba^{2+} , Pb^{2+} , Sr^{2+} insoluble; Ag^+ , Ca^{2+} slightly soluble
Insoluble anions	Exceptions
Hydroxides (OH^-)	Group 1, NH_4^+ , Ba^{2+} , Sr^{2+} soluble; Ca^{2+} slightly soluble
Oxides (O^{2-})	Group 1, NH_4^+ , Ba^{2+} , Sr^{2+} , Ca^{2+} soluble
Sulfides (S^{2-})	Groups 1 and 2, NH_4^+ soluble
Carbonates (CO_3^{2-})	Group 1, NH_4^+ soluble
Phosphates (PO_4^{3-})	Group 1, NH_4^+ soluble

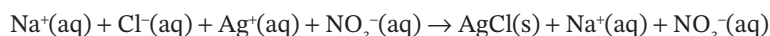
Writing ionic equations for reactions between ionic compounds

- Reactions involving ionic aqueous solutions can be represented in three different ways:

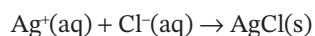
- Neutral species (balanced chemical) equation:



- Overall ionic equation:



- Net ionic equation (spectator ions removed):



CHAPTER EXAM

MULTIPLE CHOICE

- Identify** which of the following solutions will not form a precipitate when added to an aqueous solution of iron(II) sulfate ($\text{FeSO}_4(\text{aq})$).
 - Barium chloride ($\text{BaCl}_2(\text{aq})$)
 - Sodium hydroxide ($\text{NaOH}(\text{aq})$)
 - Lithium carbonate ($\text{Li}_2\text{CO}_3(\text{aq})$)
 - Ammonium nitrate ($\text{NH}_4\text{NO}_3(\text{aq})$)
- Identify** which of the following reactants will form a precipitate when added to a solution of sodium sulfate.
 - Potassium sulfate solution
 - Copper(II) chloride solution
 - Barium nitrate solution
 - Ammonium carbonate solution
- Identify** the precipitate that forms when lead(II) chloride solution is mixed with sodium iodide solution.
 - Lead solid
 - Lead(II) iodide
 - Sodium chloride
 - Sodium metal
- Identify** which of the following aqueous solutions would form a precipitate when added to an aqueous solution of barium chloride.
 - Ammonium sulfate
 - Copper(II) nitrate
 - Lithium hydroxide
 - Potassium iodide
- Identify** which of the following would form a precipitate when added to an aqueous lithium chloride solution.
 - $\text{AgNO}_3(\text{aq})$
 - $\text{Ca}(\text{NO}_3)_2(\text{aq})$
 - $\text{CuSO}_4(\text{aq})$
 - $\text{Fe}_2(\text{SO}_4)_3(\text{aq})$
- Identify** the net ionic equation that represents the mixing of an aqueous solution of barium nitrate and chromium(III) sulfate to form a white precipitate.
 - $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
 - $2\text{Cr}^{3+}(\text{aq}) + 3\text{NO}_3^-(\text{aq}) \rightarrow \text{Cr}(\text{NO}_3)_3(\text{aq})$
 - $3\text{Ba}(\text{NO}_3)_2(\text{aq}) + \text{Cr}_2(\text{SO}_4)_3(\text{aq}) \rightarrow 3\text{BaSO}_4(\text{s}) + 2\text{Cr}(\text{NO}_3)_3(\text{aq})$
 - $\text{Ba}^{2+}(\text{aq}) + 3\text{NO}_3^-(\text{aq}) + 2\text{Cr}^{3+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{Cr}_2(\text{NO}_3)_3(\text{aq})$
- Identify** the solution you would mix with calcium chloride to obtain a precipitate of the cation Ca^{2+} .
 - $\text{AgNO}_3(\text{aq})$
 - $\text{NaI}(\text{aq})$
 - $\text{NH}_4\text{Br}(\text{aq})$
 - $\text{Na}_3\text{PO}_4(\text{aq})$

8. **Identify** which of the following substances is soluble in water.
- A Barium carbonate
 - B Lithium carbonate
 - C Magnesium carbonate
 - D Zinc carbonate
9. **Identify** which of the following cations forms an insoluble compound with the sulfate anion.
- A Ba^{2+}
 - B Cu^{2+}
 - C Fe^{2+}
 - D Mg^{2+}
10. **Identify** which of the following cations forms an insoluble compound with the iodide anion.
- A Cu^{2+}
 - B Na^{+}
 - C Mg^{2+}
 - D Zn^{2+}

SHORT RESPONSE

11. A student mixed lead(II) chloride and iron(II) sulfate and noticed that a precipitate had formed at the bottom of the beaker.
- a **Determine** the neutral species equation for this reaction.
 - b **Identify** the name of the precipitate formed.
 - c **Determine** the overall ionic equation for this reaction.
 - d **Determine** the net ionic equation for this reaction.
12. A solution is known to contain either sodium carbonate or sodium chloride. Samples of the solution evolved a gas when mixed with hydrochloric acid (HCl) and formed a precipitate when mixed with $\text{Ba}(\text{NO}_3)_2$.
- Explain** which anion is present.

CROSS-CHAPTER QUESTION

13. Solid sodium chloride is found in deposits well below ground level. Cold water is pumped underground at high pressure to bring the sodium chloride to the surface as brine, a concentrated aqueous solution. Often, the brine produced contains magnesium ions as contaminants. These magnesium ions can be removed by precipitation.
- a **Explain** why it might be better to use hot water rather than cold water for this process.
 - b **Evaluate** whether sodium carbonate would be a suitable chemical to add to the brine to precipitate out the magnesium ions.
 - c **Determine** the net ionic equation for this precipitation reaction.

DATA ANALYSIS

14. Analyse data

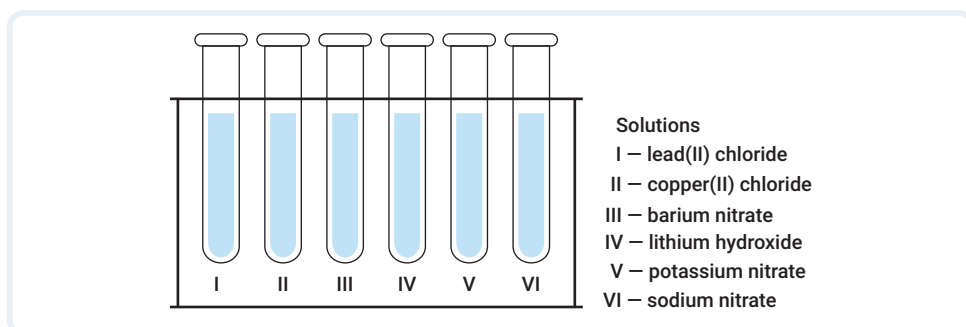
A group of Chemistry students were conducting an experiment to determine the presence of precipitates when solutions were mixed. Unfortunately some of the labels had come off the solutions that they used. Unknown solutions X and Y were either sodium chloride or iron(II) iodide. They were mixed with silver nitrate and potassium carbonate. The students recorded if a precipitate formed, and used Table 15.1.2 to help them decide which solution was which. The results for the experiment were as follows:

Solution	Silver nitrate	Potassium carbonate
Unknown X	Precipitate formed	Precipitate formed
Unknown Y	Precipitate formed	No precipitate
Barium chloride		

- Identify** unknown solutions X and Y.
- Determine** the chemical formula for the precipitate formed when unknown X was mixed with potassium carbonate. Use the results table and Table 15.1.2 to help with your determination.
- Predict** whether a precipitate will form if barium chloride is mixed with silver nitrate and potassium carbonate to complete the results table.

15. Interpret evidence

A group of Chemistry students decided to perform an experiment to test which combination of solutions would form a precipitate with ammonium sulfate solution. The set-up of their experiments were as follows.



- Identify** the two solutions that formed precipitates with ammonium sulfate.
- Classify** the products formed in test tube V as soluble or insoluble.
- Determine** the name of the products formed in test tube III.
- Compare** the results of test tube I and test tube II, in terms of the solubility of the products formed.
- Predict** the results of the experiment for each test tube if a solution of manganese(II) sulfate was used instead of ammonium sulfate solution.



Miriam Gil Albert/Shutterstock.com

SYLLABUS
DOT POINTS**SCIENCE UNDERSTANDING**

- 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 the 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 INQUIRY

Investigate:

- factors that affect solubility in aqueous solutions.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

This chapter examines the relationship between solubility and temperature for different substances, including ionic and organic compounds, as well as gases. Temperature plays a role in the solubility of substances, and this may be represented by a solubility curve. Solubility curves are used to determine whether a solution is unsaturated, saturated or supersaturated at a given temperature. For solutions of solid compounds, we use solubility curves to predict the amount of crystallisation that can occur when a solution cools. It also examines the role of intermolecular forces in determining solubility of molecular substances in water.

Practicals

- Factors affecting solubility – polarity of solute
- Effect of temperature on solubility
- Deriving the solubility curve of ammonium chloride
- Disappearing act - dissolving processes of liquids (online-only resource)

Worksheets

- Understanding solubility
- Solubility of different types of substances
- Vitamins: good or bad?



To access resources above, visit
cengage.com.au/nelsonmindtap



ASSUMED KNOWLEDGE

- ✓ The polarity of bonds within a molecule and the shape of the molecule affect its overall polarity and thus affect the intermolecular forces between molecules.
- ✓ Ionic and covalent molecular substances have different properties.
- ✓ Water is a good solvent owing to its polarity.
- ✓ Molarity can be calculated using $c = \frac{n}{V}$.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ explain why most ionic substances dissolve in water
- ✓ explain why some molecular substances dissolve in water and others do not
- ✓ relate trends in molecular size to solubility in water
- ✓ explain general trends in solubility of ionic solids, molecular compounds and gases with temperature
- ✓ analyse solubility curves to determine how much solute will dissolve, and whether a solution is saturated at a given temperature
- ✓ calculate the solubility of substances or concentration of solutions at a given temperature.

16.1 Solubility and intermolecular bonding

In Chapter 14, we identified that a solution forms when a solute dissolves in a solvent, producing a homogeneous mixture in which particles of solid solute can no longer be observed. When this occurs, the forces that have been holding together the solute particles have been overcome, and new interactions form between the water and the solute particles, causing them to mix. As we already know, not all substances are soluble in water. This varying solubility of different substances with water, or indeed any other solvent, can be explained by comparing the relative strengths of these interactions.

The process of dissolving both requires and releases energy. When water dissolves a substance, the water molecules attract the particles (molecules or ions) of the substance, and if the attraction is strong enough, the particles separate from each other. The strong attraction is due to water's polarity. It takes energy to break the bonds between the molecules or ions of the solute. However, energy is released when water molecules 'bond' to the solute molecules or ions as the substance dissolves.

If it takes more energy to separate the particles of the solute than is released when the water molecules bond to the particles, then the temperature decreases and the dissolving process is endothermic. If it takes less energy to separate the particles of the solute than is released when the water molecules bond to the particles, then the temperature increases and the dissolving process is exothermic.

Disposable hot and cold packs work on this principle (**Figure 16.1.1**). They are generally made of a small vial of a salt contained within a bag of water. When the vial is broken, the contents are released. Ammonium nitrate, commonly found in cold packs, requires more energy to dissolve than it releases, so the energy is taken in from the surroundings and the pack

becomes cold. Alternatively, the chemicals in heat packs (usually magnesium sulfate or calcium chloride) release more energy when forming new bonds than required to break old bonds. They become hot as the bond energy is released as heat.



Carolyn Jenkins/Alamy Stock Photo

FIGURE 16.1.1 Instant hot and cold packs contain small vials of salt that, when broken, dissolve in water. This causes a temperature change.

Four main factors affect the solubility of substances:

- molecular polarity
- temperature
- pressure (for gas solubility)
- molecular size.

Molecular polarity

A variety of substances dissolve in water. The types of bonds that exist when a solution forms depend on the structures of the solute. As mentioned in Chapter 14, a simple way to predict which compounds will dissolve in other compounds is the phrase ‘like dissolves like’. Polar solvents dissolve polar and ionic solutes, and non-polar solvents dissolve non-polar solutes.

As a polar solvent, water can form different intermolecular bonds, including ion–dipole bonds with ionic substances, and dipole–dipole and hydrogen bonds with other molecules that are polar. This is why water can dissolve various types of compounds, both ionic and covalent in structure. The following section explains the mechanism of dissolving for ionic and molecular compounds.

Solutions of ionic compounds

When an ionic salt such as sodium chloride dissolves in water, it dissociates to form ions. These substances are known as **electrolytes** because the ions can carry a charge in solution. When a crystal of sodium chloride is placed into water, the water’s molecules collide with the crystal lattice. **Ion–dipole bonds** are formed between the water molecules and the ions. The force of attraction to water molecules pulls each ion away from the crystal lattice. The individual ions are then surrounded by water molecules.

The anion attracts the positive end of the water molecule. More water molecules surround the anion until it is perfectly hydrated. The cation attracts the negative end of the water molecule, and a similar process occurs until the ionic bonds within the salt crystal are broken and the salt has dissolved.

electrolyte a solution that can conduct electricity because of the presence of ions

ion–dipole bond a bond formed by attraction between a charged ion and the dipole of a polar molecule

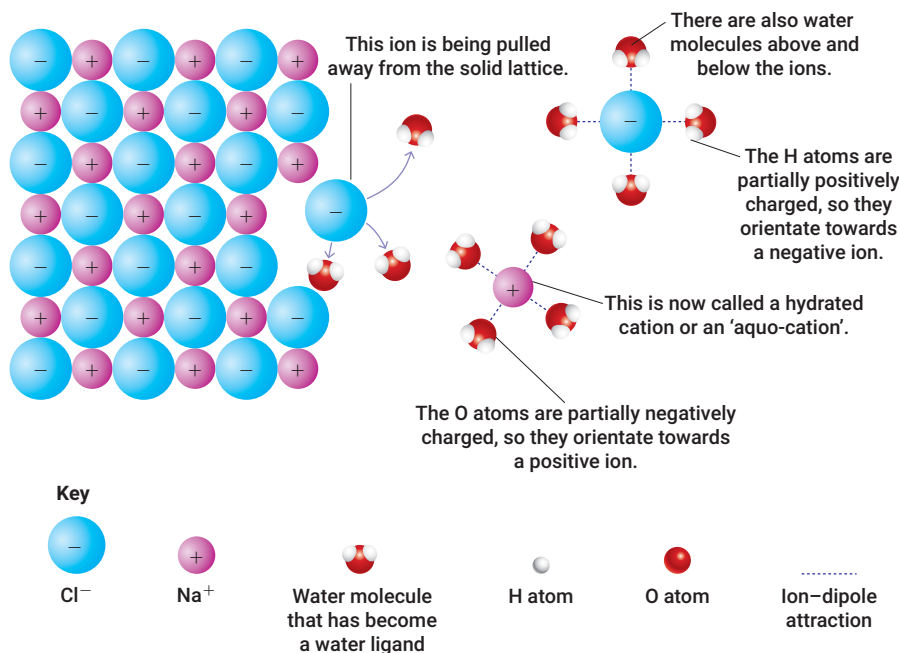


FIGURE 16.1.2 A simplified model of the solution process for a sodium chloride crystal in water



Weblink

Why are ionic compounds soluble in water?

In **Figure 16.1.2**, ion-dipole bonds are formed between water and both the Na^+ and Cl^- ions. The water molecules orient themselves so that the negative dipole (oxygen end) is attracted to the positive Na^+ ions to form ion-dipole bonds. The water molecules orient themselves the other way around near Cl^- ions, so the positive dipole (the hydrogen end) of water is attracted to the negative Cl^- ions to form ion-dipole bonds.

However, not all ionic compounds are soluble in water. You have seen this in Chapter 15 when you used the solubility rules and solubility table to determine whether a substance is soluble or insoluble in water. Also, some substances are partially soluble, meaning they do not completely dissociate but small amounts may dissolve. The amount of energy required and released in the dissolving process plays a role in solubility because of the strength of the forces of attraction between the solute and water, and within the substance itself. **Figure 16.1.3** clearly outlines the role of energy factors in controlling the solubility of ionic compounds.

Solutions of molecular compounds

Unlike ionic substances, molecular substances generally do not dissociate when they dissolve. The dissolved solute particles of molecular substances are not ions but are entire molecules. For example, when sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) dissolves, the molecules do not dissociate but remain as individual sugar molecules surrounded by water. These substances are known as **non-electrolytes**, because they are not capable of carrying a charge in solution.

A sugar crystal consists of sugar molecules held together by intermolecular hydrogen bonding forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the hydrogen bonding forces between molecules. It does not break apart the covalent bonds between atoms within the molecule because these are much stronger than the hydrogen bonds. The attractive forces between water molecules and the sugar molecules must be stronger than the attraction between the sugar molecules in the crystal for dissolving to occur. **Figure 16.1.4** shows how each individual sugar molecule has intermolecular forces of attraction to water molecules and breaks away from the other individual sugar molecules in the sugar crystal.

non-electrolyte a substance that does not form ions when it dissolves in water, so the resulting solution does not conduct electricity

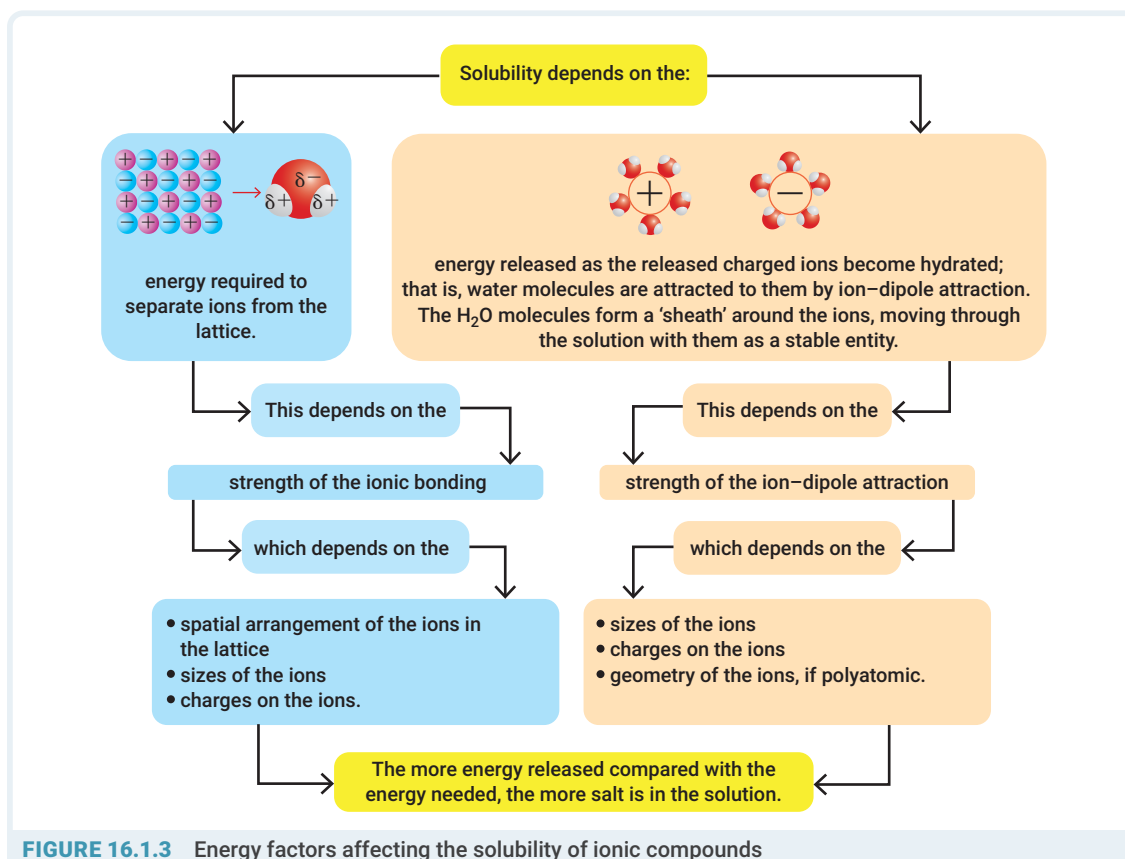


FIGURE 16.1.3 Energy factors affecting the solubility of ionic compounds

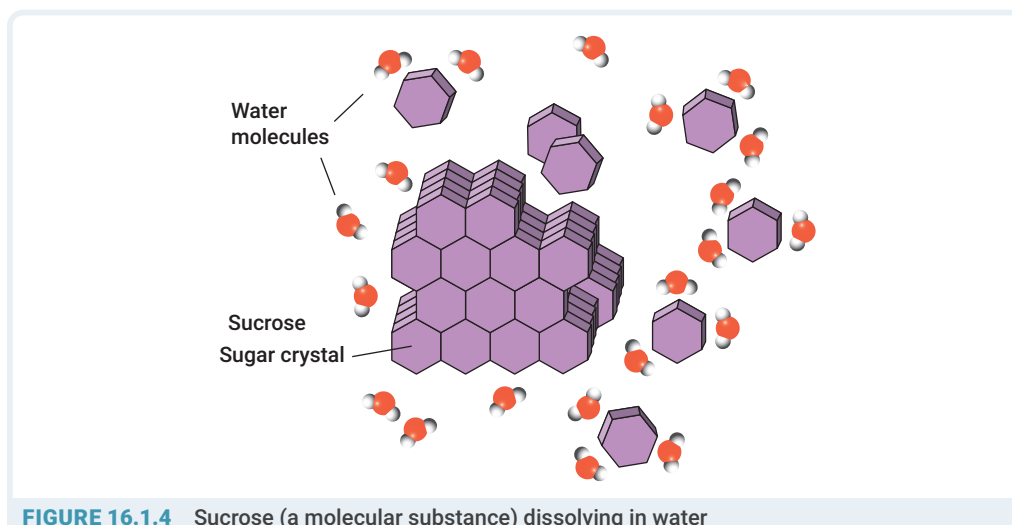


FIGURE 16.1.4 Sucrose (a molecular substance) dissolving in water



Weblinks
Solubility and
intermolecular forces

Solubility of
hydrocarbons

Note that there are other factors that affect dissolving that are not discussed in this chapter, such as how the dissolving of some salts is endothermic.

Not all molecular substances are soluble. Most organic molecules are non-polar, containing non-polar covalent bonds. Due to this, dispersion forces are the only type of intermolecular bonding that exists between them. Because they are non-polar, a strong attraction to the highly polar water molecules does not occur, so these molecules are insoluble. They are referred to as **hydrophobic**.

However, when organic molecules have polar groups attached to a carbon chain, their hydrophobic nature decreases. The polar groups can form hydrogen bonds and dipole-dipole

hydrophobic 'water-hating', not attracted to the water molecule



Worksheet
Understanding solubility

Practical activity
Disappearing act –
dissolving processes
of liquids

interactions with water molecules, increasing the solubility of the molecule as a whole. Examples of functional groups that give polarity include —OH , C=O and —NH . You will learn more about functional groups in Unit 4.

Table 16.1.1 shows the solubilities of a series of different alkanol molecules. These are molecules with a non-polar hydrocarbon chain and a polar —OH bond.

TABLE 16.1.1 The solubility of alkanols in water

Name	Molecular formula	Solubility in water (g/100 mL)
Methanol	CH_3OH	Infinite
Ethanol	$\text{C}_2\text{H}_5\text{OH}$	Infinite
Propanol	$\text{C}_3\text{H}_7\text{OH}$	Infinite
Butanol	$\text{C}_4\text{H}_9\text{OH}$	8.0
Pentanol	$\text{C}_5\text{H}_{11}\text{OH}$	2.2
Hexanol	$\text{C}_6\text{H}_{13}\text{OH}$	0.7

Molecular size

It is evident from Table 16.1.1 that larger molecules with more carbon and hydrogen atoms (and the increasing number of non-polar C—C and C—H bonds) have lower solubility of the overall molecule in water. This is due to the non-polar nature of these bonds that only have weak dispersion forces of attraction to the water molecule.

If larger organic molecules, such as glucose, vitamins, enzymes and hormones, have enough polar groups, then they may be polar enough to dissolve in water, because the polar groups have dipole–dipole attractions and hydrogen bonding, both of which form bonds with the water molecule.

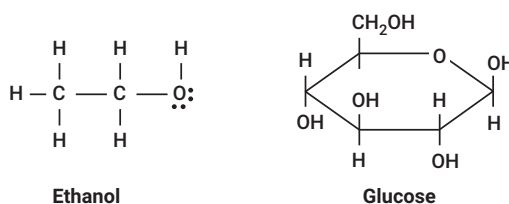
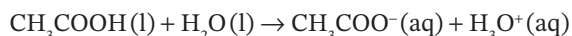


FIGURE 16.1.5 Ethanol and glucose molecules are polar because of the presence of polar groups.

It should be noted that there are also some organic molecules, such as carboxylic acids, that will ionise to form ions when added to water. These molecules can form additional ion–dipole bonds with water molecules, so they have significantly greater solubility than other organic molecules that do not ionise when added to water. This can be represented by the equation below:



Ethanoic acid

The positive hydronium (H_3O^+) and negative ethanoate ion (CH_3COO^-) formed can now form ion–dipole bonds with water molecules and are highly soluble.

Covalent network molecules are insoluble because the covalent bonds between atoms throughout the network are stronger than the forces of attraction to the water molecules.

Table 16.1.2 is a summary of the general trends in solubility of different molecular structures.

TABLE 16.1.2 The solubility of various types of substances in water

Bonding type	Solubility in water	Examples
Ionic	Most are soluble	NaCl, CuSO ₄
Covalent molecular	Soluble if hydrogen bonds and dipole–dipole interactions possible Soluble by reacting with water Otherwise insoluble	Ethanol, glucose HCl, HNO ₃ , NH ₃ Dichloromethane (CH ₂ Cl ₂)
Non-polar molecular	Most are insoluble Partially soluble if water attracted to atoms	Hydrocarbons (e.g. hexane) O ₂ , I ₂
Very large molecules	Insoluble if highly structured Soluble if large proportion of polar groups	Cellulose (fibre) Starch, glycogen, enzymes
Network covalent substances	Insoluble	Diamond, SiO ₂
Metals	Insoluble Unless they react with water	Pb, Au Li, Na, K, Ca, Ba

**Worksheets**

Solubility of different types of substances

Vitamins: good or bad?

PRACTICAL ACTIVITY 16.1.1**FACTORS AFFECTING SOLUBILITY – POLARITY OF SOLUTE****Introduction**

The polar nature of the water molecule is largely responsible for its ability to dissolve a range of substances. Water is able to dissolve ionic substances and polar covalent molecules. Non-polar molecules are not able to dissolve in water. In this experiment, you will examine how substances of different polarity are soluble or insoluble in water.

Research question

How does the polarity of a substance affect its ability to dissolve in water?

Aim

To test the solubility in water of a variety of substances with different polarities

Materials

- sodium chloride (NaCl)
- potassium sulfate (K₂SO₄)
- sucrose (sugar)
- calcium carbonate (CaCO₃)
- sodium thiosulfate (Na₂S₂O₃)
- ethanol (C₂H₅OH)
- hexane (C₆H₁₄)
- glycerin (C₃H₈O₃)
- spatula
- 3 droppers
- 8 test tubes
- 2 test-tube racks
- 5 mL measuring cylinder
- plastic wash bottle
- distilled water
- paper towel

Procedure

- 1 Add 3–4 cm of distilled water to each test tube.
- 2 Add a tiny amount of a substance to be tested. For solids, use an amount the size of a match head. For liquids, use one drop.
- 3 Mix the substance with the water by flicking the test tube gently. If the substance dissolves, add another tiny amount and mix gently. Repeat the process if the substance continues to dissolve.

- 4 Describe each substance as insoluble, partially soluble or soluble based on the observed results. Record these observations in your results table.
- 5 Repeat for all of the substances.

Results

Copy and complete the results table.

Substance	Formula	Ionic/polar covalent/ non-polar covalent	Insoluble/partially soluble/soluble
Sodium chloride			
Potassium sulfate			
Sucrose (sugar)			
Calcium carbonate			
Sodium thiosulfate			
Ethanol			
Hexane			
Glycerin			

Analysis of results

- 1 Identify whether each substance was insoluble, partially soluble or soluble, based on your observations.
- 2 Identify whether there is a relationship between the type of substance or polarity of the substance and whether it is soluble.

Interpretation

- 3 Compare the solubilities of the polar covalent substances with the solubilities of the ionic substances.
- 4 Determine if there is a relationship between the polarity of a substance and how soluble it is.

Liquid molecular substances, such as alcohols, can also dissolve in water. In a similar manner to the dissolution of solids, the molecules of a liquid are also separated and hydrated (surrounded by water molecules as a result of forces of attraction). This may affect the final volume of the resultant solution, as the forces of attraction between the liquid and the water may make the resultant solution more dense than the original liquids.

LEARNING CHECK 16.1

DESCRIBING

- 1 **Describe**, using a diagram, how sodium chloride (NaCl) dissolves in water.
- 2 **Recall** the name given to the forces of attraction between ions and water molecules when an ionic substance dissolves.
- 3 **Identify** the correct form that sugar takes when it is dissolved in water as:
 - A dissociated ions.
 - B charged particles.
 - C sugar molecules broken apart.
 - D individual sugar molecules surrounded by water.

APPLYING

- 4 Hydrogen bromide and methane are small covalent molecules. **Explain** why methane does not dissolve in water but hydrogen bromide does.
- 5 **Describe** what happens when sugar is added to a cup of coffee. **Identify** the solute and solvent.
- 6 KBr dissolves in water by dissociation.
- Explain** what this statement means.
 - Write an ionic equation to **describe** what happens.
 - Classify** this solution as an electrolyte or a non-electrolyte.
- 7 **Describe** what happens on a molecular level when water and ethanol mix. Explain, using the terms 'solute' and 'solvent'. **Identify** which bonds are broken and which bonds are formed. Write an equation to show this.

ANALYSING

- 8 **Deduce** why cerium sulfate has a low solubility in water, despite being ionic.
- 9 **Identify** whether the following molecular structures are generally soluble or insoluble.
- Ionic substances
 - Non-polar covalent molecules
 - Polar covalent molecules
 - Large covalent molecules with no polar groups
 - Large covalent molecules with many polar groups
 - Network covalent molecules
- 10 **Classify** the following as either soluble or not soluble in water.
- | | |
|-----------------------------|---|
| a HBr | b MgCl_2 |
| c CS_2 | d ZnCl_2 |
| e H_2SO_4 | f Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) |
| g C_8H_{18} | h NH_2CONH_2 |
- 11 **Classify** each of these as an electrolyte or a non-electrolyte, given that each dissolves in H_2O to some extent.
- | | |
|---|-------------------|
| a $\text{CH}_3\text{CH}_2\text{OH}$ (ethanol) | b BaCl_2 |
| c CaCl_2 | d I_2 |
| e KOH | |

INTERPRETING

- 12 A student obtained samples of sodium chloride, copper sulfate, ethanol and an unknown substance. The student conducted an experiment in which they dissolved the substances in deionised water and then tested the electrical conductivity of the resultant solution. The results of the experiment are outlined below.

Substance	Soluble in water? Y/N	Electrical conductivity in aqueous solution? Y/N	Type of substance? Ionic/molecular
Sodium chloride	Y	Y	Ionic
Copper sulfate	Y	Y	
Ethanol	Y	N	Molecular
Unknown	Y	N	

- Classify** copper sulfate as ionic or molecular, based on these results.
- Classify** the unknown substance as ionic or molecular, based on these results.
- Determine** which of these substances would form ion-dipole interactions when dissolved in water.
- Determine** which of these substances exists as individual molecules surrounded by water molecules when dissolved.
- Explain** why sodium chloride conducts electricity in aqueous solution, whereas ethanol does not.

16.2 The effect of temperature on solubility

Solubility can be measured as the amount of solute dissolved in a fixed amount of water. A common unit used is g/100 g – the mass of solute dissolved in 100 g of water. The solubility of a substance depends on temperature. Exactly how solubility varies with temperature depends on other factors, including the state of the solute.

Temperature dependence of the solubility of solids

The solubility of most solids in water increases with temperature.

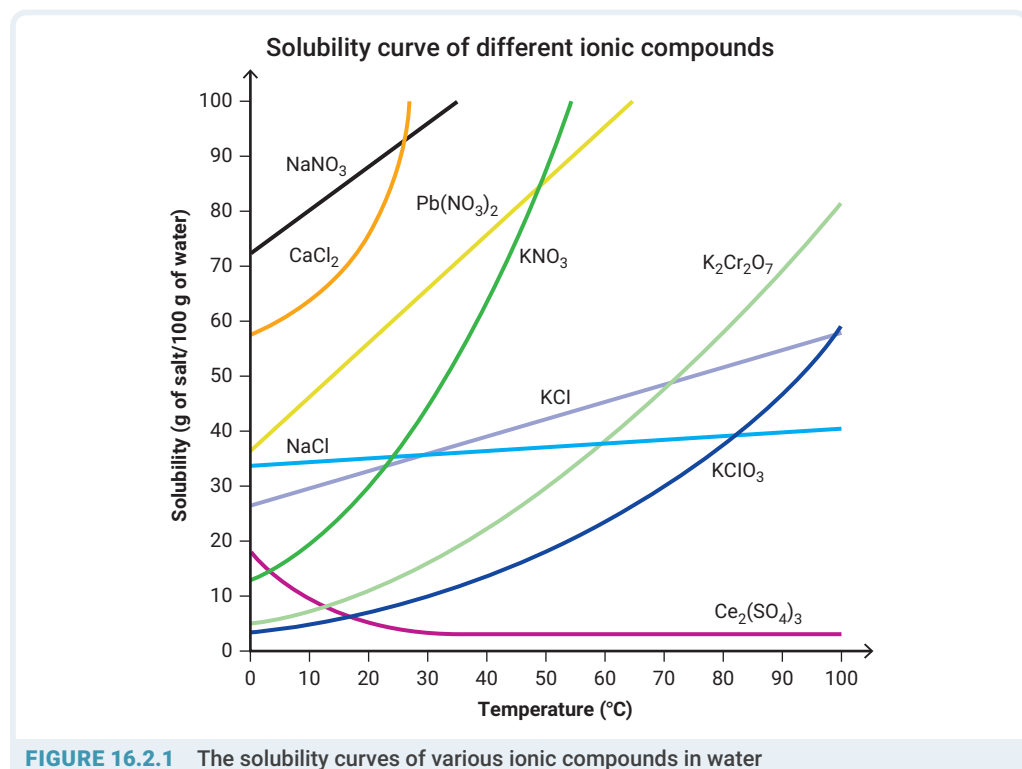
Increasing the temperature can have a marked effect on the solubility of substances, because there is consequent greater increase in the energy of the particles, and the solvent particles have more energy to break bonds in the solute.

When heated, solvent molecules collide with solute molecules with a greater impact, and effectively break apart the pieces of solute and overcome the attractive forces that hold them together. In addition, when the solute atoms are moving more energetically, they are less likely to hold together, so they dissolve more readily.

It is important to note that not all solids increase in solubility with an increase in temperature. Some substances actually decrease in solubility with an increase in temperature. This is also the case for the solubility of gases, which will be discussed later in this chapter.

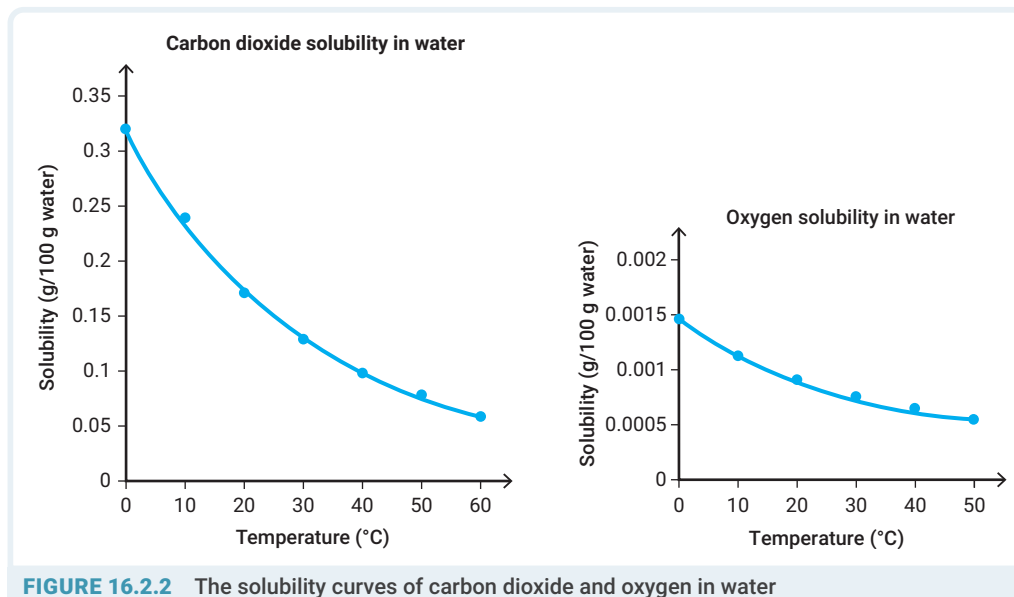
A **solubility curve** is a graph that shows the variation of solubility with temperature. **Figure 16.2.1** shows solubility curves for several ionic compounds. As shown, the solubility of different compounds increases at markedly different rates with temperature, and the solubility of cerium sulfate ($\text{Ce}_2(\text{SO}_4)_3$) actually decreases with temperature. This is due to the varying strength of the ionic bonds inside the ionic solids, which need to be overcome, and the variable nature of the ion-dipole interactions between the ions and water molecules that form during the process of dissolving.

solubility curve a graph showing the variation of the solubility of a substance in water with temperature



Solubility of gases in water

Gases are less soluble than solids in water. In addition, as the temperature is increased, the solubility of gases decreases rather than increases. **Figure 16.2.2** demonstrates this trend for the gases carbon dioxide and oxygen.



This apparently contradictory trend is explained again by considering the forces that need to be overcome and the interactions that form when gases dissolve. The forces of attraction between the molecules of most gases are negligible. When the bonds with water molecules form, heat is released in an exothermic process. An increase in temperature increases the kinetic energy of the gas particles, allowing them to move faster. This increase in movement may allow gas particles to break the forces between the gas particles and water, reducing the solubility of the gas. Furthermore, the increase in temperature has the effect of reversing the exothermic dissolution process and re-forming the gas. This concept is further explored in Unit 3.

Some of the most biodiverse oceans of the world are off the coast of Antarctica, where the water is very cold, and the solubility of oxygen in the seawater is high. One current concern of climate scientists is that increasing atmospheric temperatures cause similar changes to the temperature of seawater, decreasing the amount of oxygen dissolved, with potentially large consequences for organisms living in the sea (**Figure 16.2.3**).



FIGURE 16.2.3 The oceans around Antarctica are very rich in biodiversity.

Pressure and gas solubility

The higher the pressure, the more soluble the gas. This can be observed when opening a bottle of carbonated drink. When the pressure in the bottle is decreased as it is opened, the gas that was dissolved in the drink bubbles out of it (**Figure 16.2.4**).



FIGURE 16.2.4 The solubility of gases decreases as pressure decreases.

PRACTICAL ACTIVITY 16.2.1

EFFECT OF TEMPERATURE ON SOLUBILITY

Introduction

Temperature can have different effects on the solubility of substances, depending on the nature of the substance and whether the substance being dissolved is a solid or a gas.

Research question

How does temperature affect the solubility of potassium nitrate?

Aim

To determine the effect of changing temperature on the solubility of potassium nitrate

Materials

- 20 g potassium nitrate
- deionised water
- 4 test tubes and a test-tube rack
- stirring rod
- hot plate
- thermometer
- boss-head and clamp
- retort stand
- 10 mL measuring cylinder
- 400 mL beaker
- electronic scales
- spatula
- watch glass or weighing boat



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
The hot plate can burn.	Use tongs or clamps to handle the hot test tubes and beaker. Do not touch the hot plate while it is hot.
Broken glassware will cut.	Inspect and discard any chipped or cracked glassware.

Procedure

- 1 Label the test tubes 1–4.
- 2 Add potassium nitrate to each test tube as outlined in the table.

Test tube	Mass of KNO_3 (g)
1	2.0
2	4.0
3	6.0
4	8.0

- 3 Add 5.0 mL of deionised water to each test tube.
- 4 Add 300 mL of water to the 400 mL beaker.
- 5 Place the beaker full of water on a hot plate. Heat the water to 90°C (do not let it boil). Use the thermometer to maintain the temperature of the water at this temperature.
- 6 Attach test tube 1 to the clamp that is connected to the retort stand. Lower the test tube into the beaker of water.
- 7 Use the stirring rod to stir the mixture in the test tube until it dissolves.
- 8 When the KNO_3 is completely dissolved, add the thermometer into the test tube. Move the clamp upwards to remove the test tube from the water.

9 Record the temperature at which crystals start to form.

10 Repeat steps 6–9 for test tubes 2–4.

Results

Copy and complete the following table.

Test tube	Mass of KNO_3 (g) in 5 mL of water	Mass of KNO_3 (g) in 100 g of water	Temperature ($^{\circ}\text{C}$)
1			
2			
3			
4			

Analysing results

1 Use the data in the table to describe what happens to the temperature as the mass increases.

Interpretation

2 Define 'supersaturated' and 'saturation' in relation to describing solutions.

3 Explain why crystals form in each test tube.

4 Describe the relationship between the mass of KNO_3 dissolved and the temperature at which the crystals start to form.

Evaluation

5 Identify two possible sources of error in your experiment.

LEARNING CHECK 16.2

DESCRIBING

1 **Define:**

- a temperature
- b solubility curve.

APPLYING

2 **Contrast** the general trends in solubility for most solids and for gases with an increase in temperature.

3 **Explain** why increasing the temperature can increase the solubility of a solid in water.

ANALYSING

4 **Discuss** why scientists are concerned about the effect that rising sea temperatures, caused by climate change, might have on oxygen levels in Antarctic marine environments.

INTERPRETING

5 **Compare** the trends in solubility of carbon dioxide and oxygen in water, using Figure 16.2.2.

6 **Determine** whether heat affects the solubility of all ionic solids in the same way, using examples from Figure 16.2.1 to illustrate your answer.

16.3 Interpreting and analysing solubility curves

We can use information from solubility curves to determine predictive information about solutions, and the quantity of solute that will dissolve at a particular temperature. Solubility curves can also be used to determine the concentration of the solution when the solution is saturated, by using the key formulas from Chapters 9 and 14 involving molarity and mol.

KEY FORMULA

$$c = \frac{n}{V} \text{ and } n = \frac{m}{M}$$

WORKED EXAMPLE 16.3.1

Refer to the solubility curve in **Figure 16.3.1**.

- Identify** how much potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) will dissolve at 20°C .
- Calculate** the concentration of the saturated potassium dichromate solution at 20°C .
- Determine** the number of moles of dissolved K^+ ions present in solution at 20°C .
- Determine** how much extra potassium dichromate will dissolve at 80°C .

ANSWERS

- 1 Find the curve for $\text{K}_2\text{Cr}_2\text{O}_7$ on the graph.**
On Figure 16.3.1, it is the blue curve.
- 2 Draw a line from the x-axis.**
Draw from 20°C up to the potassium dichromate curve.
- 3 Draw a line from this point across to the y-axis.**
This indicates the amount that will dissolve at 20°C (Figure 16.3.1).
Therefore, 12 g potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) will dissolve at 20°C .
- 1 Find the solubility at 20°C .**
The solubility of $\text{K}_2\text{Cr}_2\text{O}_7$ is 12 g/100 g water, which equals 12 g/100 mL water.
- 2 Use the solubility to calculate the amount, in mol, of $\text{K}_2\text{Cr}_2\text{O}_7$.**
Convert g mL^{-1} to g L^{-1} .
 12 g/100 mL (multiply by 10 to get to 1000 mL) = $120 \text{ g/1000 mL} = 120 \text{ g L}^{-1}$.

Calculate number of moles of $\text{K}_2\text{Cr}_2\text{O}_7$ using the formula $n = \frac{m}{M}$.

The molar mass of $\text{K}_2\text{Cr}_2\text{O}_7$ is $(2 \times 39.10) + (2 \times 52.00) + (7 \times 16.00) = 294.2 \text{ g}$

The number of moles of 120 g $\text{K}_2\text{Cr}_2\text{O}_7$ is $\frac{120}{294.2} = 0.41 \text{ mol}$

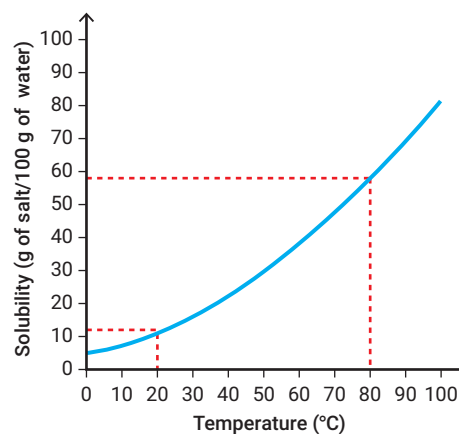


FIGURE 16.3.1 The amounts of potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) that can dissolve at 20°C and at 80°C . The difference is how much extra potassium dichromate will crystallise out of the solution on lowering the temperature from 80°C to 20°C .

4 Calculate the concentration of the solution.

$$\begin{aligned}\text{The concentration of K}_2\text{Cr}_2\text{O}_7 \text{ is } c &= \frac{n}{V} \\ &= 0.41 \text{ mole L}^{-1} = 0.41 \text{ M}\end{aligned}$$

At 20°C, the concentration of a saturated solution of $\text{K}_2\text{Cr}_2\text{O}_7$ is 0.41 M.

c 1 Determine the relationship between K^+ ions and $\text{K}_2\text{Cr}_2\text{O}_7$.

For every 1 mol of $\text{K}_2\text{Cr}_2\text{O}_7$, there are 2 mol of K^+ ions.

2 Calculate the number of moles of K^+ present.

$$\text{Number of moles of K}^+ = 2 \times 0.41 = 0.82 \text{ mol}$$

d 1 Find out how much $\text{K}_2\text{Cr}_2\text{O}_7$ dissolves at both temperatures.

58 g of the substance will dissolve in 100 g of water at 80°C.

At 20°C, 12 g of the substance will dissolve in 100 g of water.

2 Calculate the difference in solute dissolved.

$$58 - 12 = 46 \text{ g}$$

This is also the amount that would crystallise out of solution if the temperature of the saturated solution were cooled from 80°C to 20°C.

The extra amount that would dissolve at 80°C is 46 g.

PRACTICAL ACTIVITY 16.3.1

DERIVING THE SOLUBILITY CURVE OF AMMONIUM CHLORIDE

Introduction

Solubility curves are drawn from experimental data, where successive heating and cooling of solutions is used to determine the maximum amount of solute that can be dissolved at different temperatures.

Research question

How does temperature affect the solubility of potassium nitrate?

Aim

To gather data to draw a solubility curve of ammonium chloride

Materials

- ammonium chloride (NH_4Cl)
- distilled water
- 250 mL beaker
- 10 mL measuring cylinder
- large test tubes
- Bunsen burner
- tripod
- gauze mat
- bosshead and clamp
- 0–100°C thermometer
- stirring rod
- weighing container
- electronic balance



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Ammonium chloride is slightly toxic.	Work in a well-ventilated area. Wear safety glasses and wash your hands after the experiment. Dispose of all chemicals according to your teacher's directions. Do not pour them down the sink.
The Bunsen burner will get hot.	Do not use the Bunsen burner if the gas tube is damaged. Ensure long hair is tied back and the flame is away from flammable material. If you burn yourself, rinse the affected area under cold running water for 10 minutes and inform your teacher.
Broken glassware will cut.	Inspect and discard any chipped or cracked glassware, no matter how small the damage. Sweep up broken glass with a brush and dustpan; do not use your fingers.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the three listed risks to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- 1 Carefully weigh out the amount of ammonium chloride assigned to your group (4.00, 4.50, 5.00, 5.50, 6.00 or 6.50 g). Groups will share data at the end of the experiment.
- 2 Place the weighed ammonium chloride into a test tube.
- 3 Add 10 mL distilled water to the test tube.
- 4 Half-fill a beaker with tap water. Place the beaker on the gauze mat on a tripod over a Bunsen burner. Clamp the test tube securely into the beaker so that it is immersed in the water but not touching the bottom of the beaker (see [Figure 16.3.2](#)).

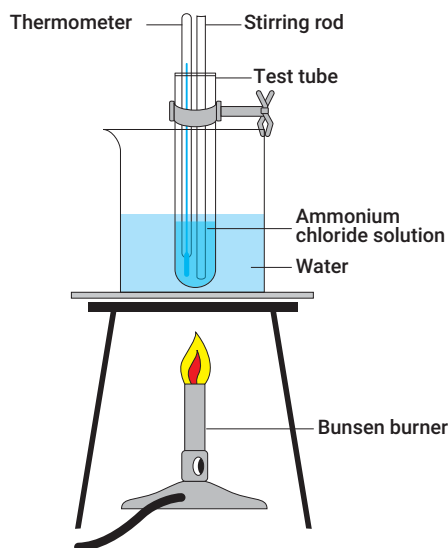


FIGURE 16.3.2 The experimental set-up for determination of a solubility curve

- 5 Warm the beaker, stirring the contents of the test tube with the stirring rod. Continue until the ammonium chloride has dissolved completely.
- 6 Place the thermometer in the test tube and allow the test tube to cool. Continue to stir it until tiny flakes of ammonium chloride start to appear. Note the temperature at which this occurs.

Results

Copy and complete the following table.

Group	Mass of ammonium chloride (g)	Mass of water (g)	Temperature of recrystallisation (°C) (x-axis)	Solubility (g/100 g H ₂ O) (y-axis)
1	4.00	10.00		
2	4.50	10.00		
3	5.00	10.00		
4	5.50	10.00		
5	6.00	10.00		
6	6.50	10.00		

Analysis of results

- 1 Use the class results to plot a graph of solubility against temperature. Plot the temperature along the horizontal x-axis (0–100°C).
- 2 From the graph, predict the solubility of ammonium chloride at 20°C, 40°C, 60°C and 80°C.

Interpretation

- 3 Describe what happens to the solubility of ammonium chloride as temperature increases.
- 4 At the point the crystals first appear, the solution is saturated. Explain why you know the solution is saturated at this point.
- 5 Identify the temperature at which 101 g of ammonium chloride would saturate 50 g of water.
- 6 Calculate how many grams of ammonium chloride would crystallise out of 100 g of water if a saturated solution of ammonium chloride at 90°C is cooled to 10°C.

Evaluation

- 7 Calculate the percentage error if the theoretical value for the solubility of ammonium chloride at 50°C is 50 g/100 g
- 8 Identify possible sources of errors in your experiment.
- 9 Describe modifications to the experiment that could improve the accuracy.

LEARNING CHECK 16.3

Where required, use the solubility curves in Figure 16.2.1 (page 405) to answer the following questions.

DESCRIBING

- 1 **Recall** what the line on a solubility curve shows.

APPLYING

- 2 **Identify** the mass of the following solutes that can dissolve in 50 mL of water.
- a KClO_3 at 65°C
 - b CaCl_2 at 25°C
 - c $\text{Pb}(\text{NO}_3)_2$ at 55°C
- 3 **Classify** a solution of 45 g/100 g of KCl at 70°C as unsaturated, saturated or supersaturated.
- 4 **Determine** which of $\text{K}_2\text{Cr}_2\text{O}_7$, KCl and $\text{Ce}_2(\text{SO}_4)_3$ is most soluble in water at 15°C .
- 5 **Identify** the maximum mass of KNO_3 that could be dissolved in 25 mL of water at 40°C .

ANALYSING

- 6 Suppose 200 mL of a saturated solution of $\text{K}_2\text{Cr}_2\text{O}_7$ is made at 70°C . **Calculate** the mass of solid that would precipitate by the cooling of this solution to 30°C .
- 7 **Identify** the mass of the following solutes that will dissolve in 100 mL of water.
- a $\text{K}_2\text{Cr}_2\text{O}_7$ at 50°C
 - b NaCl at 100°C
 - c NaNO_3 at 10°C
- 8 a **Identify** the mass of KCl and NaCl that would just dissolve in 100 mL of water at 30°C .
- b **Classify** the solution as saturated, unsaturated or supersaturated for this mass at this temperature.
- 9 **Determine** the mass of potassium nitrate (KNO_3) that will dissolve in:
- a 100 g of water to form a saturated solution at 20°C
 - b 50 g of water to form a supersaturated solution at 30°C
 - c 80 g of water to form a saturated solution at 10°C .
- 10 **Calculate** the concentration of the solutions formed in Question 9.

INTERPRETING

- 11 a **Explain**, using diagrams, how KCl dissolves in water.
- b **Determine** a balanced equation for the dissolution of KCl in water.
- c **Identify** how much KCl would dissolve in 50 mL of water at 80°C .
- d **Classify** the solution in part c if the solution were cooled to 70°C and no precipitate formed. **Justify** your answer.
- e i **Identify** the temperature you would need to heat 100 g of water so that 40 g of KCl dissolves completely.
- ii **Determine** how much KCl you would expect to precipitate out of the solution if a saturated solution at 60°C were cooled to 10°C .

12 Consider the solubility curves for potassium nitrate, lead(II) nitrate and sodium chloride (Figure 16.3.3).

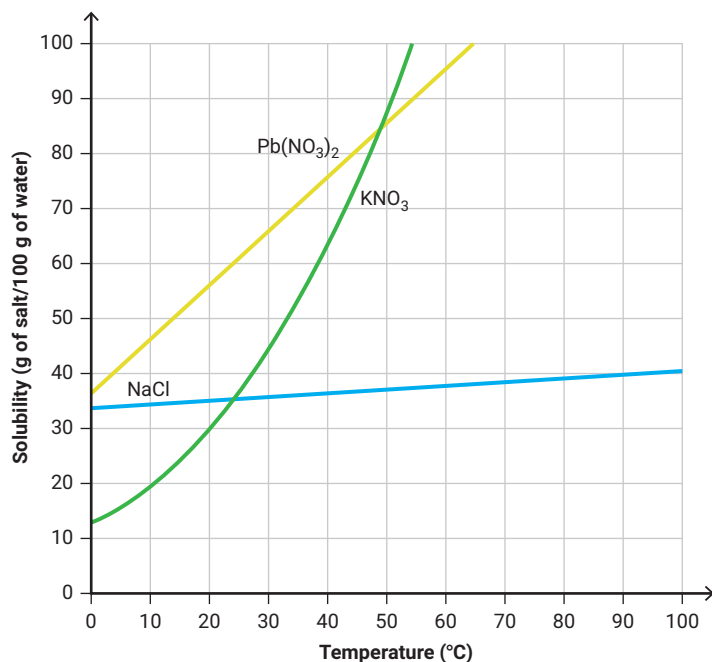


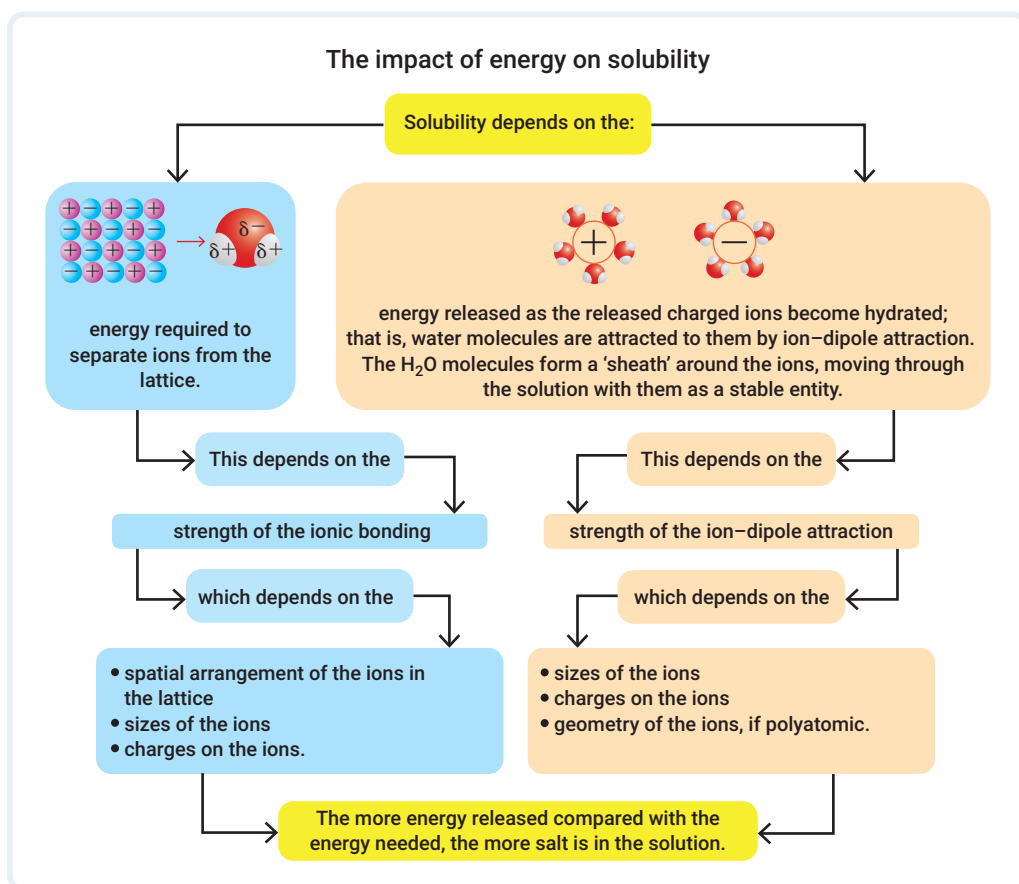
FIGURE 16.3.3 Solubility curves for three ionic substances

- Identify** the trend shown by all three solutes in this graph.
- Identify** which substance experiences the greatest increase in solubility with an increase in temperature.
- Identify** how much NaCl would dissolve in 50 mL of water at 80°C. Give evidence for your response.
- Identify** the substance that has the smallest change in solubility over the temperature increase shown on the graph.
- Determine** which substances are more soluble than potassium nitrate at temperatures below 40°C.
- Deduce** the temperature conditions that would be needed to create a situation where the solubility of lead(II) nitrate is 100 g/100 mL.
- Compare** the solubility of all three substances at 23°C.

CHAPTER SUMMARY

Solubility

- The solubility of an ionic substance depends on many factors, including the strength of ionic bonds and the strength of the ion–dipole interaction.
- Polar molecular substances can dissolve in water while non-polar molecular substances cannot.
- Ionic compounds can generally dissolve in water because of the interaction between charged ions and polar water molecules.
- Four main factors affect the solubility of substances:
 - molecular polarity
 - temperature
 - pressure (for gas solubility)
 - molecular size.



Solubility curves

- Solubility is often measured as the amount of solute that can dissolve in 100 g of water at a given temperature.
- A solubility curve shows the solubility of different substances at different temperatures.
- This can be used to help determine the degree of saturation and concentration of ions of a solution.

MULTIPLE CHOICE

- When sodium chloride dissolves in water, the water molecules will form:
A covalent bonds with the sodium chloride.
B dipole-dipole bonds with the sodium chloride.
C hydrogen bonds with the sodium and chloride ions.
D ion-dipole bonds with the sodium and chloride ions.

Questions 2–6 refer to the following information.

The solubility of sodium nitrate (NaNO_3) at different temperatures is shown in the table.

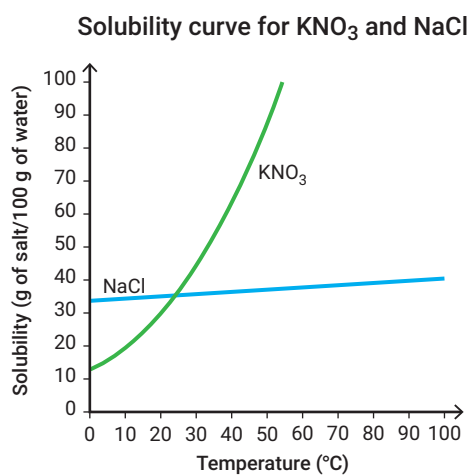
Temperature ($^{\circ}\text{C}$)	Solubility (g/100 g)
10	21
30	48
50	88
70	138

- Determine the mass of sodium nitrate that would dissolve in 50 g water at 30°C .
A 24 g
B 44 g
C 55 g
D 88 g
- Suppose 210 g of sodium nitrate is tipped into 100 g of water at 70°C to make a solution. Some solute remained at the bottom of the beaker. The solution was filtered into another beaker, leaving the undissolved particles behind.
Identify the mass of the undissolved particles.
A 72 g
B 79 g
C 113 g
D 138 g
- The saturated solution at 70°C in Question 3 was cooled to 10°C , at which point further solid precipitated. **Identify** the additional mass that would crystallise out of the solution when it was cooled.
A 27 g
B 60 g
C 117 g
D 137 g
- How many moles of sodium nitrate would be dissolved in a 50.0 mL saturated solution at 50°C ?
A 0.44
B 0.52
C 0.88
D 0.85
- What would be the approximate number of sodium ions dissolved in a 50.0 mL saturated solution of sodium nitrate at 50°C ?
A 3.1×10^{23}
B 3.1×10^{25}
C 6.2×10^{23}
D 6.2×10^{25}

7. When the molecular compound ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) is added to water, it:
 - A dissociates into ions and dissolves.
 - B forms ions by reaction with water molecules.
 - C forms hydrogen bonds between its O–H groups and water molecules.
 - D is immiscible – not able to dissolve in water and forms a separate layer.
8. **Identify** how the solubility in water of gases and most ionic compounds changes as the temperature increases.
 - A Both ionic substances and gases are less soluble.
 - B Both ionic substances and gases are more soluble.
 - C Ionic substances become more soluble and gases less soluble.
 - D Ionic substances become less soluble and gases more soluble.
9. A Chemistry student looks up the solubility curve of KNO_3 and discovers that its solubility is 35 g/100 g water at 25°C . The student dissolves 18 g of KNO_3 in 200 mL water at 25°C . The student produced a:
 - A saturated solution.
 - B unsaturated solution.
 - C supercooled solution.
 - D supersaturated solution.
10. The line on a solubility curve shows:
 - A how fast a substance dissolves.
 - B the minimum amount of solvent that can be dissolved at a given temperature.
 - C the maximum amount of solute that can be dissolved at a given temperature.
 - D the minimum amount of solute that can be dissolved over a temperature range.

SHORT RESPONSE

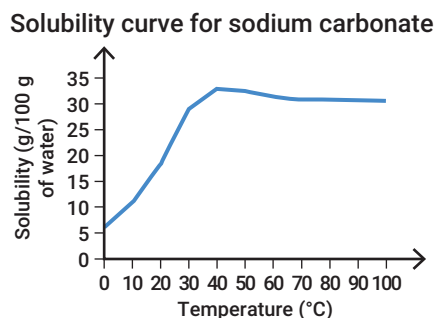
11. The graph shows a solubility curve for potassium nitrate (KNO_3) and sodium chloride (NaCl). **Contrast** the solubilities of KNO_3 and NaCl at 10°C and at 80°C .



12. **Explain** how a Chemistry student could use the information in a solubility curve to make a supersaturated solution.

CROSS-CHAPTER QUESTION

13. The graph shows part of the solubility curve for sodium carbonate (Na_2CO_3).

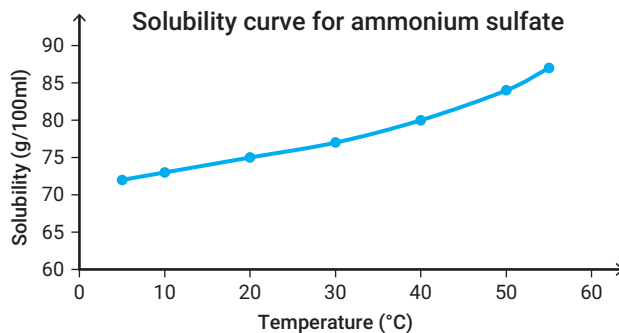


- Identify** how many grams of sodium carbonate are present in a 100 mL sample of a saturated solution of sodium carbonate at 40°C.
- A student cools the 100 mL sample from 40°C to 20°C. They filter and dry the precipitate that is formed. **Determine** the mass of the precipitate formed.
- The student adds HCl to the dried precipitate and it fizzes. **Identify** the gas formed, using a balanced equation to represent this reaction.
- The student captures the gas that was formed in the reaction between HCl and the precipitate. The gas was bubbled through a solution of limewater. **Predict** what the student will observe when the gas is bubbled into the limewater solution.
- Calculate** the number of moles of gas produced, if 15 g of sodium carbonate precipitate was reacted with HCl.

DATA ANALYSIS

14. Analyse data

The graph shows part of the solubility curve for ammonium sulfate ($(\text{NH}_4)_2\text{SO}_4$).



- A student places 205 g of solid ammonium sulfate in a beaker and adds 250 mL of deionised water. **Determine** what minimum temperature the student will have to heat the mixture to produce a saturated solution.
- Calculate** the mass of solid that will crystallise when 75 mL of a saturated solution of ammonium sulfate at 50°C is cooled to 30°C.

- c A 5.5 M solution of ammonium sulfate is heated to a temperature of 40°C. **Classify** this as solution unsaturated, saturated or supersaturated. **Explain** your reasoning.
- d **Predict** how much ammonium sulfate will dissolve in 100 mL water at 60°C.

15. Analyse data

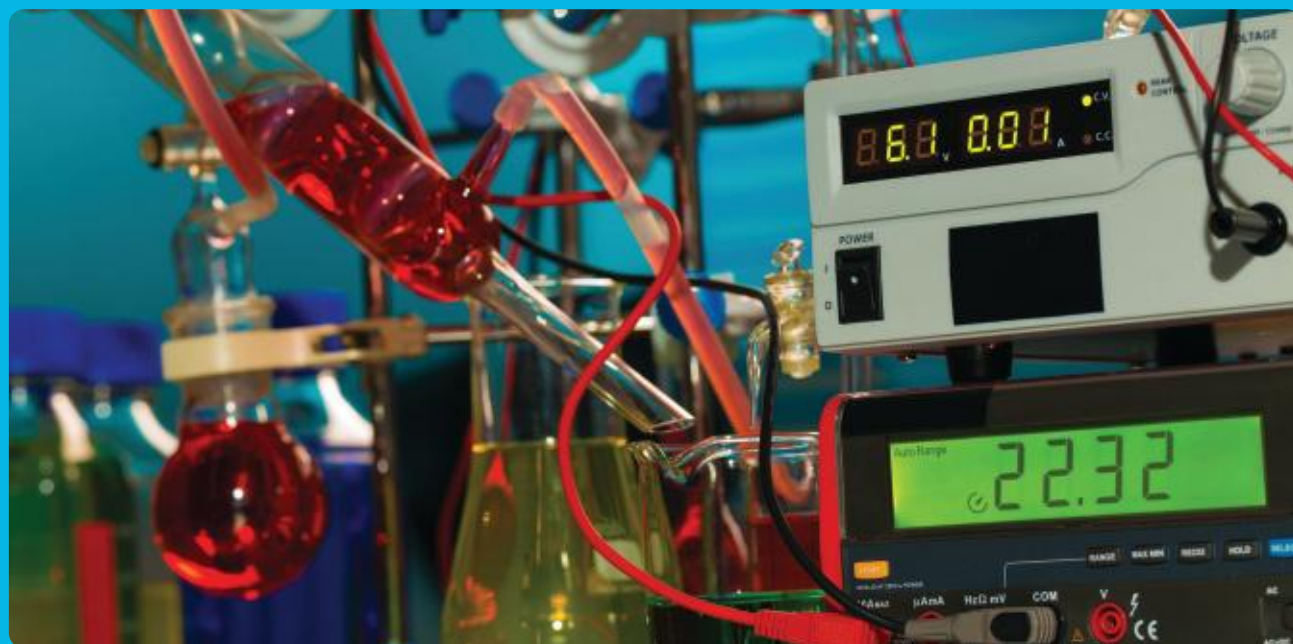
Potassium nitrate is a soluble substance. Its solubility values are listed in the following table.

Solubility (g/100 g water)	Temperature (°C)
8.5	0
31.5	20
105	60

- a **Identify** how much potassium nitrate you would need to make 60 mL of a saturated solution at 60°C.
- b A 90 g sample of potassium nitrate is dissolved in 350 mL at 20°C. **Determine** if the solution is saturated. Give reasons for your answer.
- c **Calculate** the molarity of a 100 g saturated potassium nitrate solution at 0°C.
- d **Infer** how temperature affects the way sodium nitrate dissolves in water.

CHAPTER 17

Reactions of acids and bases



science photo/Shutterstock.com

SYLLABUS DOT POINTS

SCIENCE UNDERSTANDING

- Determine balanced chemical and ionic equation (including states) for the reactions of acids with bases, metals and carbonates.

SCIENCE INQUIRY

Investigate:

- the reactions of acids with bases, metals and carbonates.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Acids and bases play an important part in our lives. Acids and bases affect our health, playing a crucial role in the chemical reactions in our blood. On a larger scale, an understanding of acid and base reactions is important in the regulation of our water supply. Acid–base reactions follow patterns that scientists use to predict the products easily.

Industrialisation on a global scale has produced gases that cause acid rain, which has a negative impact on the environment. Scientists have investigated the chemistry of acid rain and how this problem can be solved.

Practical

- Investigating acid reactions

Worksheets

- Neutralisation reaction
- Reactions of acids



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Balanced chemical equations have the same number of atoms of each element on the reactant and product sides of the equation.
- ✓ Word equations can be written as balanced chemical equations.
- ✓ Acids are substances that ionise in water to release H^+ ions.
- ✓ Ionic substances generally exist as ions in an aqueous solution and therefore are written as anions and cations.
- ✓ Strong acids are acids in which all molecules ionise in solution; weak acids are acids in which only a small percentage of the molecules ionise in solution.
- ✓ Strong bases are bases in which almost all of the species have dissociated; weak bases are bases in which only a small proportion of the species produce hydroxide ions in solution.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ understand that acid–base reactions follow generalised patterns that can be used to write balanced equations
- ✓ understand that neutral species and net ionic equations can be written to show what is happening to the aqueous ions in chemical reactions
- ✓ apply the general pattern for acid–base neutralisation, and acid–metal, acid–carbonate and acid–hydrogen carbonate reactions to write equations.
- ✓ investigate the reactions of acids with bases, metals and carbonates
- ✓ explore the chemistry behind the formation of acid rain and its effects on the natural and built environment.

17.1 Reactions of acids

Acids undergo several different types of reactions. With acid–base reactions, once scientists understand the pattern of the reaction, they can predict most of the products. This makes it easier to write and balance equations. This chapter focuses on acid reactions with bases, metals and carbonates.

Representing acid reactions

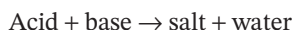
Reaction equations

As outlined in Chapter 15, reactions can be represented as neutral species equations, complete ionic equations and net ionic equations. In the following section, the various equation types are shown for the reactions between a strong acid and a base (Worked example 17.1.1), a strong acid and a weak base (Worked example 17.1.2), a metal and an acid (Worked example 17.2.1), an acid and a carbonate (Worked example 17.2.2), and an acid and a hydrogen carbonate (Worked example 17.2.3).

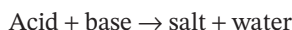
Acid–base reactions

Neutralisation reactions

Neutralisation is a common reaction between acids and bases, following the general pattern:

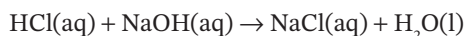


In this reaction, **salts** and water are formed. If the right amounts of acid and base react, the resulting solution is **neutral** (neither acidic nor basic). This generalised neutralisation reaction pattern can be used to predict the products of most reactions between an acid and a base. The formula for the neutralisation reaction is:



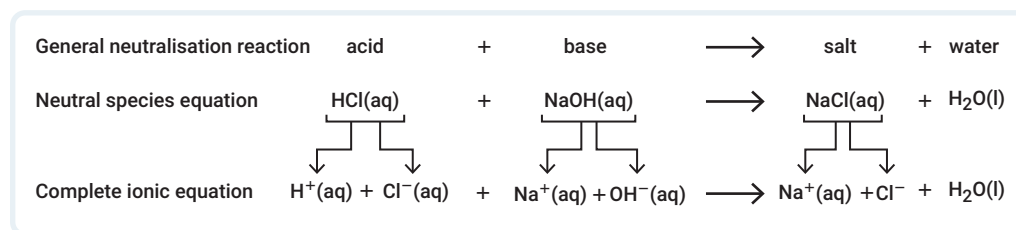
You may recall from Chapter 14, that strong acids become ionised and bases dissociate completely in water. According to the Arrhenius model, acids release a H^+ ion, and bases produce an OH^- ion. If a strong acid reacts with a strong base, then the hydrogen ions and hydroxide ions released **neutralise** each other when water is formed. An acid–base indicator can be used to show when neutralisation has occurred.

A typical reaction between acids and bases is the reaction between the strong acid HCl and the strong base NaOH :

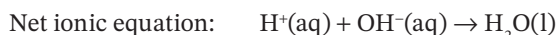


In this reaction, the acid (HCl) reacts with the base (NaOH) to form sodium chloride salt (NaCl) and water. It is important to note that not all ‘salts’ are sodium chloride, but they are formed from the combination of the positive ion from the base and the negative ion from the acid. Water is formed between the OH^- ion from the base and the H^+ ion from the acid.

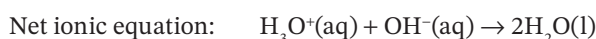
Chemical reactions are more accurately represented by ionic equations. Using the example above, the acid–base reaction is represented in a full ionic equation. Since the acid and base are both strong, they fully ionise/dissociate and are represented as ions. The NaCl product is soluble, so it also dissociates in water.



Sodium and chloride ions are the spectator ions in the reaction, since they produce the salt that remains in aqueous form. Removing these spectator ions leaves the following as the net ionic equation.



Since hydronium ions (H_3O^+) are actually formed when H^+ attaches to a lone pair of electrons on the oxygen atom in the water molecule, the net ionic equation can also be represented as:



neutralisation the reaction between an acid and a base to form a salt and water

salt a substance formed from the positive ion (cation) of a base and the negative ion (anion) of an acid

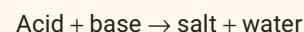
neutral neither acidic nor basic



Syllabus link
Chapter 18 goes into further detail about the Arrhenius model of acids and bases.

KEY FORMULA

Neutralisation reactions



Worksheet
Neutralisation reaction

Weblinks
Acid–base reactions
Neutralisation reactions

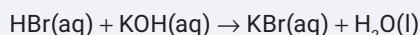
WORKED EXAMPLE 17.1.1

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between hydrogen bromide and potassium hydroxide (a strong acid and a strong base).

ANSWER

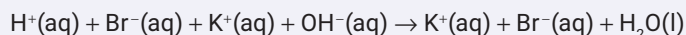
1 Write the neutral species equation.

Acid + base $\rightarrow$ salt + water



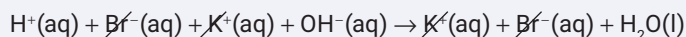
hydrogen bromide + potassium hydroxide $\rightarrow$ potassium bromide + water

2 Write the complete ionic equation.



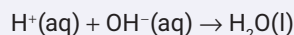
3 Write the net ionic equation.

Delete species that are common to both sides.



The $\text{K}^+(\text{aq})$ and $\text{Br}^-(\text{aq})$ ions in the complete ionic equation are spectator ions.

This means that the reaction between an acid and a base is simply the reaction between the $\text{H}^+(\text{aq})$ ions and $\text{OH}^-(\text{aq})$ ions:



dLibrary.com

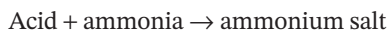


FIGURE 17.1.1 Antacids contain a base, which neutralises excess acid in the stomach.

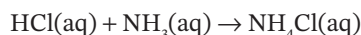
Neutralisation reactions are important in our bodies and in our lives. If too much hydrochloric acid is produced in our stomach, then we can have heartburn or indigestion. This problem is commonly fixed by taking an antacid, which contains a base, such as magnesium hydroxide or aluminium hydroxide (Figure 17.1.1). The added base neutralises the excess stomach acid.

Another common, but not typical, acid-base reaction is the reaction between an acid and ammonia (a weak base). This reaction does not adhere to the generalised neutralisation reaction because the hydrogen ion attaches to the ammonia molecule (NH_3), to produce the ammonium ion ($\text{NH}_4^+(\text{aq})$).

Instead, it has the general formula:

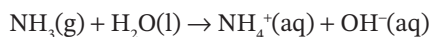


For example:



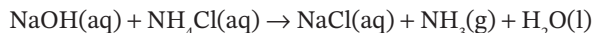
Hydrochloric acid + ammonia $\rightarrow$ ammonium chloride

Ammonia is classified as a base because when it reacts with water, hydroxide ions (OH^-) are produced. Ammonia is considered to be a weak base:



Ammonia + water $\rightarrow$ ammonium ion + hydroxide ion

The ammonium ion (NH_4^+) is considered a weak acid. This means it reacts with a base to produce a salt and water. However, ammonia gas is also produced in the reaction as seen in the following example:



Sodium hydroxide + ammonium chloride $\rightarrow$ sodium chloride + ammonia + water

Ammonium salts such as ammonium sulfate and ammonium nitrate are commonly used as nitrogen fertilisers. Alkaline compounds (soluble bases) such as lime (CaO) are often used to prevent soils from becoming too acidic. One reason a farmer does not treat a field with an ammonium compound at the same time as using an alkaline compound is that the ammonium salt will react with the base to give off ammonia. The ammonia will escape into the atmosphere as a gas, and not be available to fertilise the plants.

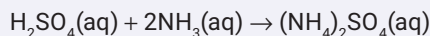
WORKED EXAMPLE 17.1.2

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between sulfuric acid (H_2SO_4) and ammonia (NH_3) (a strong acid and a weak base).

ANSWER

1 Write the neutral species equation.

Acid + ammonia $\rightarrow$ ammonium salt



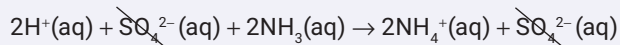
Sulfuric acid + ammonia $\rightarrow$ ammonium sulfate

2 Write the complete ionic equation.

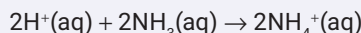


3 Write the net ionic equation.

Delete species that are common to both sides:



Write the net ionic equation.



Divide by 2 to simplify the equation.



LEARNING CHECK 17.1

DESCRIBING

- 1 State the general word equation to represent a neutralisation reaction between an acid and a base.

APPLYING

- 2 Determine neutral species, complete ionic and net ionic equations for the reactions between:

- a hydrochloric acid (HCl) and potassium hydroxide (KOH)
- b sulfuric acid (H_2SO_4) and magnesium hydroxide (Mg(OH)_2)
- c nitric acid (HNO_3) and lithium hydroxide (LiOH)
- d hydrofluoric acid (HF) and barium hydroxide (Ba(OH)_2).

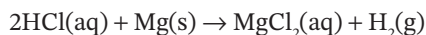
17.2 Other acid reactions

Acid–metal reactions

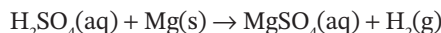
Acids react with some metals. Not all metals are reactive, but for those that are, the reaction between an acid and a metal can be represented by the general equation:



The following reactions are examples of reactions between an acid and magnesium metal:



Hydrochloric acid + magnesium $\rightarrow$ magnesium chloride + hydrogen gas

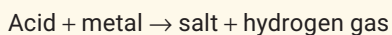


Sulfuric acid + magnesium $\rightarrow$ magnesium sulfate + hydrogen gas

Since hydrogen gas is produced by these reactions, a **pop test** is commonly used to confirm that the gas produced is hydrogen. A small sample of hydrogen gas is collected and exposed to a flame, which initiates a combustion reaction between the hydrogen gas and oxygen in the surrounding air. Water is produced and the combustion reaction is accompanied by the telltale ‘pop’ sound (Figure 17.2.1).

KEY FORMULA

Acid–metal reactions



pop test an audible test to determine the presence of hydrogen gas.

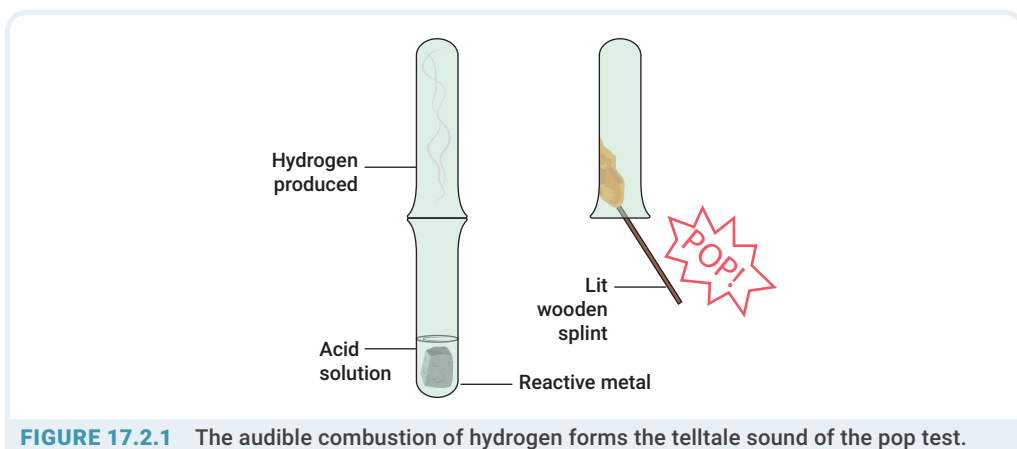


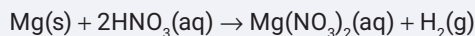
FIGURE 17.2.1 The audible combustion of hydrogen forms the telltale sound of the pop test.

WORKED EXAMPLE 17.2.1

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between magnesium and nitric acid (HNO_3) (a metal and an acid).

ANSWER

1 Write the neutral species equation.



2 Write the complete ionic equation.



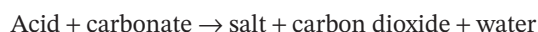
3 Write the net ionic equation.



Acid-carbonate reactions

The acid-carbonate reaction is another type of acid reaction with predictable results. These reactions are very useful in cooking. Food products such as bread and sponge cakes have a honeycomb structure, which contains bubbles. These bubbles are formed by carbon dioxide (CO_2) gas causing the mixture to rise during cooking. Carbon dioxide is formed by adding a carbonate (CO_3^{2-}) or hydrogen carbonate (HCO_3^-) and an acid to the cooking mixture.

When an acid and carbonate react, they produce a salt, carbon dioxide and water. The reaction between an acid and a carbonate can be represented by the general equation:



The following two reactions are examples of reactions between an acid and calcium carbonate:

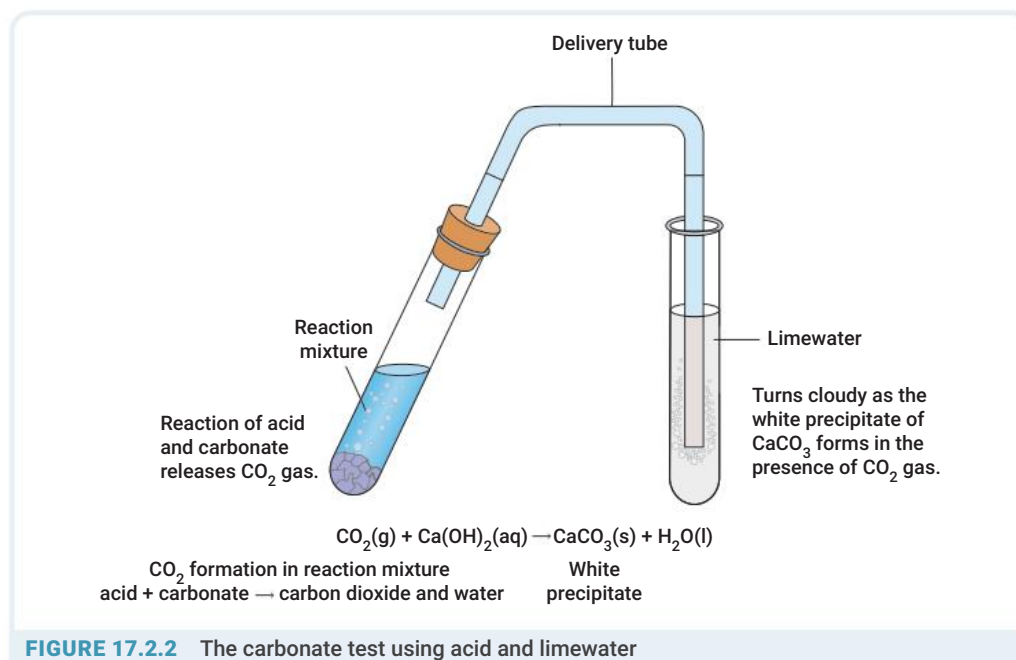


Hydrochloric acid + calcium carbonate $\rightarrow$ calcium chloride + water + carbon dioxide

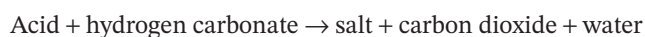


Sulfuric acid + calcium carbonate $\rightarrow$ calcium sulfate + water + carbon dioxide

Since carbon dioxide gas is produced by these reactions, a limewater test is commonly used to confirm that the gas produced is carbon dioxide. Carbon dioxide gas produced by the reaction is bubbled through limewater, an aqueous solution of $\text{Ca}(\text{OH})_2$. If carbon dioxide is present, it forms a white precipitate (CaCO_3), which turns the limewater milky (Figure 17.2.2). (You learnt about this reaction in Chapter 15.)

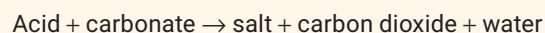


When an acid and a hydrogen carbonate (HCO_3^-) react, they also produce a salt, water and carbon dioxide according to the general equation:



KEY FORMULA

Acid-carbonate reactions



Worksheet
Reactions of acids

Weblink

Reactions between
carbonates and acids

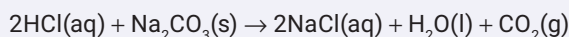
WORKED EXAMPLE 17.2.2

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between hydrochloric acid and sodium carbonate (an acid and a carbonate).

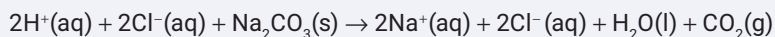
ANSWER

1 Write the neutral species equation.

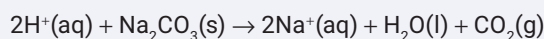
Acid + hydrogen carbonate $\rightarrow$ salt + carbon dioxide + water



2 Write the complete ionic equation.



3 Write the net ionic equation.



KEY FORMULA

Reactions of acid and hydrogen carbonate

Acid + hydrogen carbonate $\rightarrow$ salt + carbon dioxide + water

The following reactions are examples of reactions between an acid and sodium hydrogen carbonate:



Hydrochloric acid + sodium hydrogen carbonate $\rightarrow$ sodium chloride + water + carbon dioxide



Sulfuric acid + sodium hydrogen carbonate $\rightarrow$ sodium sulfate + water + carbon dioxide

In cooking, a relatively safe weak acid, such as acetic acid (the acidic component in vinegar) or citric acid (the acid in citrus fruits) is used, together with sodium bicarbonate (commonly called bicarb soda or baking soda), to release carbon dioxide during baking.



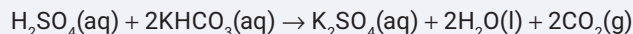
FIGURE 17.2.3 Sodium bicarbonate releases carbon dioxide, which makes the cake rise.

WORKED EXAMPLE 17.2.3

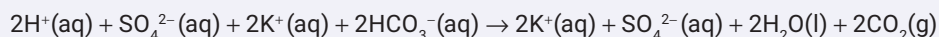
Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between sulfuric acid and sodium hydrogen carbonate (an acid and a hydrogen carbonate).

ANSWER

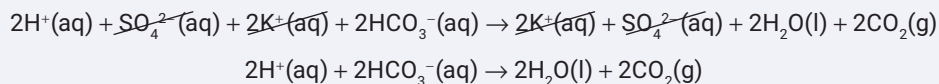
1 Write the neutral species equation.



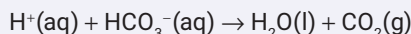
2 Write the complete ionic equation.



3 Write the net ionic equation.



Divide through by 2 to simplify the equation.



PRACTICAL ACTIVITY 17.2.1

INVESTIGATING ACID REACTIONS

Introduction

Acids react with bases, metals, hydrogen carbonates and carbonates in predictable ways. During this experiment, you will investigate a range of acid reactions to examine these generalised reaction patterns.

Research question

How do acids interact with different types of substances?

Aim

To investigate some common reactions of acids

Materials

- 2 cm strip of magnesium
- CaCO_3 (marble chips)
- 0.1 M HCl in a dropper bottle
- 1 M HCl in a dropper bottle
- 0.1 M NaOH in a dropper bottle
- limewater
- universal indicator and colour chart
- 5 test tubes and a test-tube rack
- stopper with glass tube and hose
- wooden splint/taper
- matches
- sandpaper



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
HCl is very corrosive.	Wear gloves, safety glasses and a laboratory coat.
Hydrogen is flammable.	Only use an inverted test tube to collect the hydrogen gas.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the two listed risks to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

Neutralisation reaction

- 1 Place a test tube in the test-tube rack and add 20 drops of 0.1 M HCl.
- 2 Add 1 drop of universal indicator. Observe the colour.
- 3 Add 0.1 M NaOH to the test tube, drop by drop. Observe any colour changes. Record your observations.

Acid-metal reaction

- 1 Clean the 2 cm magnesium strip with sandpaper and place it into a test tube in the test-tube rack.
- 2 Add 5 cm of 1 M hydrochloric acid. Invert a clean test tube over the top to collect the gas.
- 3 Remove the inverted test tube, keeping it inverted and place a lit wooden splint at the mouth of the inverted test tube.
- 4 Listen for the 'pop' sound. If an audible pop is heard, the collected gas is hydrogen. Record your observations.

Acid-carbonate reaction

- 1 Place a small number of marble chips into a test tube in the test-tube rack.
- 2 Place 5 cm of limewater into a second test tube.
- 3 Add 3 cm of 1 M HCl to the test tube containing the marble chips. Quickly place the stopper with the glass tube and hose in the top of this test tube and bubble the gas into the limewater test tube. Record your observations.

Results

Copy and complete the results table.

Reaction	Products formed	Observations
Acid + base		
Acid + metal		
Acid + carbonate		

Analysis of results

- 1 Construct balanced chemical equations for each reaction in this experiment.
- 2 Determine what your observations revealed about the products formed for each of the reactions.
- 3 Compare the observations for the three tests.

Interpretation

- 4 Determine whether these three experimental reactions followed the generalised patterns for acid reactions. Give reasons for your response.
- 5 Discuss whether a different acid, such as sulfuric acid, would yield similar results.

LEARNING CHECK 17.2

DESCRIBING

1 Describe:

- a which ion is produced when an acid ionises
- b a salt
- c a neutralisation reaction
- d the acid-carbonate reaction
- e the acid-metal reaction
- f the pop test.

APPLYING

2 Determine balanced chemical equations (including states) for the following acid reactions.

- a $\text{HNO}_3(\text{aq}) + \text{MgCO}_3(\text{s}) \rightarrow$
- b $\text{H}_3\text{PO}_4(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow$
- c $\text{CH}_3\text{COOH}(\text{aq}) + \text{Mg}(\text{s}) \rightarrow$

3 Determine the product(s) missing from these word equations.

- a Hydrochloric acid + zinc metal $\rightarrow$ zinc chloride + _____
- b Sulfuric acid + calcium carbonate $\rightarrow$ calcium sulfate + _____ + carbon dioxide
- c Hydrochloric acid + sodium hydrogen carbonate $\rightarrow$ _____ + water + _____

4 Determine the ionisation reaction for the following acids in water, by writing the ions these acids will form and the states they will be in.

- a Nitric acid (HNO_3)
- b Hydrochloric acid (HCl)
- c Phosphoric acid (H_3PO_4)

5 Determine the overall chemical equations, complete ionic equations and net ionic equations for the following acid reactions.

- a $\text{HNO}_3(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow$
- b $\text{H}_3\text{PO}_4(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow$
- c $\text{HCl}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow$

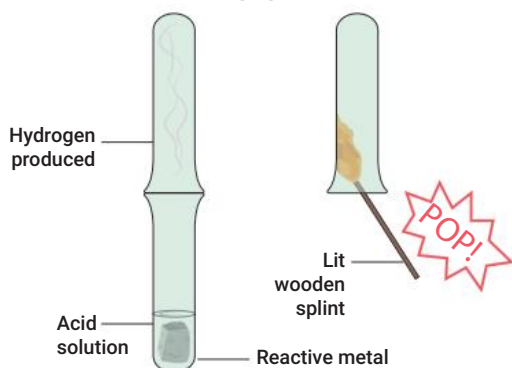
CHAPTER SUMMARY

Representing reactions with acids

- Reactions with acids can be represented in different ways:
 - General reaction equation:
acid + base $\rightarrow$ salt + water
 - Word equation:
hydrochloric acid + sodium hydroxide $\rightarrow$ sodium chloride + water
 - Neutral species equation:
 $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
 - Complete ionic equations:
 $\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O(l)}$
 - Net ionic equation:
 $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)}$

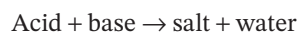
Reactions of acids

The pop test



Neutralisation reaction

The general equation for reactions between an acid and a base is:

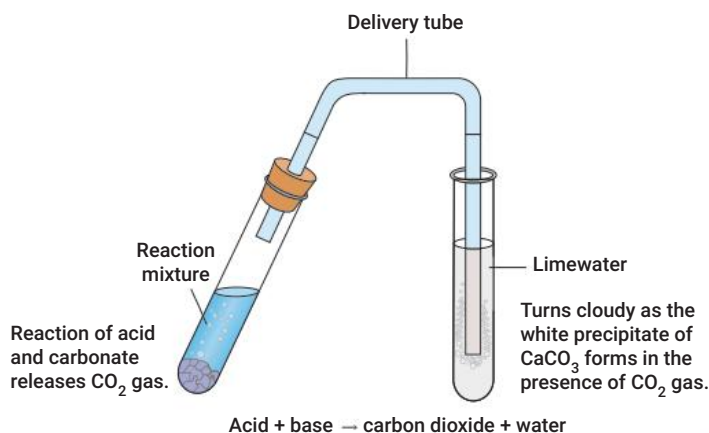


Acid-metal reactions

- The general equation for reactions between acids and metals is:
$$\text{Acid} + \text{metal} \rightarrow \text{salt} + \text{hydrogen gas}$$
- The pop test can be used to confirm acid-metal reactions. The hydrogen produced in this reaction will cause a 'pop' when exposed to a flame.

Acid-carbonate and acid-hydrogen carbonate reactions

The limewater test



- The general equation for reactions between acids and carbonates is:
$$\text{Acid} + \text{carbonate} \rightarrow \text{salt} + \text{carbon dioxide} + \text{water}$$
- The general equation for reactions between an acid and a hydrogen carbonate is:
$$\text{acid} + \text{hydrogen carbonate} \rightarrow \text{salt} + \text{carbon dioxide} + \text{water}$$
- Carbon dioxide produced by acid-carbonate and acid-hydrogen carbonate reactions can be confirmed by the limewater test, which results in the formation of a cloudy precipitate (calcium carbonate).

MULTIPLE CHOICE

- Identify** the products of the acid-carbonate reaction.
A Salt
B Salt + water
C Salt + hydrogen
D Salt + carbon dioxide + water
- Identify** the salt that would be formed in the following reaction:
$$2\text{H}_3\text{PO}_4(\text{aq}) + 3\text{Na}_2\text{CO}_3(\text{aq}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + \underline{\hspace{2cm}}$$

A $\text{H}_2\text{CO}_3(\text{aq})$
B $\text{Na}_2\text{PO}_4(\text{aq})$
C $\text{Na}_3\text{PO}_4(\text{aq})$
D $\text{NaOH}(\text{aq})$
- Identify** the gas produced when sodium hydrogen carbonate reacts with hydrofluoric acid.
A O_2
B H_2
C CO_2
D Cl_2
- Identify** the substance missing from the word equation:
$$\text{acid} + \text{metal} \rightarrow \text{salt} + \underline{\hspace{2cm}}$$

A Ion
B Water
C Oxygen gas
D Hydrogen gas
- Identify** the acid that reacts with potassium carbonate to produce potassium nitrate.
A Nitric acid
B Sulfuric acid
C Hydrochloric acid
D Hydrofluoric acid
- Identify** the salt produced when sulfuric acid reacts with sodium hydroxide.
A Sodium chloride
B Sodium sulfate
C Calcium sulfate
D Sodium hydrogen carbonate
- Identify** the term that describes the reaction between an acid and a base.
A Colouration
B Ionic reaction
C Water formation
D Neutralisation reaction
- Classify** the reaction type shown in the following word equation:
$$\text{Sulfuric acid} + \text{calcium hydroxide} \rightarrow \text{calcium sulfate} + \text{water}$$

A Acid-base reaction
B Acid-metal reaction
C Acid-carbonate reaction
D Acid-hydrogen carbonate reaction

9. **Identify** the name of the salt that forms when calcium carbonate reacts with nitric acid.
- A Calcium bromide
 - B Calcium nitrate
 - C Calcium nitrite
 - D Nitric calcide
10. When sulfuric acid is added to a solid sample of blue-green copper(II) carbonate, the reaction mixture bubbles and the solid dissolves to form a clear, blue solution of copper sulfate. **Identify** the balanced chemical equation that symbolises this reaction.
- A $\text{H}_2\text{SO}_4(\text{aq}) + \text{CuCO}_3(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{CO}_2(\text{g})$
 - B $\text{H}_2\text{SO}_4(\text{aq}) + \text{CuCO}_3(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq})$
 - C $\text{H}_2\text{SO}_4(\text{aq}) + \text{CuCO}_3(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - D $\text{H}_2\text{SO}_4(\text{aq}) + \text{CuCO}_3(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

SHORT RESPONSE

11. With the aid of neutral species equations and complete ionic equations, **determine** the net ionic equation for the reaction between calcium hydrogen carbonate powder ($\text{Ca}(\text{HCO}_3)_2(\text{s})$) and nitric acid ($\text{HNO}_3(\text{aq})$).
12. **Construct** neutral species, complete ionic and net ionic equations for the following reactions.
- a Nitric acid solution and solid magnesium
 - b Sulfuric acid solution and sodium hydrogen carbonate solution
 - c Phosphoric acid solution and magnesium powder
 - d Potassium hydroxide solution and sulfuric acid solution

CROSS-CHAPTER QUESTION

13. A student was mixing acetic acid (vinegar, CH_3COOH) with baking soda (NaHCO_3). They used a stopper and glass tube to direct the resultant gas into some limewater ($\text{Ca}(\text{OH})_2$), and noticed the limewater turned milky.
- a **Construct** a balanced chemical equation for the reaction between vinegar and baking soda.
 - b **Determine** what caused the limewater to turn milky.
 - c **Construct** a balanced chemical equation for the reaction between the gas formed and the limewater ($\text{Ca}(\text{OH})_2(\text{aq})$) to make this precipitate.

DATA ANALYSIS

14. **Analyse data**

A student was recording their observations of a neutralisation reaction between 1 M hydrochloric acid and 1 M sodium hydroxide. They placed 3 mL of hydrochloric acid into a test tube with 1 drop of bromothymol blue indicator solution. They observed the colour change as they added sodium hydroxide to the test tube, 1 mL at a time. Their observations are outlined in the following table.

TABLE 1 Results of experiment

Reaction mixture in test tube	Colour observed
3 mL HCl + 0 mL NaOH	Yellow
3 mL HCl + 1 mL NaOH	Yellow
3 mL HCl + 2 mL NaOH	Greenish yellow
3 mL HCl + 3 mL NaOH	Green
3 mL HCl + 4 mL NaOH	Greenish blue
3 mL HCl + 5 mL NaOH	Blue

The following table shows the colours of some indicators in an acid and in a base.

TABLE 2 Indicator colour ranges

Indicator	Colour in acid	Colour in base
Phenolphthalein	Colourless	Pink
Bromothymol blue	Yellow	Blue
Phenol red	Yellow	Red
Methyl orange	Red	Yellow

- Determine** the colour at the beginning of the reaction, when there was only HCl in the test tube.
- Identify** the trend in the colour change results as the sodium hydroxide was added.
- Identify** the colour of the indicator solution when the HCl was neutralised by the NaOH.
- Determine** the amount of NaOH added to neutralise the HCl solution.
- Infer** the colour of the solution past neutralisation if phenol red indicator was used and NaOH was continually added.

15. Analyse data

A baker mistakenly erased the labels from their sodium chloride and sodium hydrogen carbonate powders. The baker conducted an experiment in which they mixed the powder with vinegar to see if they could tell these substances apart. The results are as follows.

TABLE 3 Results of identification experiment

Substance	Appearance	Reaction with vinegar
Unknown A	Fine white crystalline powder	No reaction
Unknown B	Fine white crystalline powder	Bubbles formed

The gas formed in the experiment was tested by bubbling it into a container of limewater. The limewater turned milky in colour. The gas was also tested by combusting it with a lit splint. The flame went out in the presence of the gas.

- Compare** the results in appearance and reaction with vinegar for the two unknown substances.
- Identify** which of the two substances is the sodium hydrogen carbonate. Explain your response.
- Infer** the name of the gas produced for unknown B.



H_Ko/Shutterstock.com

SYLLABUS
DOT POINTS**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.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Acids and bases are all around us. At home, in industry and in our environment, acids and bases are a part of our lives. These substances vary greatly in their degree of acidity and basicity and it is important to be able to identify this. The Swedish scientist Svante Arrhenius recognised that the acidity of a solution was related to the concentration of hydrogen ions (H^+) in solution but the calculations required to compare solutions were cumbersome. The Danish biochemist Soren Sorensen devised a scale to make writing the hydrogen ion concentration in solution more manageable. This became known as the pH scale. pH is a measure of the degree of acidity or alkalinity of a solution.

Practical

- Investigating the properties of strong and weak acids

Worksheets

- pH of common substances
- Comparing strengths



To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Ionic compounds can dissociate in solution.
- ✓ Some molecular compounds can ionise in solution.
- ✓ Concentration refers to the amount of particles in a solution and is measured in moles per litre.
- ✓ Concentrated solutions have many particles dissolved in a certain volume.
- ✓ Dilute solutions have few particles dissolved in a certain volume.

LEARNING OUTCOMES

By the end of this chapter, you should be able to:

- ✓ describe properties of acids and bases
- ✓ define the Arrhenius model of acids and bases
- ✓ state that pH is dependent on hydrogen ion concentration
- ✓ identify that the pH scale is a logarithmic scale
- ✓ apply the pH scale to explain differences in acidity of solutions
- ✓ determine the acidity of a solution based on the pH and concentration of hydrogen ions
- ✓ apply colour changes of indicators to identify the pH of a solution
- ✓ identify strong and weak acids and bases
- ✓ distinguish between concentration and strength of acids and bases.

18.1 pH and the pH scale

acid a substance that produces hydrogen ions (H^+) in aqueous solution

Shutterstock.com/Feliks



FIGURE 18.1.1 Coffee beans contain caffeine, which is a base. The two most common species of coffee are *Coffea arabica* and *Coffea canephora*, the latter having a higher caffeine content.

base a substance that either contains the oxide ion (O^{2-}) or the hydroxide ion (OH^-), or produces hydroxide ions in an aqueous solution

Most people think of **acids** as being highly corrosive and something that should be avoided, yet many substances in the world are acidic. Rain is slightly acidic, as are many foods and drinks. Soft drink contains carbonic acid, citrus fruits contain citric acid and many foods contain vitamin C, which is ascorbic acid. Acids are also important components of many of the compounds that make up living things. Proteins are long chains of amino acids and the stomach produces hydrochloric acid to help digest food.

Bases are also part of the natural environment, although they are not as

commonly recognised as acids. Ammonia is a base in volcanic gases and is produced by rotting plant and animal matter. Caffeine, in coffee and tea, is a base (**Figure 18.1.1**), as is nicotine, which is in tobacco.

Definitions of acids and bases

Acids and bases are families of substances. Members of each family can be recognised by their common properties.



Weblink

What is the pH scale?

Acids have common properties such as a sour taste and can:

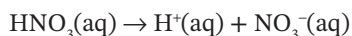
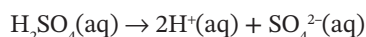
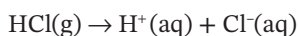
- sting or burn the skin
- conduct electricity in solution
- turn blue litmus red.

Bases have common properties such as a bitter taste and can:

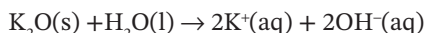
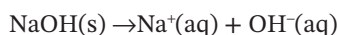
- feel slippery in aqueous solution
- conduct electricity in solution (not all bases are soluble)
- turn red litmus blue.

The effects of acids and bases on vegetable dyes such as litmus were initially used to classify substances as acids and bases. While practical definitions are useful in the laboratory, chemists attempt to define substances by what they are as well as by what they do. The definition of an acid has changed several times over the last 250 years. In the 18th century, French scientist Antoine Lavoisier was one of the first to try to establish the chemical definition of an acid. In the late 19th century, Swedish chemist Svante Arrhenius attempted to explain acids and bases in terms of the particles they produce in aqueous solution. He proposed that acids are substances that ionise in solution to produce hydrogen ions (H^+) and that acids are strong if most of the acid molecules ionise and weak if only a few of the acid molecules ionise. Arrhenius defined a base as a substance that in solution produces hydroxide ions (OH^-). Note that not all bases dissolve in water, so this definition has been modified to include insoluble bases.

Common acids are hydrochloric acid (HCl), sulfuric acid (H_2SO_4) and nitric acid (HNO_3). When these acids are added to water to form acidic solutions, the molecules ionise, producing H^+ ions, as shown in the following equations:



Common soluble bases (alkalis) are sodium hydroxide (NaOH), barium hydroxide (Ba(OH)_2), potassium oxide (K_2O) and ammonia (NH_3). These substances either contain, or in aqueous solution produce, the hydroxide ion (OH^-). For example:



pH

The term 'pH' stands for 'power of hydrogen'. This name came about because the pH scale is based on the concentration of hydrogen ions in solution. Danish biochemist Soren Sorensen devised the scale to make writing the hydrogen ion concentration in solution more manageable.

pH is a measure of how acidic or alkaline (another way of saying 'basic') a solution is.

Notice the relationship between the pH value and the superscript number in the hydrogen ion concentration (**Table 18.1.1**). They are the same number because pH is based on the hydrogen ion concentration of a solution.

pH a scale used to measure the acidity of a solution: pH = 1 is strongly acidic; pH = 7 is neutral; pH = 14 is strongly alkaline

TABLE 18.1.1 The relationship between pH and hydrogen ion concentration

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
H^+ concentration (mol L^{-1})	1	10^{-1} or 0.1	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}



Weblink
The pH scale

concentrated solution a large amount of solute dissolved in a given volume of solvent

dilute solution a small amount of solute dissolved in a given volume of solvent



Syllabus link
Concentration was introduced in Chapter 14.

The mathematical relationship between pH and hydrogen ion concentration is readily apparent:

$$\text{pH} = \log_{10}[\text{H}^+]$$

that is, if $[\text{H}^+] = 10^{-x}$

then $\text{pH} = x$

Recall from Chapter 14, concentration is a measure of the amount of solute dissolved in a given amount of solvent. A **concentrated solution** has a large amount of solute dissolved in a given volume of solvent. Conversely, a **dilute solution** only has a small amount of solute dissolved in a given volume of solvent. Therefore, acids have a high concentration of hydrogen ions and a low pH compared to bases, which have a low concentration of hydrogen ions and a high pH. The pH scale defines a substance as acid, base or neutral (neither an acid nor a base). Acids have a pH of less than 7. Bases have a pH greater than 7. A substance that has a pH equal to 7 is neutral.

As shown in Table 18.1.1, the pH scale ranges from 0 to 14. The lower the pH, the more acidic a solution is; the higher the pH, the more basic the solution is. The pH scale is logarithmic and, as a result, each whole pH value represents a change in concentration by a factor of 10 compared to an adjacent value. For acids ($\text{pH} < 7$), a solution is 10 times more acidic than the next higher value. For example, a solution with a pH 4 is 10 times more acidic than a solution with a pH 5 and 100 times (10×10) more acidic than one with pH 6. For bases ($\text{pH} > 7$), each is 10 times more basic than the next lower whole value. For example, pH 10 is 10 times more basic than pH 9 and 100 times (10×10) more basic than pH 8. This means that a change in 1 unit on the pH scale is equivalent to a 10 times change in the hydrogen ion concentration. For example, a change from pH 1 to 2 is equivalent to a change in the hydrogen ion concentration from 0.1 to 0.01 mol L⁻¹.

Although the pH scale is based on the hydrogen ion concentration, it can also be used to determine alkalinity and the hydroxide ion (OH⁻) concentration. A pH of 7 represents a neutral solution. If equal concentrations of H⁺ and OH⁻ ions were to react, the resultant solution should be neutral. At pH 7, the concentration of H⁺ is 10⁻⁷ mol L⁻¹ so the concentration of OH⁻ must also be 10⁻⁷ mol L⁻¹.

TABLE 18.1.2 The relationship between pH and hydroxide ion concentration

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
H ⁺ concentration (mol L ⁻¹)	1 or 10 ⁰	10 ⁻¹ or 0.1	10 ⁻²	10 ⁻³	10 ⁻⁴	10 ⁻⁵	10 ⁻⁶	10 ⁻⁷	10 ⁻⁸	10 ⁻⁹	10 ⁻¹⁰	10 ⁻¹¹	10 ⁻¹²	10 ⁻¹³	10 ⁻¹⁴
OH ⁻ concentration (mol L ⁻¹)	10 ⁻¹⁴	10 ⁻¹³	10 ⁻¹²	10 ⁻¹¹	10 ⁻¹⁰	10 ⁻⁹	10 ⁻⁸	10 ⁻⁷	10 ⁻⁶	10 ⁻⁵	10 ⁻⁴	10 ⁻³	10 ⁻²	10 ⁻¹	1 or 10 ⁰

As the H⁺ ion concentration decreases, the OH⁻ ion concentration increases. Also note that the concentration superscripts for H⁺ and OH⁻ at a given pH always add up to 14.

Measuring pH



Weblink
The colours and chemistry of pH indicators

Worksheet
pH of common substances

Many natural and synthetic indicators are used to identify whether a solution is acidic or alkaline, as well as its degree of acidity. Synthetic indicators include methyl violet, methyl orange, litmus, bromothymol blue and phenolphthalein. Their colour changes are shown in **Figure 18.1.2**. The actual range of acidity or alkalinity over which an indicator changes colour varies from one indicator to another. For example, a slightly acidic solution is purple in methyl violet, yellow in methyl orange, red in litmus, yellow in bromothymol blue and colourless in phenolphthalein. Some colour changes of common indicators are included in the *Formula and data book*. Chemists have developed a universal indicator that produces several colour changes depending on the pH of the solution. **Table 18.1.3** shows the pH of some common substances with the associated colour change of a universal indicator.

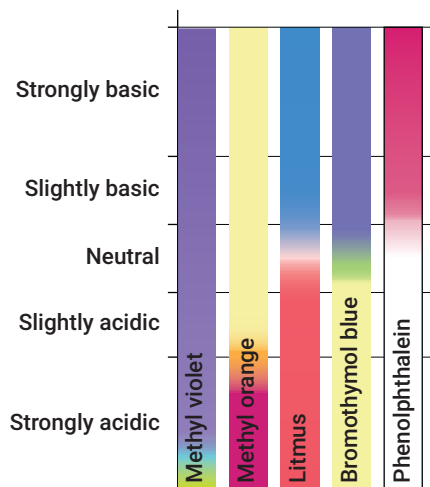


FIGURE 18.1.2 Colour changes in a range of indicators

TABLE 18.1.3 pH of some common substances

Colour of universal indicator	$[H_3O^+]$	pH	Substance	Classification
	1	0	Car battery acid 1 mol L ⁻¹ hydrochloric acid	ACID
	10 ⁻¹	1	0.1 mol L ⁻¹ hydrochloric acid	
	10 ⁻²	2	Stomach acid	
	10 ⁻³	3	Vinegar Lemon juice	
	10 ⁻⁴	4	Soft drinks Soda water	
	10 ⁻⁵	5	Wine Black coffee	
	10 ⁻⁶	6	Rain water Milk, saliva	NEUTRAL
	10 ⁻⁷	7	Pure water	
	10 ⁻⁸	8	Blood Sea water	ALKALINE
	10 ⁻⁹	9	Bore water Baking soda solution	
	10 ⁻¹⁰	10	Toilet soap	
	10 ⁻¹¹	11	Laundry detergents	
	10 ⁻¹²	12	Household ammonia Dishwashing machine powders	
	10 ⁻¹³	13	Chlorine bleach solutions	
	10 ⁻¹⁴	14	Oven cleaners 1.0 mol L ⁻¹ sodium hydroxide	

LEARNING CHECK 18.1

DESCRIBING

- 1 **Define:**
 - a acidic
 - b basic
 - c pH
 - d pH scale
 - e indicator.
- 2 State the pH values for acidic substances, basic substances and neutral substances.
- 3 **Describe** the relationship between the concentration of an acid, the concentration of a base and pH.

APPLYING

- 4 Consider the pH ranges 0–3, 4–6, 8–10 and 11–14. **Determine** the pH range that would apply to each of the following solutions.
 - a Concentrated hydrochloric acid
 - b Distilled water
 - c Dilute ammonia solution
 - d Oven cleaner
- 5 Use Figure 18.1.2 to **identify** which of the following statements are correct.
 - A A substance that turns phenolphthalein pink is more acidic than one that turns methyl orange pink.
 - B A substance that turns methyl orange yellow is less acidic than one that turns methyl orange pink.
 - C A substance that turns phenolphthalein colourless is neutral.
 - D A substance that turns litmus red is more acidic than one that turns bromothymol blue a blue colour.

18.2 The Arrhenius model

The first modern definition of acids and bases was proposed by Svante Arrhenius, a Swedish scientist, in the late 19th century.

Arrhenius suggested that *an acid is a substance that ionises in water to produce H^+ ions*. In other words, an acid increases the concentration of hydrogen ions in solution.

Arrhenius suggested that *a base is a substance that dissociates in water to produce OH^- ions*. In other words, a base is a substance that increases the concentration of hydroxide ions in solution.

However, not all compounds containing hydrogen are acids and not all hydrogen atoms in an acid are released as H^+ ions.

It is only the hydrogen atoms in very polar bonds that are released as H^+ ions. For example, the $H-Cl$ bond is very polar due to the difference in electronegativity values of the hydrogen and chlorine atoms. This means the $H-Cl$ bond is easily ionisable and explains why hydrochloric acid is considered a strong acid. All hydrochloric acid molecules ionise into H^+ and Cl^- ions (Figure 18.2.1).

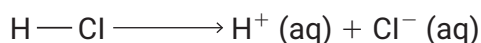


FIGURE 18.2.1 When hydrochloric acid ionises, the polar $H-Cl$ bond breaks, releasing the H^+ ion.



Syllabus link
Chapter 11 discussed the polarity of different compounds



Weblink
Arrhenius acids and bases

A weak acid such as ethanoic (acetic) acid has only one ionisable hydrogen, even though there are four hydrogen atoms in the molecule (**Figure 18.2.2**). Only a small number (<1%) of the ethanoic acid molecules ionise into ions. For equal concentrations of hydrochloric acid and ethanoic acid solutions, the hydrochloric acid solution contains far more hydrogen ions than ethanoic acid does.

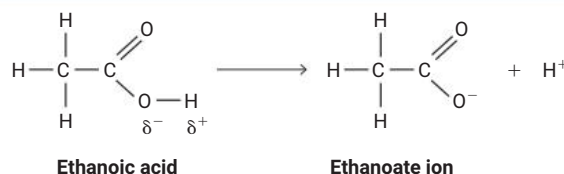


FIGURE 18.2.2 When ethanoic acid ionises, only the polar O–H bond breaks, releasing the H^+ ion. The three H–C bonds are polar and these hydrogen atoms are not released.

With experience, you will be able to recognise these ionisable hydrogens even in the most complex of molecules. In a weak acid solution, there is ionisation of only a few of the acid molecules into ions. That means that in the solution, there will be molecules of the acid, H^+ ions and the corresponding anion. For example, a solution of ethanoic acid contains CH_3COOH molecules, H^+ ions and ethanoate ions (CH_3COO^-) (**Figure 18.2.3**).

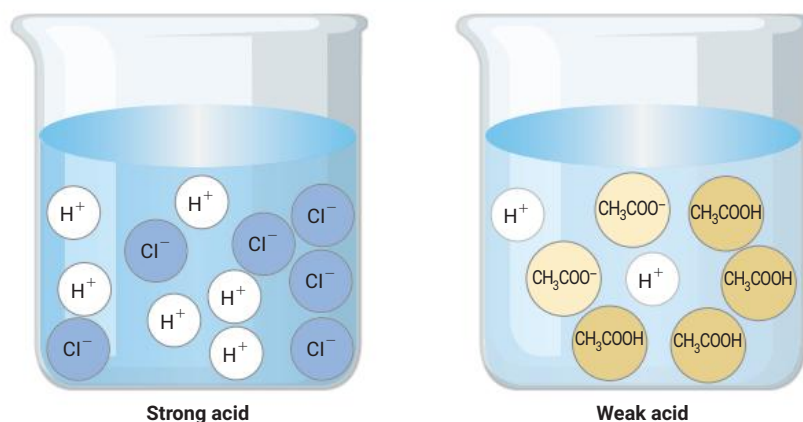
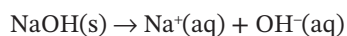
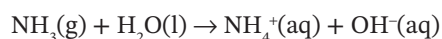


FIGURE 18.2.3 Strong and weak acids. In hydrochloric acid solution, all molecules ionise into H^+ and Cl^- . In ethanoic acid solution, only some of the ethanoic acid molecules ionise into H^+ and CH_3COO^- .

According to the Arrhenius theory, a strong base is one in which all the base particles dissociate in water. In other words, the strength of a base is determined by the proportion of molecules that dissociate to produce OH^- in water. A strong base is one for which a large proportion of the molecules dissociate to produce OH^- ions. Sodium hydroxide (NaOH) and potassium oxide (K_2O) are strong bases – all their molecules dissociate to produce OH^- ions:



The Arrhenius definition of a weak base is one in which only a few of the base particles dissociate in water to produce hydroxide ions. In other words, in weak bases only a small fraction of the molecules or ions react to produce OH^- ions. An example of a weak base is ammonia (NH_3):



In this case, even though NH_3 does not contain OH^- ions, the NH_3 molecule reacts with water to produce OH^- ions, thereby conforming to the Arrhenius definition of a base.

LEARNING CHECK 18.2

DESCRIBING

- 1 Define 'acid' and 'base' according to the Arrhenius model.

APPLYING

- 2 Consider the following statement:
'All substances that contain a hydrogen can act as acids.'
Comment on the accuracy of this statement.
- 3 For each of the following examples, **identify** whether the substance is acting as an acid or a base.
- a $\text{KOH(aq)} \rightarrow \text{K}^{\text{+}}(\text{aq}) + \text{OH}^{-}(\text{aq})$
 - b $\text{HClO}_4(\text{aq}) \rightarrow \text{H}^{\text{+}}(\text{aq}) + \text{ClO}_4^{-}(\text{aq})$
 - c $\text{Ba(OH)}_2(\text{aq}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$

18.3 Recognising strong and weak acids and bases

Strong and weak acids

As you know by now, the strength of acids depends on the proportions of molecules which ionise. Some examples of commonly used strong acids include hydrochloric acid (HCl), nitric acid (HNO₃) and sulfuric acid (H₂SO₄).

A weak acid is one in which only a few molecules ionise. Some weak acids may have many hydrogen atoms but only one, or possibly two, ionisable hydrogen atoms and very few of these will actually ionise. For example, potassium hydrogen phthalate (C₈H₅KO₄), or KHP for short, is an important acid used in quantitative analysis. Its structure may appear daunting at first but, on closer inspection, the ionisable hydrogen atom is easy to spot. It is the one attached to the oxygen atom (**Figure 18.3.1**). Hydrocyanic acid (HCN), carbonic acid (H₂CO₃) and hydrosulfuric acid (H₂S) are three other examples of weak acids. HCN ionises to produce H⁺ and CN⁻ ions while H₂S ionises to produce H⁺ and HS⁻ ions. The pH of strong acids is relatively low compared to that of weak acids, which have a higher pH (closer to 7).

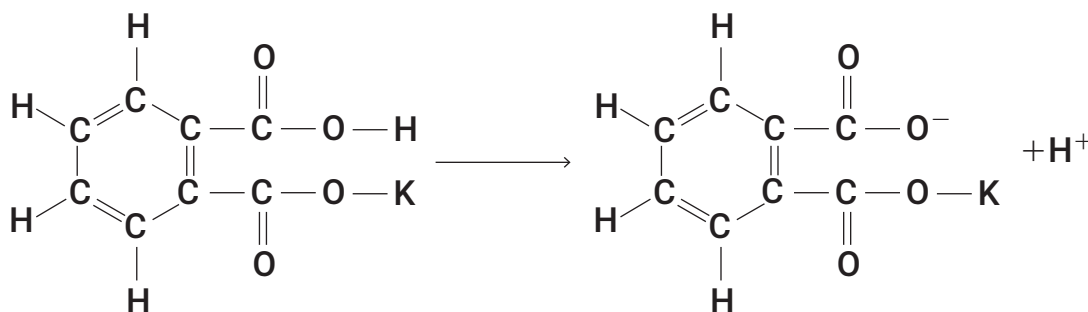
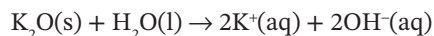


FIGURE 18.3.1 The ionisable hydrogen atom in potassium hydrogen phthalate (KHP) is the one attached to the oxygen atom. This O—H bond is very polar, making it relatively easy to break, thereby releasing the hydrogen ion. As it is a weak acid, very few of these bonds actually break.

Strong and weak bases

Strong bases include the hydroxides and some oxides of groups 1 and 2. For example, potassium oxide (K_2O) reacts in water to produce potassium (K^+) and hydroxide (OH^-) ions:



Ammonia (NH_3) is a weak base and many weak bases are ‘variations’ of ammonia, where one or more hydrogen atoms has been replaced; for example, with a carbon chain. Ethanamine (ethylamine, C_2H_7N) is a weak base that is used extensively in the chemical industry. It is used in the manufacture of compounds, herbicides and pharmaceutical products. (Figure 18.2.2). In this example, ethanamine is a weak base as only some of the base molecules ionise to produce hydroxide ions.



Worksheet
Comparing strengths

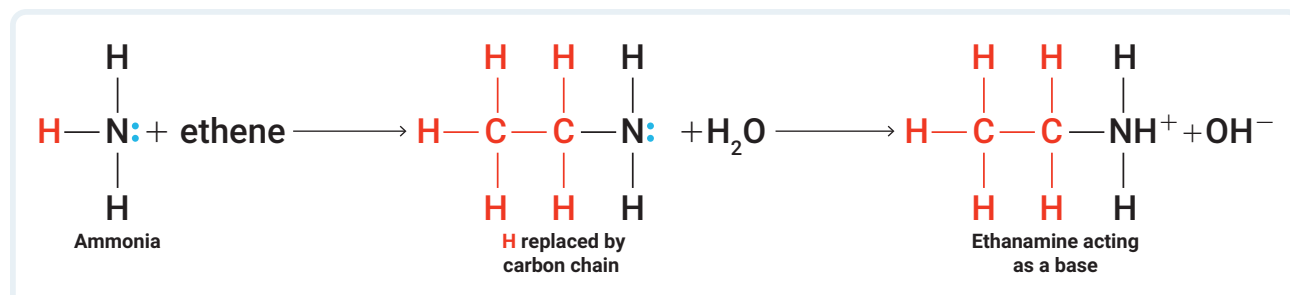


FIGURE 18.3.2 Ethanamine (ethylamine) is a weak base and can be thought of as an ammonia molecule that has had one of its hydrogen atoms (in red) replaced by a carbon chain ($CH_3CH_2^-$).

Strength and conductivity

Because all acids and bases do not dissociate/ionise to the same extent, conductivity can be used as a test for strength. Conductivity refers to the ability to conduct electricity in solution. To conduct electricity, a solution needs to have free-flowing ions to carry the electric current. If the acid or base conducts electricity strongly, this means that many of the acid or base molecules have produced ions, implying that it is a strong acid or base. If the acid or base conducts electricity weakly, few ions have been produced and therefore it is a weak acid or base. A simple conductivity test can be used to determine if a solution conducts, or a more advanced conductivity meter will provide a quantitative measure of conductance. The units of conductivity are siemens per metre ($S m^{-1}$).

TABLE 18.3.1 The conductivities of a strong and a weak acid and base

Substance (1% solution at 25°C)	Chemical formula	Conductivity ($S m^{-1}$)
Ethanoic (acetic) acid	CH_3COOH	0.064
Hydrochloric acid	HCl	10.3
Sodium hydroxide	$NaOH$	5.23
Ammonia	NH_3	0.049

The difference between strength and concentration

As you learnt in Chapter 14, the strength of an acid or a base does not refer to its concentration. The terms ‘concentrated’ and ‘dilute’ are used to describe the concentration of an acid or a base solution, whereas the terms ‘strong’ and ‘weak’ refer to the percentage of acid or base molecules that have ionised or dissociated in solution to produce either H^+ or OH^- ions.



Syllabus link
Chapter 14 discussed the differences between concentration and strength.

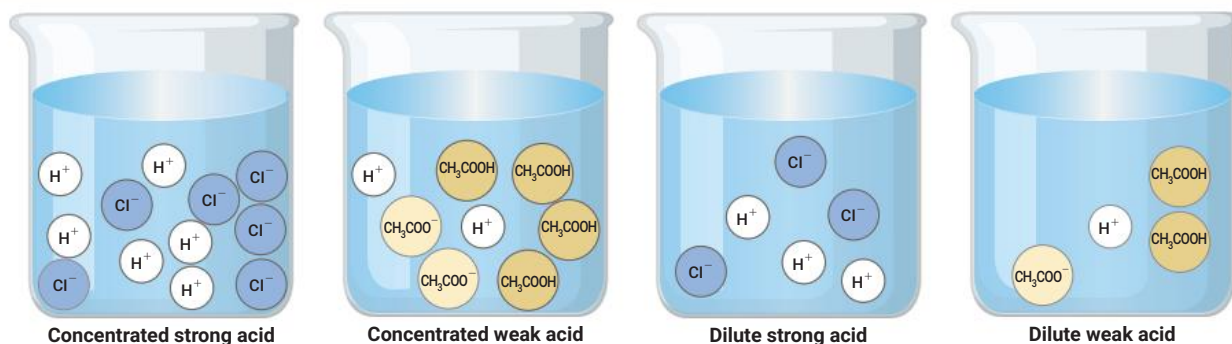


FIGURE 18.3.3 A concentrated strong acid has many acid molecules and they all ionise. A concentrated weak acid has many acid molecules but only some of them ionise. A dilute strong acid has relatively fewer acid molecules and they all ionise. A dilute weak acid has relatively fewer acid molecules and only some of them ionise.

PRACTICAL ACTIVITY 18.3.1

INVESTIGATING THE PROPERTIES OF STRONG AND WEAK ACIDS

Introduction

In this activity you will examine the strength of acids in three different ways: indicators, pH meter and conductance.

Research question

How can the strengths of different acids be distinguished?

Aim

To determine the relative strengths of some strong and weak acids

Equipment per group

- 70 mL of 0.1 mol L⁻¹ solutions of hydrochloric acid (HCl), nitric acid (HNO₃), sulfuric acid (H₂SO₄), ethanoic (acetic) acid (CH₃COOH), citric acid (C₆H₈O₇) and ascorbic acid (vitamin C, C₆H₈O₆)
- universal indicator solution and pH colour chart
- distilled water
- 100 mL beaker
- 6 test tubes and a test-tube rack
- battery or power pack
- electrical leads
- ammeter or multimeter
- carbon rod electrodes
- pH meter



What are the risks in doing this experiment?

Dilute solutions of hydrochloric acid, sulfuric acid and nitric acid can cause irritation to skin and eyes.

How can you manage risks to stay safe?

Wear safety glasses and protective clothing. Do not allow contact with skin or clothing. If contact with skin occurs, wash with water.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Ask your teacher to check your table before you proceed.

Procedure

Part A: Testing the substances with the indicator

- Set up the test tubes in the rack and label each one with the name of an acid being investigated. Place a few millilitres of each acid into its corresponding test tube.
- Add a drop of universal indicator to each acid and record the colour change.

- 3 Compare the colour of each test tube with the pH colour chart and try to assign a pH value for each acid.
- 4 Record the pH value.

Part B: Testing the substances with a pH meter

- 1 Empty the test tubes from the previous experiment and rinse them with distilled water.
- 2 Place a few millilitres of each acid into its corresponding test tube.
- 3 Place the tip of the pH meter into each acid (remembering to rinse the meter between each test using distilled water) and record the reading.

Part C: Conductance test of each acid

- 1 Set up the apparatus as shown in **Figure 18.3.4**.
- 2 Half-fill the beaker with the acid to be tested.
- 3 Switch on the power pack and set to 4 V, or connect the battery.
- 4 Record the reading for the electrical current.

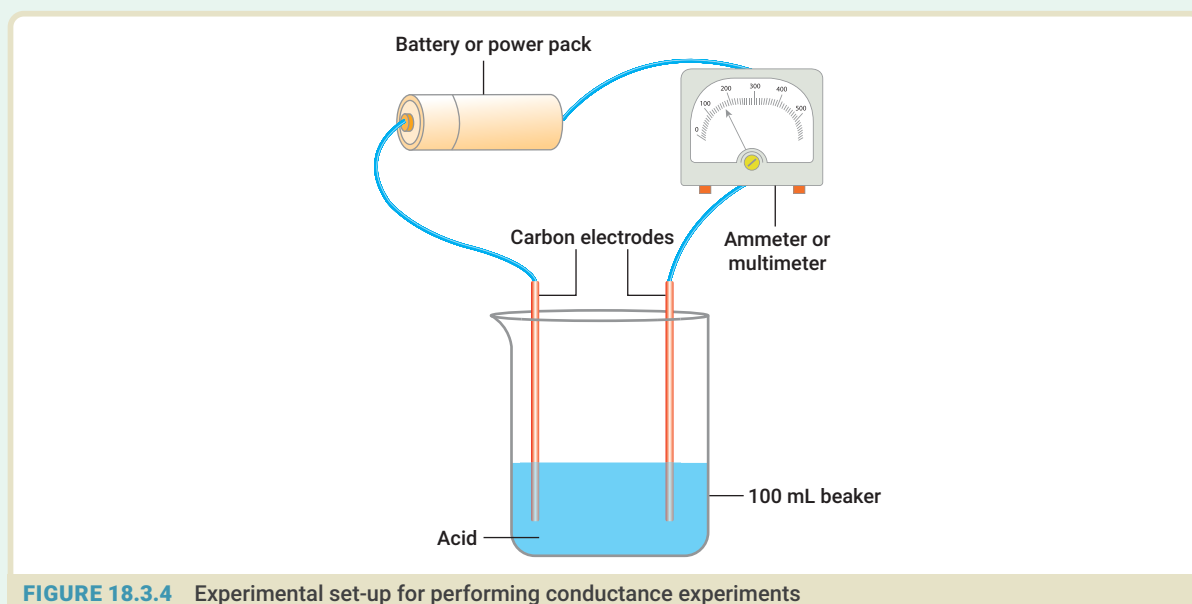


FIGURE 18.3.4 Experimental set-up for performing conductance experiments

Results

Draw up a data table to record the pH colour, pH value from the pH colour chart, pH reading from the pH meter and the current value for each acid.

Interpretation

- 1 Identify any significant differences between the pH values obtained from the universal indicator and the pH meter. Suggest possible reasons for these differences.
- 2 Is there a trend in the current readings for the various acids? If so, what is it?
- 3 Explain why the conductance experiment was carried out by using your understanding of the Arrhenius theory of acids.
- 4 Explain your results of the conductance experiment using the Arrhenius theory.

LEARNING CHECK 18.3

DESCRIBING

- 1 Define:
 - a strong acid
 - b weak acid
 - c strong base
 - d weak base.
- 2 List two examples of strong acids and bases.
- 3 List two examples of weak acids and bases.

APPLYING

- 4 Phosphoric acid is mostly used in fertilisers. Phosphorus in phosphoric acid that is released into the soil is known to support plant growth. Phosphoric acid used in fertilisers is considered a weak acid; however, it is available in concentrated solutions. **Explain** how phosphoric acid solution can be considered weak and concentrated.

ANALYSING

- 5 The table below gives the electrical conductivity values (a measure of the number of ions in solution) for a number of aqueous solutions of the same concentration.

Aqueous solution	Electrical conductivity (S m^{-1})
HNO_3	108
NaOH	93.1
Mg(OH)_2	0.001
NaCl	30.2
NH_3	1.0
CH_3COOH	0.8

- a How does electrical conductivity relate to acid and base strength? **Justify** your explanation by referring to the table.
- b **Explain** why Mg(OH)_2 has a very low electrical conductivity.

Comparing acids and bases

	Acids	Bases
Behaviour	Produce hydrogen ions in solution	Produce hydroxide ions in solution
Properties	Sour taste Corrosive $\text{pH} < 7$	Bitter taste Corrosive Feel slippery in aqueous solutions $\text{pH} > 7$
Strength	All molecules of strong acids ionise. Examples: hydrochloric acid, sulfuric acid, nitric acid Weak acids have few molecules which ionise. Examples: acetic acid, carbonic acid	All particles of a strong base dissociate. Examples: sodium hydroxide, barium hydroxide, potassium oxide Weak bases have few particles which dissociate. Examples: ammonia, ethanamine

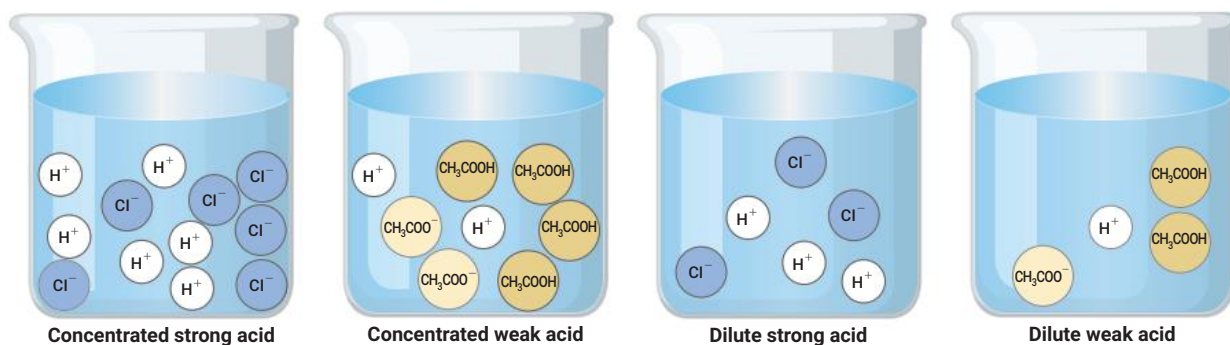
pH

- pH is based on the hydrogen ion concentration in solution.
- The pH scale ranges from 0 to 14 and is logarithmic; a change of 1 unit on the scale is equivalent to a 10 times change in hydrogen ion concentration.
- Acids have a $\text{pH} < 7$, bases have a $\text{pH} > 7$, and neutral substances have a pH of 7.
- Indicators such as universal indicator and litmus are used to indicate whether a solution is acidic or alkaline, as well as its degree of acidity.

Strength vs concentration

- Strength refers to the percentage of the acid or base molecules that ionise in water, whereas concentration refers to the quantity of solute in a given volume of solution.
- Since acids and bases of different strengths ionise differently, we can identify the strength of acids/bases by comparing their conductivities.

Concentration vs strength



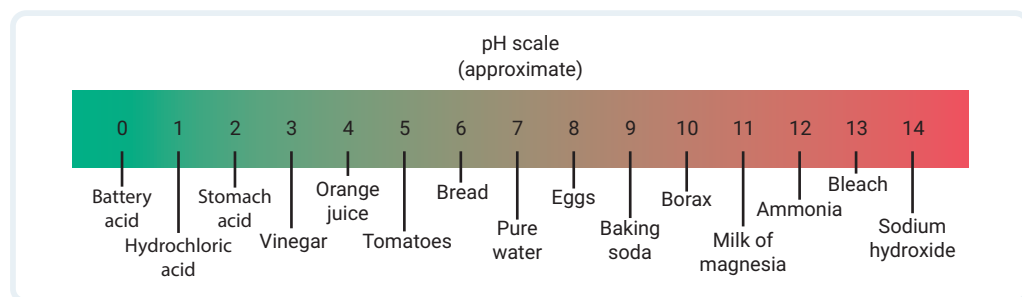
CHAPTER EXAM

MULTIPLE CHOICE

- Acids taste:
 - bitter.
 - salty.
 - sour.
 - sweet.
- Bases make litmus paper turn:
 - blue.
 - green.
 - red.
 - yellow.
- According to the Arrhenius theory of acid and bases, a base is a substance that:
 - tastes bitter.
 - has a pH above 7.
 - can donate a pair of electrons.
 - ionises in water to produce hydroxide ions.
- Which type of solution is one with a pH of 10?
 - Acidic
 - Basic
 - Neutral
 - Salty

Questions 5 and 6 refer to the following information.

The diagram shows the pH values of some substances.



- Which of the following statements is true?
 - Vinegar is 10 times more acidic than normal stomach acid.
 - Tomatoes are two times more acidic than pure water.
 - Baking soda is considered a very strong base.
 - Bread is 100 times more acidic than eggs.
- Identify** which indicator in the following table could be used to distinguish between orange juice (pH 4) and vinegar (pH 3).

	Indicator	pH range and colour change
A	Methyl orange	3.2–4.4 red → yellow
B	Bromocresol green	3.8–5.4 yellow → blue
C	Thymol blue	1.2–2.8 red → yellow
D	Cresol red	7.0–8.8 yellow → red

7. **Identify** which diagram represents the weakest acid.

A

B

C

D

● Acid ● Anion • H⁺

8. The following table shows data from an investigation into the properties of some acids and bases.

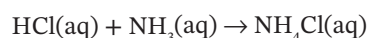
Solution	Electrical conductivity	Original colour of litmus paper	Colour of litmus paper after dipping in solution	pH
1	Very high	Red	Blue	10.0
2	Low	Blue	Red	6.5
3	Moderate	Red	Red	5.4
4	Very high	Blue	Red	2.0

Based on the data, which solution is a strong base?

- A** Solution 1
B Solution 2
C Solution 3
D Solution 4
9. Two bases are compared.
 Base I is 0.1 mol L⁻¹ sodium hydroxide.
 Base II is 1.0 mol L⁻¹ ammonia.
 How does base I compare with base II?
- A** Base I is stronger and more dilute than base II.
B Base I is weaker and more concentrated than base II.
C Base II is stronger and more concentrated than base I.
D Base I is stronger and more concentrated than base II.
10. Which of the following solutions would have the lowest pH?
- A** 1.0 mol L⁻¹ acetic acid
B 0.1 mol L⁻¹ acetic acid
C 1.0 mol L⁻¹ hydrochloric acid
D 0.1 mol L⁻¹ hydrochloric acid

SHORT RESPONSE

11. In a weak acid such as ethanoic (acetic) acid (CH₃COOH), **identify** the kind of bond present that enables it to release hydrogen ions.
12. A major problem with Arrhenius's theory is that, if equal amounts of acid and base react, the resulting solution should be neutral with pH 7. Consider the reaction between equal amounts of hydrochloric acid and ammonia:



The resulting solution from this reaction contains ammonium (NH₄⁺) and chloride ions (Cl⁻). Ammonium is known to be mildly acidic.

The pH of the resulting solution is about pH 4. How would you **explain** this?

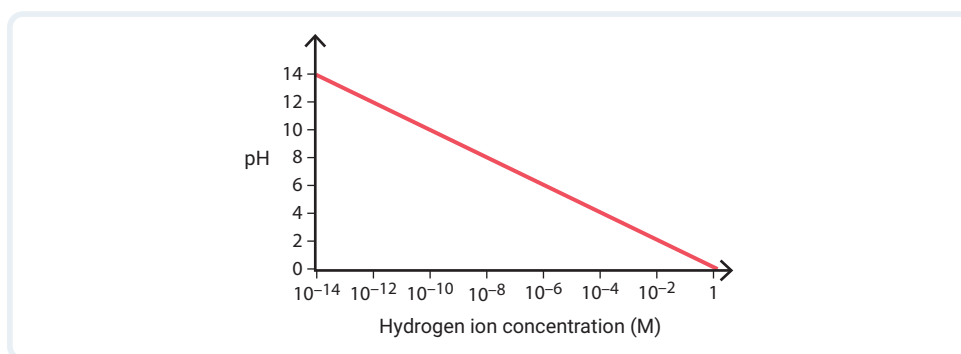
CROSS-CHAPTER QUESTION

13. A solution was prepared by dissolving 2.55 g of sodium hydroxide in water. The solution was made up to a final volume of 250 mL.
- Calculate** the number of moles of sodium hydroxide used.
 - Calculate** the concentration of the sodium hydroxide solution.
 - What would the conductivity and relative pH (no pH value required) be of the resulting solution of sodium hydroxide?

DATA ANALYSIS

14. **Analyse data**

The effect of concentration on the pH of hydrochloric acid (HCl) is shown in the graph.



- Identify** the trend shown by this graph.
 - Write an equation to show the ionisation of hydrochloric acid.
15. **Analyse data**
- The pigments found in red cabbage can act as acid–base indicators. Depending on the type of substance being tested, a different colour change can be observed in a red cabbage solution.

pH	Colour
1	Red
3	Pink
5	Light purple
7	Dark purple
9	Green

A student conducted an experiment to identify the acidity of soil samples taken from different areas. Red cabbage indicator was used for this experiment. The results are shown below.

Sample	Observation
A	Pink
B	Green
C	Pink/light purple

It has been identified that soil irrigated with wastewater can cause the soil pH to decrease. Based on the results above, **identify** which soil sample is most likely taken from an area with high wastewater irrigation. **Explain** your answer.

SCIENCE AS A HUMAN ENDEAVOUR

Syllabus dot point

- Appreciate that blood plasma is an aqueous solution containing a range of ionic and molecular substances.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Blood plasma – ‘liquid gold’

Dissolved solutes are essential for many of life’s processes. Blood is the body’s chemical transport system that relies on molecules being transported in aqueous solution. Plasma makes up about 55% of blood, with other solid components such as red and white blood cells making up the remaining 45% (Figure 1). Plasma is an aqueous solution containing a range of ionic and molecular substances. Plasma itself is composed primarily of water, with solute chemistry playing a large role in how substances are moved around the body.

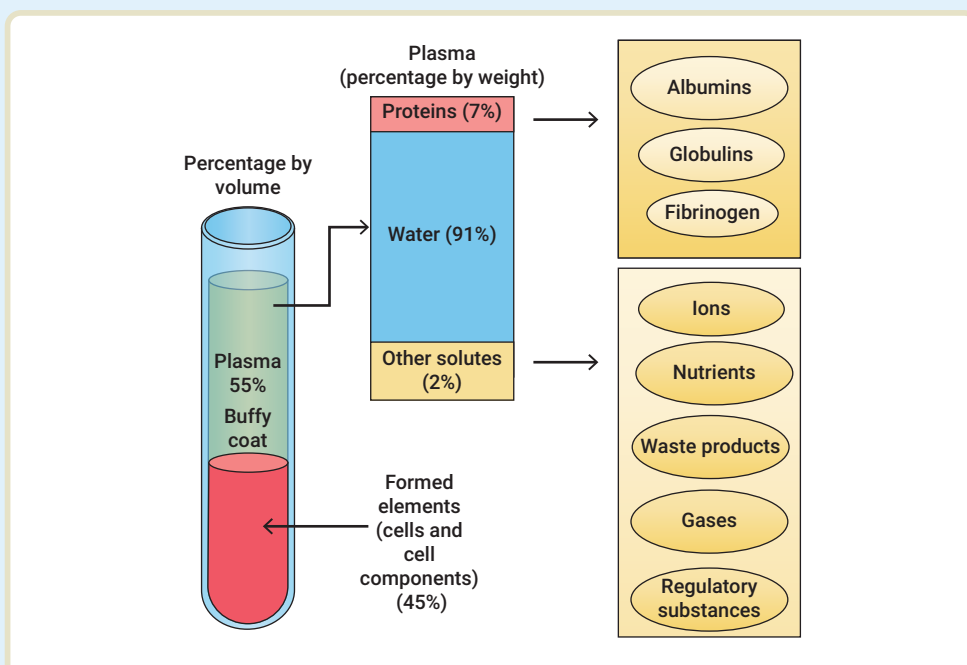


FIGURE 1 Components of blood; plasma makes up 55% of the total volume of blood.

Plasma is composed mainly of water, but also carries enzymes, nutrients, proteins, hormones, respiratory gases and ionic salts. The role of plasma is to transport these molecules around the body, providing oxygen and nutrients to cells and removing wastes. Aside from keeping nutrients flowing around the body, plasma contains dissolved ions with an important role in regulating blood chemistry and enabling muscles to work properly. Beyond its circulatory function, plasma is a treasure trove of important proteins that not only prevent uncontrolled bleeding but are also defenders against bacteria, viruses and other harmful agents (Table 1).

TABLE 1 Various components that make up plasma

Plasma components	Functions and examples
Water	Acts as a solvent and suspending medium for blood components
Proteins	Maintain osmotic pressure (albumin), destroy foreign substances (antibodies and complement), transport molecules (albumin and globulins) and form clots (fibrinogen)
Ions	Involved in osmotic pressure (sodium and chloride ions), membrane potentials (sodium and potassium ions) and acid–base balance (hydrogen, hydroxide and bicarbonate ions)
Nutrients	Source of energy and ‘building blocks’ of more complex molecules (glucose, amino acids and triglycerides)
Waste products	Breakdown products of protein metabolism (urea and ammonia salts) and red blood cells (bilirubin)
Regulatory substances	Catalyse chemical reactions (enzymes) and stimulate or inhibit many body functions (hormones)

To be healthy, it is vital that normal blood solute concentrations and pH levels are maintained. Changes in the solute concentrations in the blood can indicate a range of health conditions. For example, changed sugar levels in plasma can indicate diabetes. Pathologists test blood plasma levels and compare sample plasma concentrations to reference values that reflect the normal ranges in plasma levels (Figure 2). These comparisons are analysed to infer the presence of disease.

In Australia, donated plasma is used to make 15 medical products. There are three main types of product: immunoglobulins, clotting factors and albumin. Immunoglobulins are used to treat immune disorders and to help the immune system fight off disease. Clotting factors can be used to treat bleeding disorders in which blood doesn’t clot normally. Albumin is used to stabilise the fluids in someone suffering from blood loss.

Dragos Condrea/Alamy Stock Photo



FIGURE 2 Pathologists test blood plasma levels to determine if solute concentrations are in the normal range.

Reference

Wong, J., & Dennington, P. (2017). Plasma – the liquid gold running through our veins, Australian Academy of Science, <https://www.science.org.au/curious/people-medicine/plasma>

Rates of reactions



ivankifa/Shutterstock.com

SYLLABUS
DOT POINTS

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, to 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.

SCIENCE INQUIRY

- Investigate factors that affect the rate of chemical reactions.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Introduction

Just as some people run faster than others, some chemical reactions occur faster or slower than others. As scientific understanding of what happens during a chemical reaction has deepened, scientists have applied this knowledge to control the speed, or rate, of a reaction to suit our purposes.

Practicals

- Modelling a reaction
- Effect of concentration on the rate of reaction
- Effect of surface area on rate of reaction
- Effect of temperature on the rate of reactions
- Effect of a catalyst on the rate of a reaction

Worksheets

- Collision theory
- Modelling successful and unsuccessful collisions
- Constructing energy profile diagrams
- Graphing and interpreting results
- Factors that affect rate of reaction



 Nelson MindTap

To access resources above, visit
cengage.com.au/nelsonmindtap

ASSUMED KNOWLEDGE

- ✓ Concentration is a measure of the number of particles in a given volume.
- ✓ mol L^{-1} is a unit of concentration.
- ✓ Gas pressure is due to collisions of the gas particles with the walls of their container.
- ✓ The pressure and volume of a gas are inversely related.
- ✓ Chemical reactions involve breaking bonds in reactants and forming new bonds in the products.
- ✓ Chemical reactions can be exothermic or endothermic.
- ✓ When chemical reactions occur, mass and energy are conserved.
- ✓ The symbol ΔH represents enthalpy change in a chemical reaction.

LEARNING OBJECTIVES

By the end of this chapter, you should be able to:

- ✓ identify the rate of reaction as a change in one quantity, including the concentration, mass or volume of reactants or products, over time
- ✓ calculate the rate of a chemical reaction using data in tables or graphs
- ✓ explain collision theory
- ✓ recognise and sketch Maxwell–Boltzmann distribution curves
- ✓ define ‘activation energy’
- ✓ relate activation energy and activated complex
- ✓ explain how the nature of reactants is related to activation energy
- ✓ explain the relationship between activation energy and rate of reaction
- ✓ recognise and sketch energy profile diagrams, identifying enthalpy change, activation energy and activated complex
- ✓ define ‘catalyst’
- ✓ describe and investigate factors that affect the rate of a reaction, including temperature, concentration, pressure, surface area and presence of a catalyst
- ✓ apply collision theory to explain how concentration, pressure, temperature, surface area and catalysis affect the rate of a reaction
- ✓ represent the effect of catalysts in Maxwell–Boltzmann distribution curves and energy profile diagrams.

19.1 The rate of a chemical reaction

Rate of a reaction

rate how quickly one quantity (a dependent quantity) changes compared to another (an independent quantity)

concentration the number of particles in a given volume

Rate is how quickly one quantity changes compared with another quantity. The rate at which a car moves is how many kilometres the car travels in an hour. The rate at which someone loses weight might be the weight loss (often in kilograms) per week. Another common rate is rate of pay, which describes how much someone is paid per time period. In terms of reactions, the rate is how quickly a reaction occurs (i.e. how the quantity or **concentration** of one or more reactants or products changes with time (**Figure 19.1.1**).



Mark Fergus Photography

FIGURE 19.1.1 Which sugar would dissolve the fastest? How could you measure it accurately to compare the three?

Measuring the rate of a reaction

To calculate any rate, divide the change in one quantity by the change in the other quantity:

$$\text{Rate} = \frac{\text{change in quantity 1}}{\text{change in quantity 2}}$$

For example, at the 2020 Tokyo Olympic Games, Emma McKeon won the 100-metre freestyle swim in 51.96 seconds. This means her average rate of swimming, or speed, was:

$$\begin{aligned}\text{Rate} &= \frac{100}{51.96} \\ &= 1.92 \text{ m s}^{-1}\end{aligned}$$

To measure the rate of a reaction, you need something measurable to change, such as:

- mass
- colour
- volume
- pH
- concentration.

If you can measure this quantity at different times during the reaction then you can determine how quickly it is changing during the reaction (**Figure 19.1.2**). For example, in the reaction between solid calcium carbonate ($\text{CaCO}_3(\text{s})$) and a solution of hydrochloric acid ($\text{HCl}(\text{aq})$), the mass of the solid decreases as it reacts and forms the products calcium chloride in solution ($\text{CaCl}_2(\text{aq})$), carbon dioxide gas ($\text{CO}_2(\text{g})$) and water ($\text{H}_2\text{O}(\text{l})$), according to the equation:



Science Photo Library/Martyn F. Chillmaid

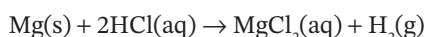
FIGURE 19.1.2 Two beakers with the same reactants. Which one is reacting fastest? How do you know?

A number of different quantities, including change in mass, change in concentration of H^+ ion and how much gas is produced, can be used to measure the rate of this reaction. You could measure the mass of the solid at the start of the reaction and at end of the reaction, and compare it to the time it takes for the reaction to occur. You can then calculate the rate of the reaction. For example, if 2.00 g of calcium carbonate takes 48.0 s to completely react, the average rate of reaction over that time interval would be:

$$\begin{aligned}\text{Rate} &= \frac{\text{change in mass}}{\text{change in time}} \\ &= \frac{2.00}{48.0} \\ &= 4.17 \times 10^{-2} \text{ g s}^{-1}\end{aligned}$$

However, a rate does not have to be calculated for the overall reaction. You can also calculate the rate at various intervals of the reaction by taking measurements throughout the reaction time. These rates may then be used to calculate the rate directly, or you can graph the changes that occur during the reaction.

Consider the reaction between magnesium metal and hydrochloric acid in solution:



The mass of the hydrogen gas produced during the reaction can be calculated by the mass lost from a flask (Figure 19.1.3 and Table 19.1.1).

Once the changes that occur during the reaction are graphed (e.g. the mass of hydrogen gas produced), the slope or **gradient** represents the rate of a reaction because it shows how quickly the reactants or products change. The graph in Figure 19.1.4 shows that reaction rate was greatest at the start of the reaction and gradually slowed to zero at the end of the reaction. The raw data in Table 19.1.1 supports this, showing that the changes in mass are greatest at the start of the reaction.

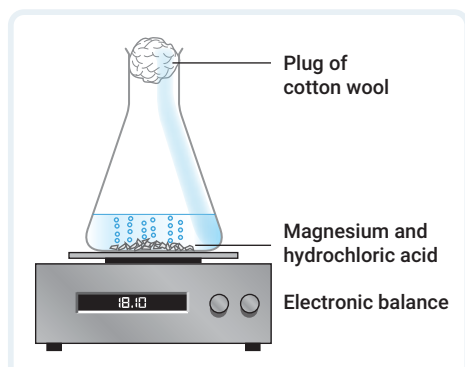


FIGURE 19.1.3 Measuring the mass of hydrogen gas lost during the reaction

gradient the steepness of the line on a graph, which can be used as a measure of the rate of change of the quantity measured on the y-axis (the dependent variable) relative to that on the x-axis (the independent variable)

TABLE 19.1.1 Changes in mass during the reaction between magnesium and hydrochloric acid

Time (s)	Mass of flask and contents (g)	Mass of hydrogen gas released (g)
0	20.00	0.00
10	18.60	1.40
20	18.30	1.70
30	18.20	1.80
40	18.15	1.85
50	18.10	1.90
60	18.06	1.94
70	18.04	1.96
80	18.02	1.98
90	18.01	1.99
100	18.00	2.00
110	18.00	2.00
120	18.00	2.00

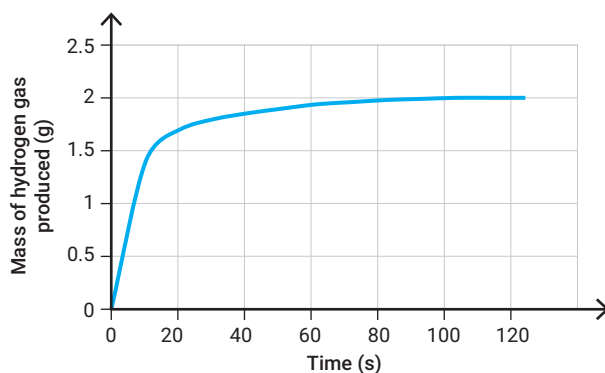


FIGURE 19.1.4 Mass of hydrogen gas produced during the reaction

If concentration had been used to measure the rate of this reaction, then the following key formula would apply.

KEY FORMULA

Rate of reaction can be measured as:

$$\text{Rate of reaction} = \frac{\text{increase in product concentration } (\Delta[\text{P}])}{\text{time taken}}$$

or

$$\text{Rate of reaction} = \frac{\text{decrease in reactant concentration } (-\Delta[\text{R}])}{\text{time taken}}$$

LEARNING CHECK 19.1

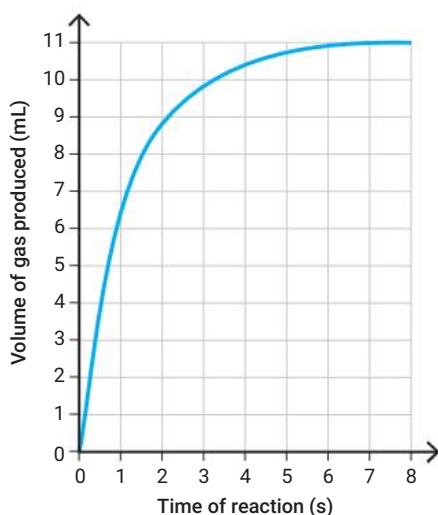
DESCRIBING

- 1 **Define** 'rate'.
- 2 **Identify** the data you would need to record to calculate or compare the rates of different reactions.
- 3 **Describe** when the rate is the highest during a reaction.
- 4 **Outline** how to calculate the average rate of a reaction over a given time interval.
- 5 **Explain** how the slope (or gradient) of a graph indicates the rate.

APPLYING

- 6 During a reaction, 20 g of product forms in 4 minutes. **Calculate** the reaction rate.
- 7 Copy the following graph and draw tangents to the curve at times 0, 1 and 5 seconds to **calculate** the rate of reactions at each of these times.

Production of carbon dioxide gas during the decomposition of copper(II) carbonate



- 8 A student was interested in finding out if the rate of a reaction would be the same if a different quantity was measured. They performed the reaction between magnesium metal and hydrochloric acid described above, using the same quantities and concentrations. However, instead of measuring the mass of hydrogen gas produced they used a pH meter to measure the change in H^+ concentration. Table 19.1.2 shows the data they obtained.

TABLE 19.1.2 Changes in concentration of H^+ during reaction

Time (s)	0	10	20	30	40	50	60	70	80	90	100	110	120
$[\text{H}^+]$ mol L^{-1}	2.0	0.60	0.30	0.20	0.15	0.10	0.06	0.04	0.02	0.01	0.00	0.00	0.00

- a Use the data in Table 19.1.2 to draw a graph of concentration of H^+ vs time.
- b **Calculate** the overall rate of reaction using the data in:
 - i Table 19.1.1
 - ii Table 19.1.2.
- c **Compare** the graph you drew in part a with Figure 19.1.5 and the rates you calculated in part b to **comment** on the question, 'Does the rate of this reaction vary with the quantity being measured?'
- d **Suggest** how you could check if your answer applied to other reactions.

19.2 What is needed for a chemical reaction to occur?



Weblink
Chemical reactions

To understand why some reactions occur faster than others, it is important to understand what happens to the reacting species during a reaction.

PRACTICAL ACTIVITY 19.2.1

MODELLING A REACTION

Research question

Can models be used to demonstrate what occurs during a chemical reaction?

Aim

To use molecular models to represent a chemical reaction

Materials

- molecular model kit (e.g. molymod®)

Procedure

- 1 Using the kit, make two molecules of hydrogen gas (H_2) and one molecule of oxygen gas (O_2) as shown in **Figure 19.2.1**.
- 2 Use these molecules to make two molecules of H_2O .
- 3 Describe what had to happen for the reactants (H_2 and O_2) to become the product (H_2O).

Interpretation

- 1 Based on the reaction you modelled, describe what has to happen for the reactants in a chemical reaction to become products?

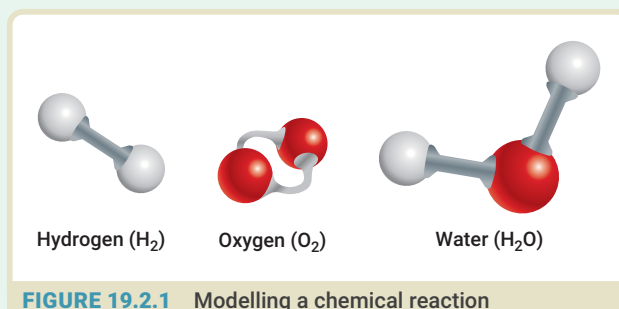
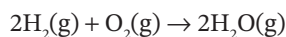


FIGURE 19.2.1 Modelling a chemical reaction

Consider the reaction between hydrogen gas and oxygen gas that you modelled in Practical activity 19.2.1:



For this reaction to occur:

- bonds in hydrogen molecules must break to separate the hydrogen atoms
- bonds in oxygen molecules must break to separate the oxygen atoms
- new bonds must form between the hydrogen and oxygen atoms to make the water molecules.

For this reaction to occur, there needs to be enough energy to overcome the forces of attraction holding the atoms of the reactants, hydrogen and oxygen, in their original bonds; that is, enough energy to break the covalent bonds within the hydrogen molecules and oxygen molecules.

This reaction also requires the newly separated hydrogen and oxygen atoms to contact each other to form the new bonds of the product. This means that two hydrogen atoms must contact one oxygen atom to form a water molecule.



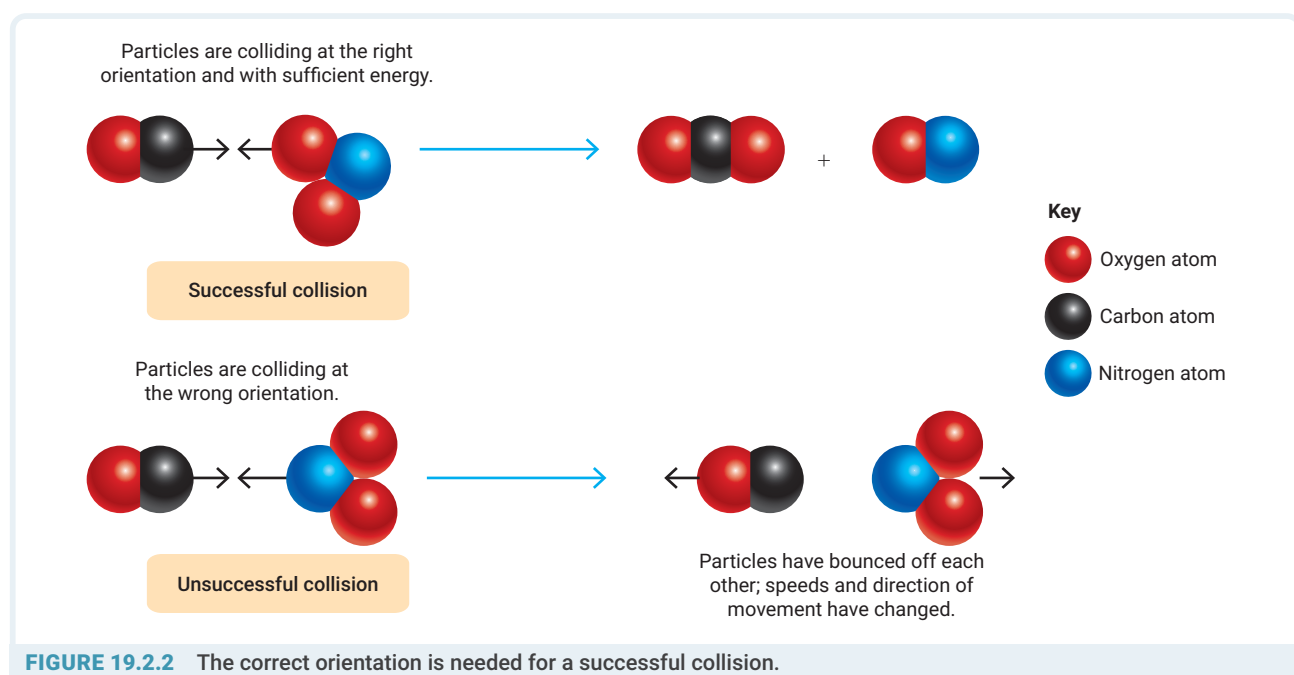
Worksheet
Collision theory

Weblink
Collision theory

collision theory for a reaction to occur, the particles must collide with sufficient energy and in the required orientation

Collision theory

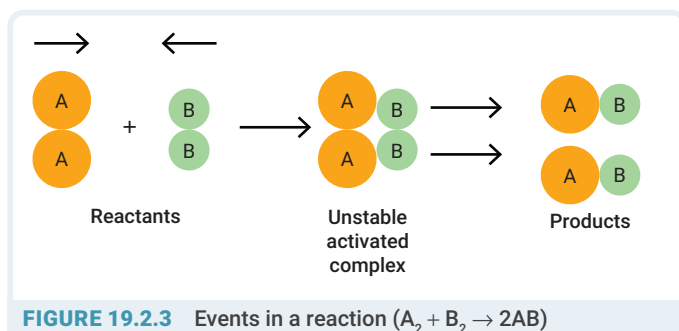
To understand why reactions occur quickly or slowly, you need to know what happens during a chemical reaction. The **collision theory** is a model that allows us to explain what happens during a chemical reaction and the factors that affect it. The collision theory states that, for a reaction to occur, the reactant particles must collide with sufficient energy and at the required orientation. Collisions between reactant particles are classified as either 'successful' or 'unsuccessful', as shown in **Figure 19.2.2** – the reaction between carbon monoxide and nitrogen dioxide to form carbon dioxide and nitrogen monoxide. A successful collision has enough energy and the correct orientation to break bonds in the reactants and make new bonds in the products. In an unsuccessful reaction, the energy, the orientation or both are not satisfactory. The more successful collisions there are in a given time period, the higher the reaction rate because more reactants are forming products.



activated complex the intermediate state of a chemical reaction in which bonds are breaking and forming

During the reaction, as reactants form products, a brief intermediate state occurs when bonds are breaking and forming. This produces a temporary, highly unstable structure called the **activated complex** (**Figure 19.2.3**). An activated complex is a structure that has the maximum energy during the reaction.

The activated complex quickly forms new bonds to produce the products, releasing energy. This means that the required bonding atoms must be positioned next to each other to form the new bonds. The orientation of the colliding particles needs to be correct.



Worksheet
Modelling successful and unsuccessful collisions

Energy profile diagrams

The changes in energy that occur as atoms and molecules collide during the progress of reaction can be shown in an **energy profile diagram**. The total energy possessed by a chemical substance is called **enthalpy**, represented by the symbol H . Enthalpy can be thought of as being mainly the chemical energy stored in a substance. Although we cannot measure the total enthalpy of a substance, we can measure changes in enthalpy. The change in enthalpy for a chemical reaction is represented by the symbol ΔH . In a reaction, the reactants, activated complex and products each have their own enthalpies. These can be represented on an energy profile diagram, which shows the relative enthalpies of the components.

Energy is needed to break the bonds of the reactants. The activated complex has more enthalpy than the reactants. Similarly, energy is lost from the activated complex when the bonds of the products are formed. The activated complex also has more enthalpy than the products (Figure 19.2.4).

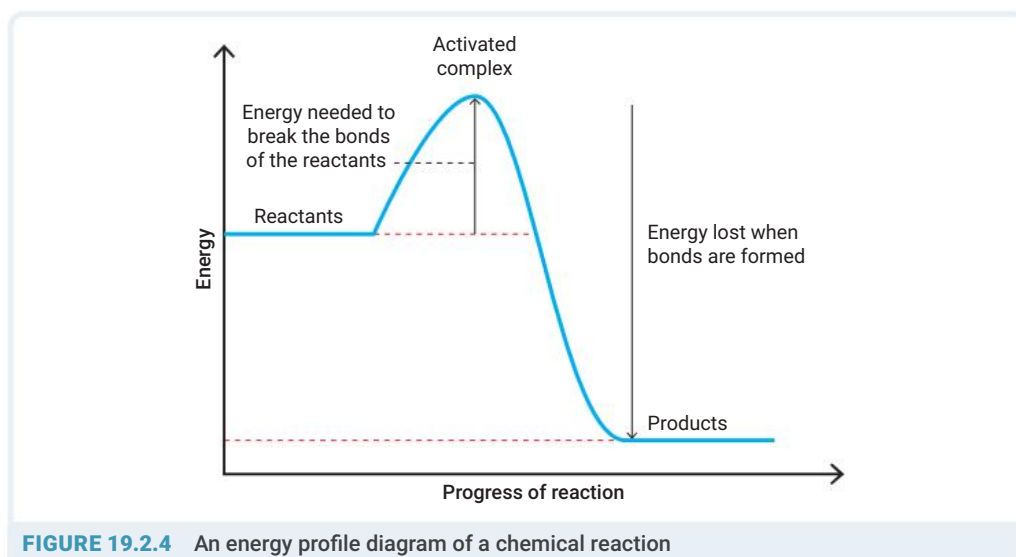


FIGURE 19.2.4 An energy profile diagram of a chemical reaction

If more energy is lost when products form than is gained to make the activated complex, then the enthalpy difference is lost from the system to the surroundings. This occurs in an exothermic reaction and leads to a negative ΔH value. In contrast, if more energy is gained than is lost, the excess must be gained from the surroundings. This occurs in an endothermic reaction and leads to a positive ΔH value (Figure 19.2.5).

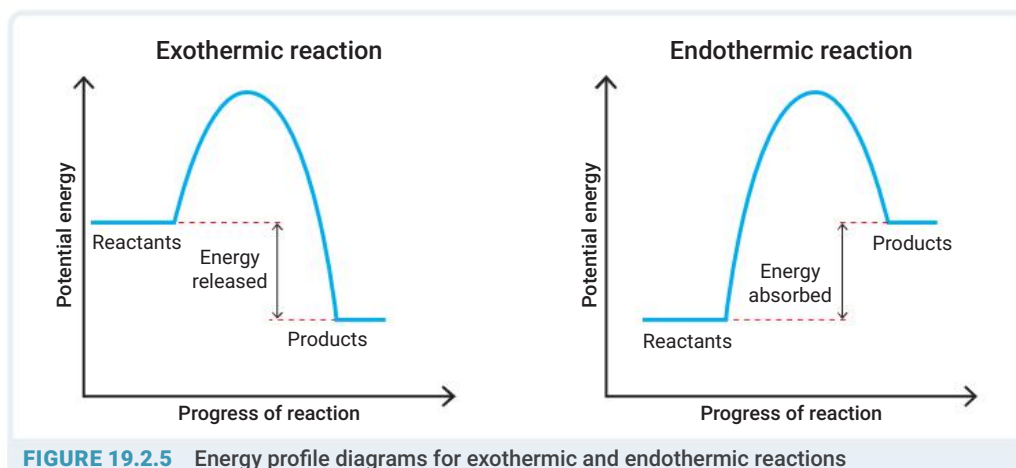


FIGURE 19.2.5 Energy profile diagrams for exothermic and endothermic reactions

energy profile diagram a diagram showing relative enthalpies (energies) during the progress of a reaction

enthalpy (H) the total energy content of a chemical substance



Weblink
Energy profile diagrams



Worksheet
Constructing energy profile diagrams

activation energy (E_a) the minimum energy required for a chemical reaction to occur

Activation energy

For a reaction to be successful, enough energy has to be absorbed to form the activated complex. The amount of energy needed to form the activated complex is called the **activation energy** and is often represented as E_a . This is the energy needed for the bonds of the reactants to break, so the activation energy is the minimum energy required for a reaction to occur. The activation energy can be determined from the difference in energy between the activated complex and the reactants (**Figure 19.2.6**).

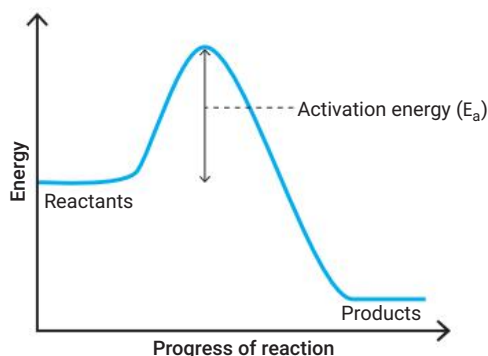


FIGURE 19.2.6 The activation energy of a reaction

Energy and the Maxwell–Boltzmann distribution

As just mentioned, energy is needed for the activated complex to form. Typically, the source of activation energy is heat, which reactant molecules absorb from the surroundings. This thermal energy speeds up the motion of the reactant molecules, increasing the forces and frequency of their collisions. So why do some molecules have more energy than others and can form an activated complex?

At a given **temperature**, particles have different amounts of kinetic energy, with the average kinetic energy of all the particles determining the temperature. This is represented in the **Maxwell–Boltzmann distribution**, which graphically shows the kinetic energies of the particles at a particular temperature. (**Figure 19.2.7**).

temperature a measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

Maxwell–Boltzmann distribution a frequency distribution of the energy value of a sample of particles at a particular temperature



Weblink
Maxwell–Boltzmann
distribution

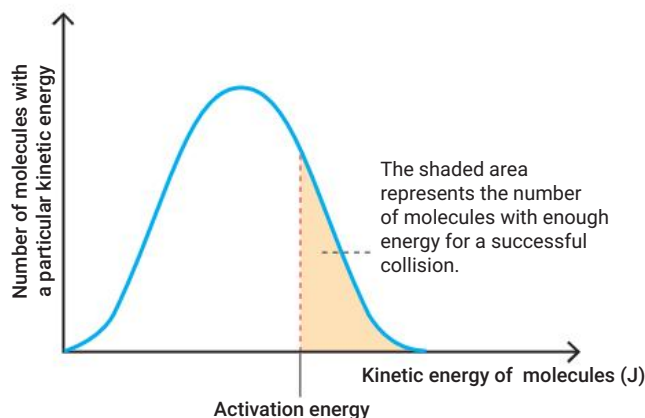


FIGURE 19.2.7 The Maxwell–Boltzmann distribution shows the kinetic energy of molecules at a particular temperature.

The Maxwell–Boltzmann distribution shows that a few particles have low kinetic energy, and a few have high kinetic energy. Most particles have a kinetic energy value in the middle range.

Only those particles with a kinetic energy equal to or greater than the required activation energy will be able to form an activated complex. On the graph in Figure 19.2.7, the proportion of particles with that amount of energy is represented by the shaded yellow area under the curve.

Activation energy and nature of reactants

The nature of reactants refers to what the reactants are and what has to happen to them to form products. If the bonds of the reactants that need to be broken are weak or there are only a few of them, then they will only require a small amount of energy to break. This means that the activation energy will be low and a larger number of particles will have enough energy for successful collision. Conversely, if the bonds of the reactants are strong or there are a lot of bonds, then more energy is needed to break them so there will be fewer particles with enough energy for successful collisions.

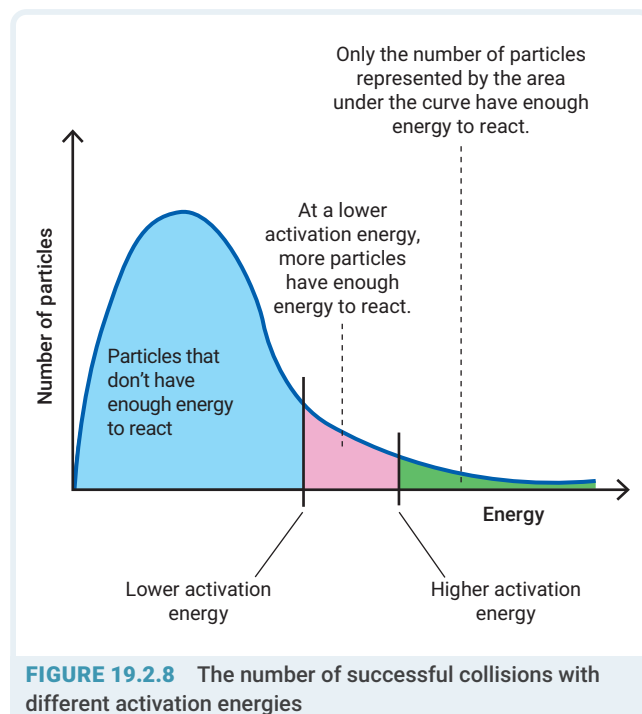
The relationship between activation energy and number of successful collisions is shown in **Figure 19.2.8**.

Activation energy and rate of reaction

The rate of a reaction is closely related to the activation energy. A reaction can only be completed once an activated complex has formed. Therefore, the more energy that is needed for this to occur (i.e. the higher the activation energy barrier), the fewer molecules there will be with enough energy to overcome the barrier.

Generally, reactions with a high activation energy occur more slowly because fewer molecules have sufficient energy to overcome the activation energy barrier and form an activated complex. In fact, many reactions with high activation energies do not proceed at all until more energy is added to system. For example, for a fuel to begin burning, a source of energy, such as a spark or lit match, needs to be added. This energy input gets some of the molecules over the activation energy barrier. The energy then released by these molecules as they form products then gets other molecules over the barrier and the reaction becomes self-sustaining.

Conversely, if the activation energy barrier is low, the reaction is more likely to occur quite readily.



LEARNING CHECK 19.2

DESCRIBING

- 1 Outline collision theory.
- 2 **Describe:**
 - a activation energy
 - b activated complex.
- 3 **Sketch** a Maxwell–Boltzmann distribution.
- 4 **Explain** why a successful collision requires:
 - a sufficient energy
 - b the correct orientation.
- 5 **Explain** the relationship between magnitude of the activation energy, chemical nature of the reactants and rate of reaction.

APPLYING

- 6 **Draw** a fully labelled energy profile diagram for a reaction with an activation energy of:
 - a 1000 J and $\Delta H = 500$ J
 - b 200 J and $\Delta H = -300$ J.

19.3 Factors that affect the rate of a chemical reaction

Collision theory tells us that the rate of a reaction depends on the number of successful collisions between reactant particles in a given period of time. The number of successful collisions depends on the:

- total number of collisions
- proportion of collisions that are successful.

If either of these factors increases, the reaction rate can increase. Remember that reactant particles need to collide at a certain orientation with a level of energy that meets the activation energy for a reaction to be successful. As such, simply increasing the total number of collisions may not increase the rate of reaction. Conversely, if either of these factors decreases, the reaction rate will decrease.

How can the total number of collisions be increased? One obvious way is to increase the number of reacting particles. What about speeding up the particles? If the particles have more kinetic energy, they would move faster so there would be more collisions.

How can the number of successful collisions be increased? One way would be to lower the energy needed for a successful collision, i.e. lower the activation energy. Another way would be to give the particles more energy so more had enough energy to overcome the activation energy barrier.

In this section, we will investigate factors that increase the number of collisions and/or make the collisions more successful.

Concentration and pressure

Recall that concentration refers to the number of particles in a given volume. In a gas, the concentration of the gaseous particles determines its **gas pressure**. The higher the concentration, or pressure, the higher the number of collisions. Although the percentage of

gas pressure the force per unit area exerted by gas particles as they collide with the walls of their container

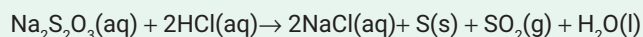
successful collisions remains the same, the total number of successful collisions increases because the overall number of collisions increase. As the concentration, or pressure, of the reactants increases, the reaction rate increases.

PRACTICAL ACTIVITY 19.3.1

EFFECT OF CONCENTRATION ON THE RATE OF A REACTION

Introduction

Sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) reacts with dilute hydrochloric acid (HCl) to produce sodium chloride (NaCl), sulfur (S), sulfur dioxide (SO_2) and water (H_2O):



At a high enough concentration of sulfur, the solution becomes cloudy. The time taken for the solution to become cloudy can be used to measure the rate of the reaction.

Research question

How does changing concentration affect the rate of reaction between sodium thiosulfate and dilute hydrochloric acid?

Aim

To investigate the effect of concentration on the rate of the reaction between sodium thiosulfate and dilute hydrochloric acid

Materials

For one trial:

- 45 mL of 0.25 mol L⁻¹ sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$)
- 45 mL of 0.125 mol L⁻¹ sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$)
- 45 mL of 0.0625 mol L⁻¹ sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$)
- 15 mL of 2 mol L⁻¹ hydrochloric acid (HCl)
- 3 × 100 mL conical flasks
- 50 mL measuring cylinder
- 10 mL measuring cylinder
- stopwatch
- white paper
- pen or marker pen



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Chemicals could splash in your eyes.	Wear safety glasses at all times.
2 mol L ⁻¹ HCl is corrosive to skin and clothes.	Wear gloves and an apron.
SO_2 gas irritates the nose, throat and airways.	Conduct the experiment in a well-ventilated area and do not inhale fumes.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the three risks listed to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- 1 Draw a large cross in the centre of the piece of white paper.
- 2 Accurately measure 45 mL of 0.0625 mol L⁻¹ sodium thiosulfate in a measuring cylinder.
- 3 Pour the sodium thiosulfate into the conical flask.
- 4 Accurately measure 5 mL of 2 mol L⁻¹ HCl in a measuring cylinder.
- 5 Have a stopwatch ready to start timing.
- 6 Pour the hydrochloric acid into the conical flask, starting the stopwatch immediately.
- 7 Quickly mix the two solutions by carefully swirling the conical flask.

- 8 Place the conical flask onto the cross and record the time it takes before you can no longer see the cross when looking from above. (Be careful not to breathe in any fumes from the reaction.)
- 9 Repeat steps 2–8 for the other two solutions of sodium thiosulfate.
- 10 If time and resources allow, repeat steps 2–8 for repeat trials.
- 11 Place your solutions in the waste beaker, and rinse the conical flask.

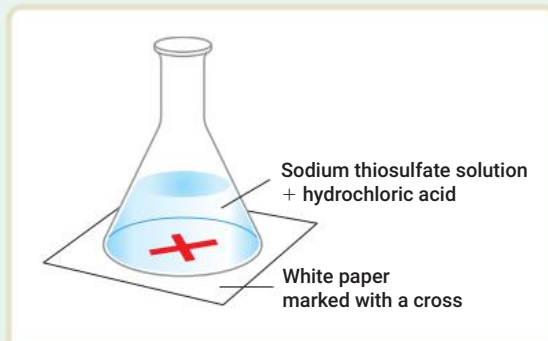


FIGURE 19.3.1 The experimental set-up for the reaction between sodium thiosulfate and dilute hydrochloric acid

Results

Copy and complete the following results table.

Concentration of sodium thiosulfate (mol L^{-1})	Time taken for the cross to disappear (s)			
	Trial 1	Trial 2	Trial 3	Average
0.0625				
0.125				
0.250				

Analysis of results

- 1 Graph the average time taken for the cross to disappear against the concentration of sodium thiosulfate.
- 2 What happened to the time taken for the cross to disappear as the concentration of sodium thiosulfate increased?
- 3 How is the time taken for the cross to disappear related to the rate of the reaction?

Interpretation

- 4 What is the relationship between the concentration of sodium thiosulfate and the rate of the reaction?
- 5 Does your data support the theoretical trend?
- 6 Were there any anomalies? If yes, why do you think they occurred?
- 7 Use collision theory to explain the relationship between concentration and the rate of the reaction.

Evaluation

- 8 Where might errors have occurred in the method that would have affected the accuracy of the results?
- 9 How could the method be improved to increase the accuracy and reliability of the results?
- 10 How could you further investigate the effect of concentration on the rate of the reaction in this reaction?
- 11 How could you further investigate the effect of concentration on the rate of reactions in general?

Changing the volume may also change the concentration. In a solution, adding water dilutes the solution, thereby decreasing the concentration. Alternatively, evaporating off some of the water decreases the volume and increases the concentration. In a gas, if the volume of the container is increased, then the concentration of all the gases decreases. Alternatively, if the volume of the container is decreased, then the concentration of all the gases increases.

During a reaction, as the reactants form the product, their concentration decreases. By applying the collision theory, scientists can predict that reaction rate will decrease as time progresses during the reaction. On a graph of the concentration of reactants or products against time, the reaction rate is indicated by the slope of the graph. The slope will be steepest at

the start of the reaction and will gradually decrease to zero at the completion of the reaction (Figure 19.1.4).

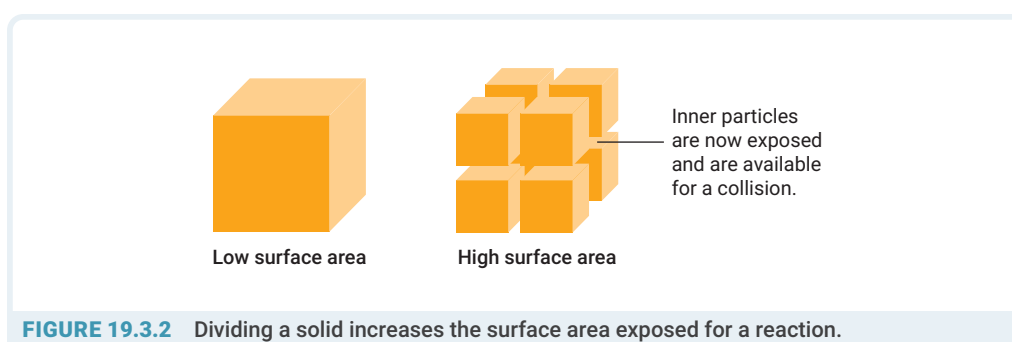
Surface area

For a reaction to occur, the reactant particles need to collide. When one of the reactants is a solid, only the particles on the surface of the solid are available for collision. Once these have reacted, then the particles beneath become exposed, and so on.

If the solid is divided or crushed, the inner particles become exposed. Although the amount of reactant remains the same, the total **surface area** has increased. More of the solid is available to react (Figure 19.3.2). There are more collisions and the reaction rate increases.

This method is used to increase the reaction rate in many everyday activities: chopping wood into smaller pieces of kindling to start the fire; cutting potatoes into small pieces to decrease cooking time; and using granulated sugar instead of sugar cubes in drinks.

surface area the total area of all surfaces



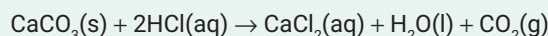
Worksheet
Graphing and
interpreting results

PRACTICAL ACTIVITY 19.3.2

EFFECT OF SURFACE AREA ON RATE OF REACTION

Introduction

Calcium carbonate ($\text{CaCO}_3(\text{s})$) reacts with dilute hydrochloric acid (HCl) according to the equation:



Bubbles of carbon dioxide are visible during the reaction. When no more bubbles are produced, the reaction has stopped. The time it takes for the bubbles to stop can be measured to compare the rates of reactions.

In this experiment, you will change the surface area of a fixed mass of solid CaCO_3 . By comparing the time taken for the reactions to occur, you can determine the effect of surface area on the rate of the reaction.

Research question

How does increasing the surface area of calcium carbonate affect the rate of reaction with dilute hydrochloric acid?

Aim

To determine the effect of surface area on the rate of the reaction between calcium carbonate and dilute hydrochloric acid

Materials

- 3 chips of calcium carbonate
- calcium carbonate powder
- 50 mL of $2 \text{ mol L}^{-1} \text{ HCl}$
- 2 large test tubes and test-tube rack
- 10 mL measuring cylinder
- stopwatch
- electronic balance



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Chemicals may splash in your eyes.	Wear safety glasses at all times.
2 mol L ⁻¹ HCl is corrosive to skin and clothes.	Wear gloves and an apron.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the two risks listed to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- 1 Measure 8 mL of HCl in a measuring cylinder.
- 2 Weigh and record the mass of three calcium carbonate chips.
- 3 Place the calcium carbonate chips in one test tube.
- 4 Have a stopwatch ready to start timing.
- 5 Pour the hydrochloric acid into the test tube and start the stopwatch.
- 6 Stop the stopwatch when no more bubbles are produced, and record this time.
- 7 Carefully weigh the same mass of calcium carbonate powder as that of the calcium carbonate chips.
- 8 Repeat steps 3–6 for the powder.
- 9 If time permits, repeat steps 1–8 for a number of trials.
- 10 Place all the product solutions into the waste beakers as directed by your teacher.

Results

Copy and complete the following results table.

Form of calcium carbonate	Time for the reaction to complete (s)			
	Trial 1	Trial 2	Trial 3	Average
Chips				
Powder				

Analysis of results

- 1 Which form of calcium carbonate took the least time to react?
- 2 Which form of calcium carbonate has the fastest rate of reaction?
- 3 Which form of calcium carbonate had the greatest surface area? Explain why.

Interpretation

- 4 What is the relationship between surface area and the rate of reaction?
- 5 Do your results support the theory about the effect of surface area on the rate of reaction?

Evaluation

- 6 Was this a fair test? Explain why or why not.
- 7 Were the results reliable? Explain why or why not.
- 8 How could the method be improved to more fairly and accurately test the effect of surface area on the rate of reaction between calcium carbonate and dilute hydrochloric acid?
- 9 How could you further investigate the effect of surface area on the rates of reactions?

A similar effect to increasing the surface area is seen when the reactants are stirred. Stirring mixes the reactants together, exposing more of the reactants to each other. This increases the number of collisions and the reaction rate. This phenomenon is demonstrated when adding sugar to a cup of tea or coffee. By stirring the sugar, the rate at which it dissolves increases.



Mark Fergus Photography

FIGURE 19.3.3 Stirring has the same effect as increasing the surface area because more particles are exposed to one another.

Temperature

The temperature of a substance is a measure of the average kinetic energy of the particles. This energy is due to movement and depends on the speed, or velocity, of the particles. As the temperature of a substance increases, so does the average kinetic energy and the average speed of the particles. For example, applying heat to a liquid or gas causes the particles to move faster, collide more energetically, increase the average kinetic energy and, in turn, the temperature (**Figure 19.3.4**).

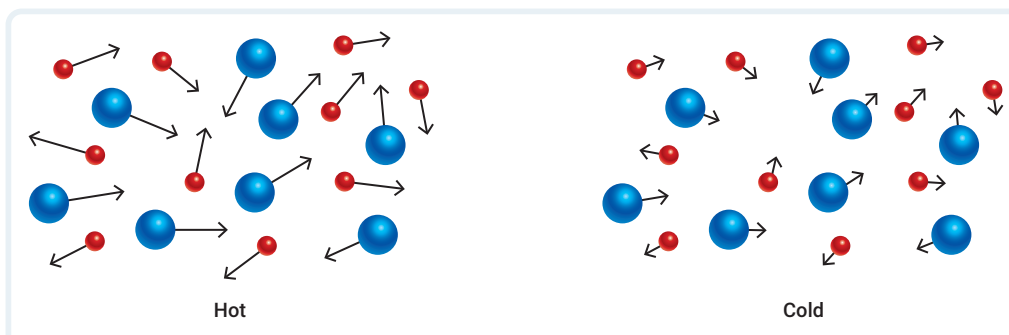


FIGURE 19.3.4 Temperature changes the speed at which particles move (as represented by the length of the arrows).



Weblink

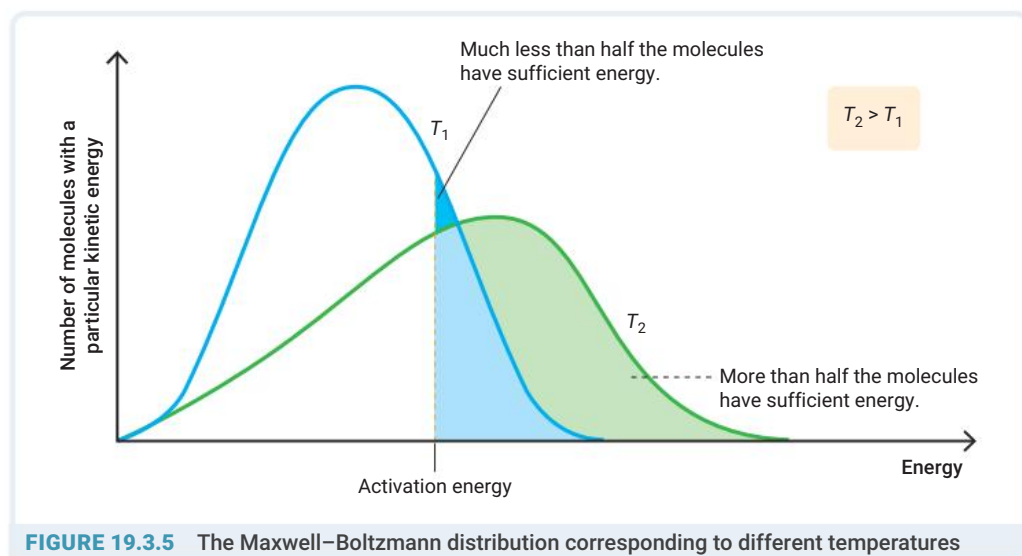
Factors affecting rate of reaction simulation

As the temperature of reactants increases, both the number of collisions and the percentage of successful collisions increase, because temperature has a dual effect on reaction rate.

Increased percentage of successful collisions

As the temperature increases, the average kinetic energy of the particles increases. More particles have enough energy to overcome the activation energy, form the activated complex and become products. A greater percentage of collisions are successful, increasing the reaction rate.

Figure 19.3.5 shows that at a higher temperature (T_2) the Maxwell-Boltzmann distribution moves to the right, increasing the area under the curve to the right of the activation energy. The proportion of particles with enough energy for a successful collision increases.



Note that although the shape of the curve has changed, the total number of particles in the reaction remains constant.

Increased number of collisions

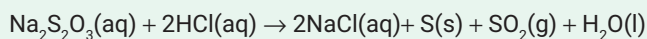
At a higher temperature, the particles have a greater average speed. They collide with each other more frequently, increasing the number of collisions and the rate of the reaction. However, increasing the number of collisions does not affect the reaction rate as much as the effect of the increased energy, which enables a greater number of successful collisions because there are more particles with the energy equal to, or greater than, the activation energy.

PRACTICAL ACTIVITY 19.3.3

EFFECT OF TEMPERATURE ON THE RATE OF REACTIONS

Introduction

You have previously seen the reaction between sodium thiosulfate and dilute hydrochloric acid in Practical activity 19.3.1.



In this experiment, you will be investigating the effect of temperature on the rate of the reaction.

Research question

How does temperature affect the rate of reaction between sodium thiosulfate and dilute hydrochloric acid?

Aim

To investigate the effect of temperature on the rate of the reaction between sodium thiosulfate and dilute hydrochloric acid

Materials

- 150 mL of 0.25 mol L⁻¹ sodium thiosulfate (Na₂S₂O₃)
- 15 mL of 2 mol L⁻¹ hydrochloric acid (HCl)
- crushed ice
- hot water
- stopwatch
- 2 × 1 L beakers (or similar container for ice bath and hot water bath)
- 3 × 100 mL conical flasks
- 3 test tubes
- 50 mL measuring cylinder
- 10 mL measuring cylinder
- white paper
- pen or marker pen
- thermometer



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Chemicals may splash in your eyes.	Wear safety glasses at all times.
2 M HCl is corrosive to skin and clothes.	Wear gloves and aprons.
SO ₂ gas irritates the nose, throat and airways.	Conduct experiment in a well-ventilated area, do not breathe in fumes.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the three risks listed to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- 1 Draw a large cross in the centre of a blank piece of white paper.
- 2 Accurately measure 45 mL of the 0.25 mol L⁻¹ Na₂S₂O₃ in a measuring cylinder.
- 3 Pour the Na₂S₂O₃ into a conical flask.
- 4 Accurately measure 5 mL of the 2 mol L⁻¹ HCl in a measuring cylinder.
- 5 Pour the HCl into a test tube and carefully place the test tube upright into the conical flask containing the Na₂S₂O₃. Ensure the HCl does not spill into the conical flask.
- 6 Place some crushed ice and a small amount of water into the 1 L beaker, and carefully place the conical flask containing the test tube into the ice bath. Ensure that there is enough ice to reach the height of the solution in the flask.
- 7 Repeat steps 2–6 with the second conical flask and test tube, using hot water instead of ice to make a hot water bath.
- 8 Repeat steps 2–5 with a third conical flask and test tube but leave at room temperature.
- 9 Measure and record the temperature of the Na₂S₂O₃ in the first flask.
- 10 Have a stopwatch ready to start timing.
- 11 Carefully remove the test tube containing the HCl from the first flask and pour the HCl into that flask, starting the stopwatch immediately.
- 12 Quickly mix the two solutions by carefully swirling the conical flask.
- 13 Place the conical flask onto the cross and record the time it takes before you can no longer see the cross when looking from above. (Be careful not to breathe in any fumes from the reaction.)
- 14 Repeat steps 9–13 for the other two solutions.
- 15 If time and resources allow, repeat steps 2–14 for repeat trials.
- 16 Place your solutions in the waste beaker and rinse the conical flask.

Results

Copy and complete the following results table.

Temperature of sodium thiosulfate (°C)	Time taken for the cross to disappear (s)			
	Trial 1	Trial 2	Trial 3	Average

Analysis of results

- 1 Plot a graph of the average time taken for the cross to disappear against the temperature of the sodium thiosulfate.
- 2 What happened to the time taken for the cross to disappear as the temperature increased?

Interpretation

- 3 What is the relationship between temperature and the rate of the reaction?
- 4 Does your data support the theory about the effect of temperature on the rate of reactions?

Evaluation

- 5 Where may errors have occurred in the experiment and how would these have affected the results?
- 6 Were your results reliable? Explain why or why not.
- 7 How could you improve the experiment if you were to do it again?
- 8 What conclusion can you make about the effect of temperature on the rate of this reaction?
- 9 How could you further investigate the effect of temperature on the rate of reactions?

catalyst a substance that affects the rate of certain reactions by providing an alternative reaction pathway



Weblink
Effect of catalysts

Worksheet
Factors that affect the rate of reaction

nanoparticle a particle with a diameter less than 100 nm that behaves as a whole unit

Catalysts

A **catalyst** affects the rate of a reaction without being consumed in the reaction. Positive catalysts increase the rate of the reaction. Negative catalysts (also referred to as inhibitors) decrease the rate of the reaction. Scientists tend to refer to positive catalysts.

There are various ways that a catalyst can affect a reaction, but all positive catalysts provide an alternative pathway that has a lower activation energy, as shown in **Figure 19.3.6a**. A lower activation energy means that, for a given temperature, more molecules have enough energy for successful collisions, as can be seen in **Figure 19.3.6b**. More successful collisions occur, so there is a faster reaction.

Cars are equipped with catalytic converters. In these devices, the exhaust fumes pass over a layer of a metal catalyst, usually platinum, rhodium and/or palladium, which is incorporated in a heat-resistant ceramic core that provides a large surface area. As the gases pass over the catalyst, the harmful combustion products of nitrogen oxides, volatile organic compounds and carbon monoxide are converted to less harmful substances (mainly carbon dioxide, water, nitrogen and oxygen). This is why the catalytic converter is an important factor in reducing toxic pollution emitted by cars.

Metal **nanoparticles** are being used more frequently as catalysts in scientific and industrial applications. A nanoparticle is a particle with a diameter less than 100 nm that behaves as a whole unit. Metal nanoparticles have an advantage over other metal catalysts because of their large surface area. Large amounts of reactants can come into contact with the catalyst, which has a greater impact on the reaction rate. Scientists research different metal nanoparticles for a number of reactions, including gold and copper nanoparticles as catalysts for fuel cells.

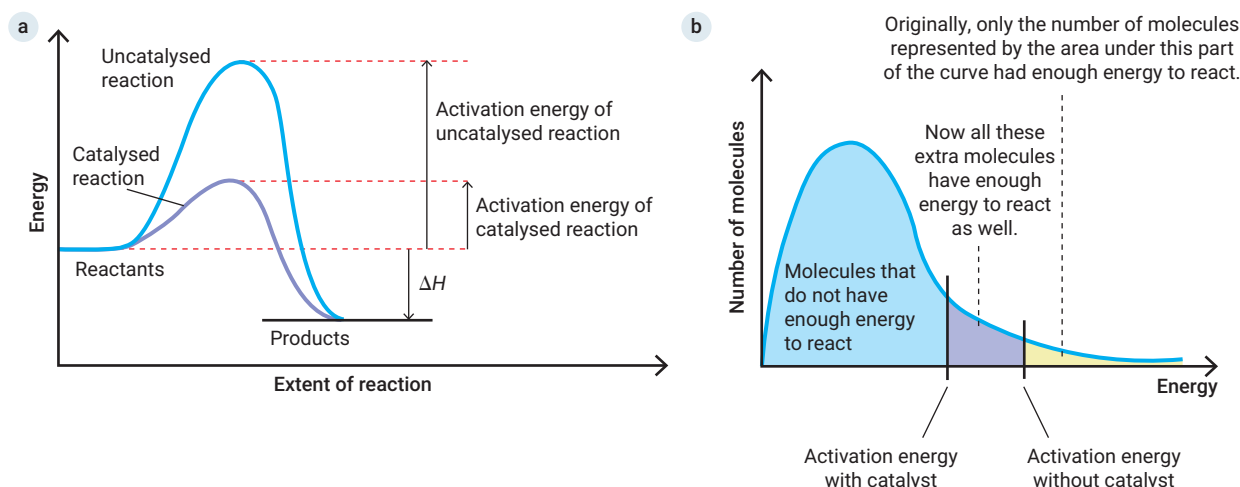


FIGURE 19.3.6 The effect of a catalyst on the number of molecules with sufficient energy for a successful collision: (a) an energy profile diagram of a catalysed and an uncatalysed reaction; and (b) a Maxwell-Boltzmann distribution curve for catalysed and uncatalysed reactions.

Catalysts in our bodies ensure that the reactions needed for a healthy life are occurring at the appropriate rate. Biological catalysts are called **enzymes** and are usually made of **protein**. Enzymes are affected by temperature and pH. As enzymes must work within a living organism, it is logical that the optimal temperature and pH is that of the organism. Outside of these conditions the enzyme is ineffective, and this can often mean that the reaction is not able to occur at the required rate.

enzyme a biological catalyst

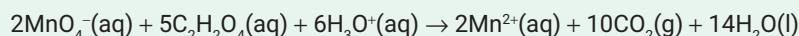
protein a compound consisting of chains of amino acid molecules; proteins are essential for life

PRACTICAL ACTIVITY 19.3.4

EFFECT OF A CATALYST ON THE RATE OF A REACTION

Introduction

Rhubarb contains oxalic acid ($\text{C}_2\text{H}_2\text{O}_4$), which reacts with acidified potassium permanganate (KMnO_4), causing it to decolourise. This can be represented by the following equation:



The time for the colour change to occur can be measured and used to calculate the reaction rate. The rhubarb may be fresh, in frozen rhubarb sticks, or a puree, which can be made by gently boiling rhubarb in water.

Manganese(II) sulfate (MnSO_4) can act as a catalyst for this reaction, increasing the rate of the reaction and decreasing the time taken for the decolourisation to occur.

Research question

How does the use of a manganese(II) sulfate catalyst affect the rate of decolourisation of rhubarb?

Aim

To observe the effect of a catalyst on the rate of the reaction between oxalic acid and acidified potassium permanganate

Materials

- 6 mL of 0.02 mol L⁻¹ of potassium permanganate (KMnO_4)
- 50 mL of 2 mol L⁻¹ sulfuric acid (H_2SO_4)
- 250 mL distilled water
- 6 mL of saturated oxalic acid ($\text{C}_2\text{H}_2\text{O}_4$)
- 1 mL of saturated manganese(II) sulfate (MnSO_4)
- 2 × 250 mL beakers
- 3 × 10 mL, 1 × 25 mL and 1 × 250 mL measuring cylinders
- stirring rod
- stopwatch



What are the risks in doing this experiment?	How can you manage these risks to stay safe?
Chemicals could splash in your eyes.	Wear safety glasses at all times.
2 mol L ⁻¹ H ₂ SO ₄ is corrosive to skin and clothes.	Wear safety glasses, gloves and an apron.
KMnO ₄ is an oxidising agent; it can irritate the skin and eyes.	Wear safety glasses, gloves and an apron.
Oxalic acid is harmful to the skin and eyes.	Wear safety glasses, gloves and an apron.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them. Expand on the four risks listed to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

Procedure

- 1 Into the first beaker, place 25 mL of 2 mol L⁻¹ H₂SO₄, 125 mL distilled water, 3 mL of 0.02 mol L⁻¹ KMnO₄ and 3 mL of saturated C₂H₂O₄.
- 2 Start the stopwatch immediately. Stir the solution. Record how long it takes for the solution to become colourless.
- 3 Into the second beaker place 25 mL of 2 mol L⁻¹ H₂SO₄, 125 mL distilled water, 3 mL of 0.02 mol L⁻¹ KMnO₄, 3 mL of saturated C₂H₂O₄ and 1 mL of saturated MnSO₄.
- 4 Start the stopwatch immediately. Stir the solution. Record how long it takes for the solution to become colourless.
- 5 Follow your teacher's instructions for disposal of chemicals.

Results

Copy and complete the following results table.

Beaker	Time taken for decolourisation (s)
No MnSO ₄	
With MnSO ₄	

Analysis of results

- 1 What was the effect of adding manganese sulfate to the reaction mixture?

Interpretation

- 2 Given that manganese sulfate was acting as a catalyst, did your results support the theory about the effect of catalysts on the rate of reactions?

Evaluation

- 3 How could the experiment be improved to obtain more accurate and reliable data?
- 4 What did you determine about the effect of manganese sulfate as a catalyst on the reaction between oxalic acid and potassium permanganate?
- 5 How could you further investigate the effect of catalysts on the rate of reaction?
- 6 How could you determine if other factors influence the effect of catalysts on the rate of reactions?

LEARNING CHECK 19.3

DESCRIBING

- 1 **Identify** the factors that affect the rate of a reaction.
- 2 **Define** 'catalyst'.
- 3 **Draw** a Maxwell–Boltzmann distribution to show the distribution of energy of particles at two different temperatures. Clearly identify which curve is for the higher temperature.

APPLYING

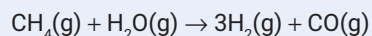
- 4 **Explain** how crushing a lump of solid affects reaction rate.
- 5 **Compare** the average kinetic energy of the particles of substances at 10°C and at 40°C.
- 6 **Explain** how the pressure of a gas is equivalent to its concentration.
- 7 Use collision theory to **explain** how each of the following factors affects the rate of a reaction.
 - a Concentration or pressure
 - b Temperature
 - c Catalyst
 - d Surface area
- 8 Ben and Amy both added a sugar cube to their cups of tea, which were the same temperature. Ben stirred his tea, while Amy left her cup to sit. After 1 minute, they each tasted their tea. **Identify** whose tea would have tasted sweeter. **Justify** your answer.

ANALYSING

- 9 Five test tubes each contained hydrochloric acid (HCl) and magnesium as shown in the following table. **Identify** which test tube would contain the fastest reaction. **Justify** your answer.

Test tube	Concentration of HCl (mol L ⁻¹)	Temperature (°C)	Form of 2 g of magnesium
1	1	25	Lump
2	2	25	Long strip
3	1	10	Strip cut into small pieces
4	0.5	40	Long strip
5	2	40	Strip cut into small pieces

- 10 A common method of preparing hydrogen for industrial purposes is to heat methane and steam with a catalyst at high temperature. The endothermic reaction is:



When 1 mol of methane was reacted with excess steam at 1100°C in a vessel of fixed volume, curve A in the following graph was obtained. When the experiment was repeated at a different temperature, curve B was obtained.

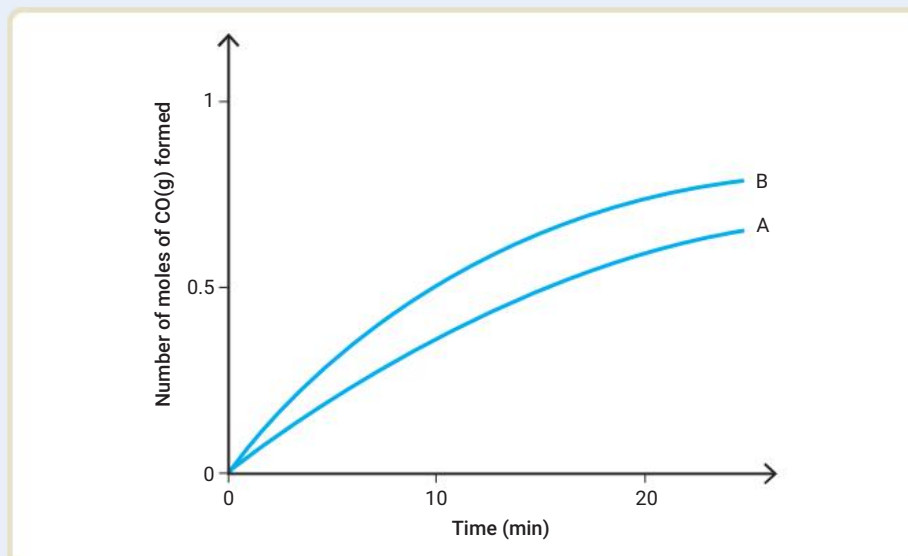


FIGURE 19.3.8 The reaction of methane and steam

- a **Determine** what information the initial slopes of the curves tell you about the initial rates at the two temperatures. **Explain** the reason for this difference.
- b **Deduce** if the temperature for curve B is greater or less than that for curve A. **Explain** your reasoning.

Rate of reaction

- The rate of a reaction can be calculated by measuring a change in concentration, mass, volume of reactants or products over time.
- Factors that affect the rate of a reaction are concentration, pressure, temperature, surface area and the presence of a catalyst.

$$\text{Rate of reaction} = \frac{\text{increase in product concentration } (\Delta[P])}{\text{time taken}}$$

or

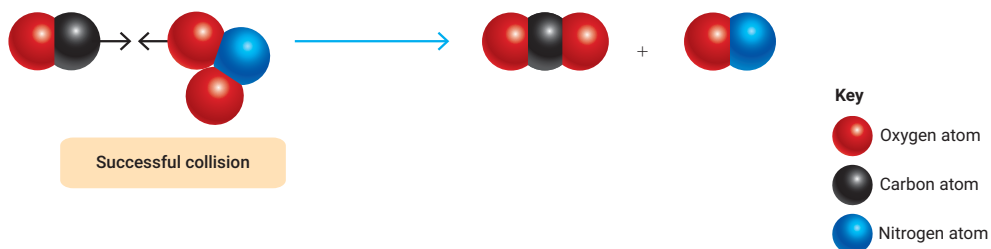
$$\text{Rate of reaction} = \frac{\text{decrease in reactant concentration } (-\Delta[R])}{\text{time taken}}$$

Collision theory

- Collision theory states for a reaction to occur, the particles must collide with sufficient energy and in the correct orientation.
- Collision theory can be used to explain why the rate of reaction is affected by concentration, pressure, temperature, surface area and the presence of a catalyst.

Illustration of how reactions occur

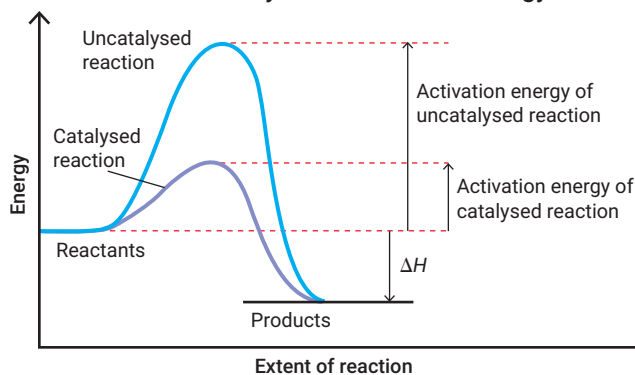
Particles are colliding at the right orientation and with sufficient energy.



Activation energy

- Activation energy, E_a , is the minimum energy required to form the activated complex.
- Catalysts lower the activation energy of a reaction by providing an alternative reaction pathway.

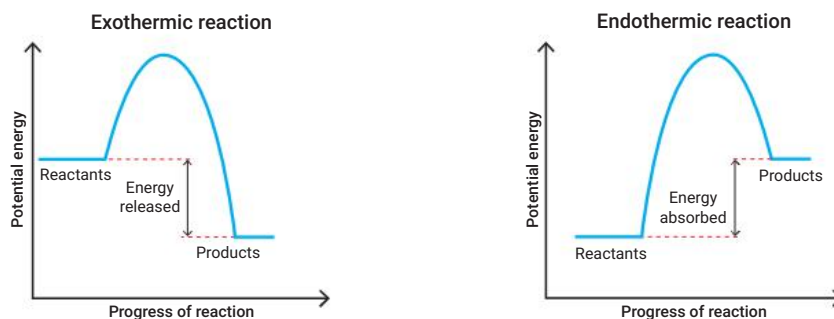
Effect of catalysts on activation energy



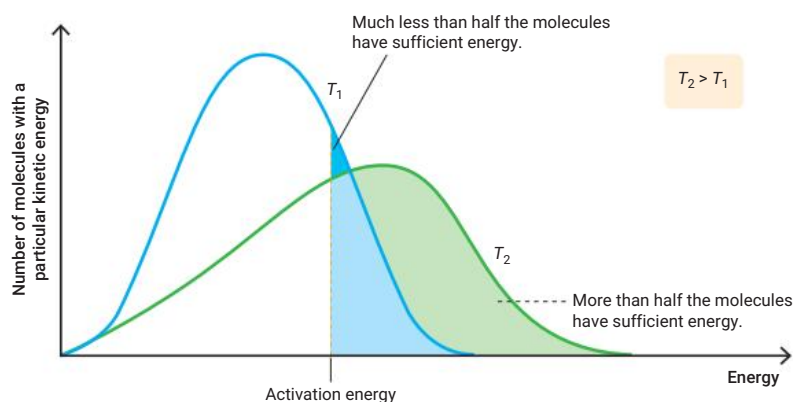
Energy profile diagrams

- An energy profile diagram can be used to show the enthalpy change during a chemical reaction.
- A Maxwell–Boltzmann distribution curve shows the kinetic energy distribution of molecules and can be used to determine the proportion of molecules that have enough energy to overcome the activation energy barrier. It can be used to show the effect of increasing the temperature or adding a catalyst to a system.

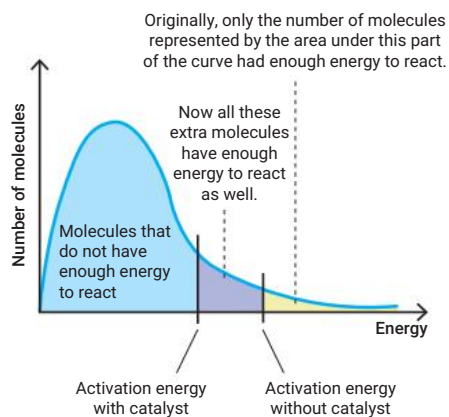
Enthalpy diagrams of exothermic and endothermic reactions



Maxwell–Boltzmann distribution curves at different temperatures

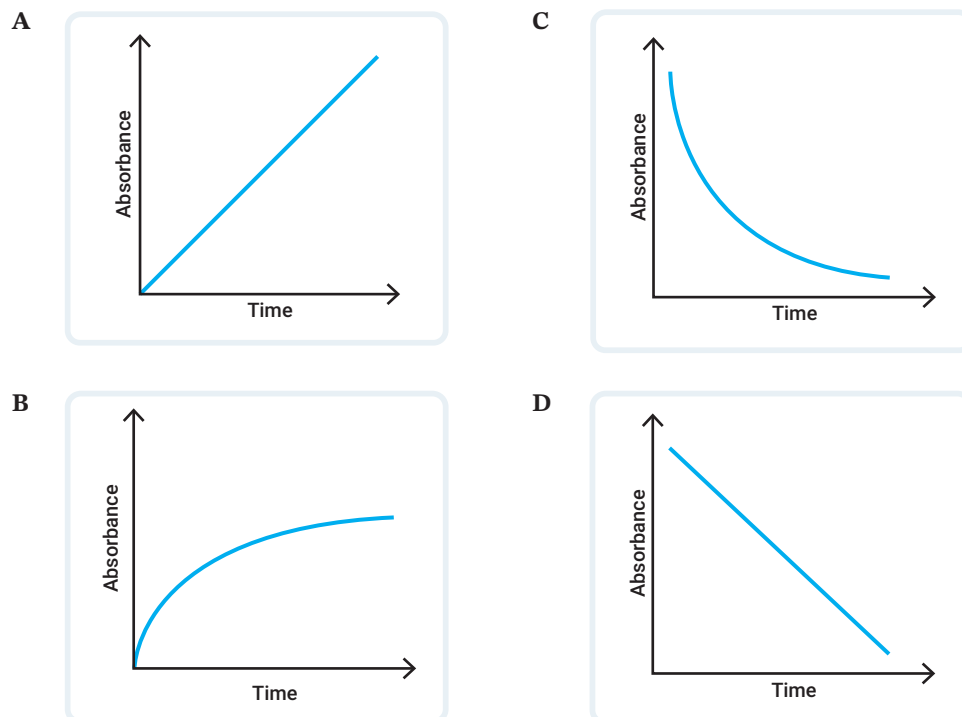


Maxwell–Boltzmann distribution curves with and without catalysts

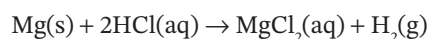


MULTIPLE CHOICE

1. Bromine liquid has a red-brown colour and reacts with methanoic acid to form a colourless substance. The decrease in colour intensity can be readily measured by a spectrometer. Which graph most accurately represents the changes in colour intensity as the reaction proceeds?



2. Magnesium metal reacts with dilute hydrochloric acid to produce a solution of magnesium chloride and hydrogen gas according to the following equation:



Which of the following steps would be least likely to increase the reaction rate?

- A** Using a suitable catalyst
B Increasing the gas pressure
C Increasing the concentration of HCl
D Using powdered magnesium rather than magnesium granules
3. An increase in temperature generally favours a higher reaction rate. Which of the following statements (I–IV) explains why this is so?
- I** Particles are moving faster, so they collide more frequently.
II Particles have more kinetic energy, so they collide with more force.
III Particles are more likely to collide with the correct orientation.
IV Particles collide more frequently with the sides of the container.
- A** I and IV
B I–IV
C I and II
D I–III

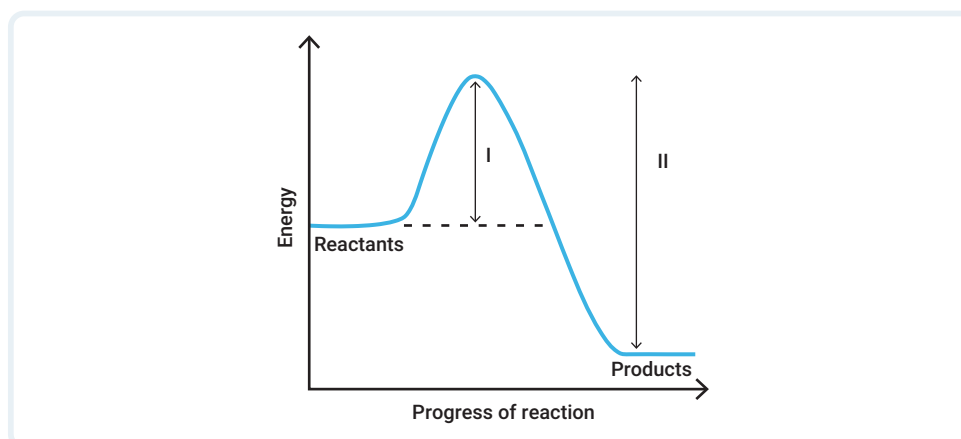
4. Which of the following statements is correct, with respect to an endothermic chemical reaction?
 - A The rate is always higher than for an exothermic reaction.
 - B The reaction causes the temperature of the surroundings to decrease.
 - C As the reaction proceeds, energy is released into the surrounding environment.
 - D The amount of energy required to break the bonds in the reactants is less than the amount of energy released as the products form.
5. A flour mill is especially prone to explosions unless suitable safety precautions are put in place. The best explanation for the increased possibility of explosions is that the fine particles of flour:
 - A have a high temperature, therefore a high kinetic energy.
 - B have a higher kinetic energy than the activation energy required for reaction.
 - C can collide with each other easily and produce a chain reaction.
 - D have an overall large surface area, causing combustion reactions to proceed rapidly.
6. Which one of the following procedures is likely to increase the rate of reaction of a gaseous system?
 - A Adding a negative catalyst to the system
 - B Increasing the volume of the system at constant temperature
 - C Increasing the pressure of the system at constant temperature
 - D Decreasing the temperature of the system at a constant volume
7. In two of the practical activities in this chapter, you used the following reaction to measure rate of reaction:



If the rate at which S(s) formed is $4.0 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$, then the rate at which:

- A HCl is consumed is $2.0 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$.
 - B $\text{Na}_2\text{S}_2\text{O}_3$ is consumed is $4.0 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$.
 - C H_2O is formed is $3.0 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$.
 - D NaCl is formed is $4.0 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$.
8. A spark plug in a car motor is necessary for the combustion of petrol to power the vehicle because:
 - A the activated complex will only form from a spark.
 - B the reaction is endothermic so needs an energy source to start it.
 - C the activation energy barrier must be overcome for the reaction to proceed.
 - D ΔH for the reaction is larger than the energy available in the surroundings.
 9. The activation energy of a reaction is usually:
 - A decreased by adding a catalyst.
 - B decreased by increasing the temperature.
 - C equal to the ΔH of the reaction.
 - D equal to the number of successful collisions.

10. The changes in energy during a chemical reaction are shown in the following diagram. I and II represent different energy changes.

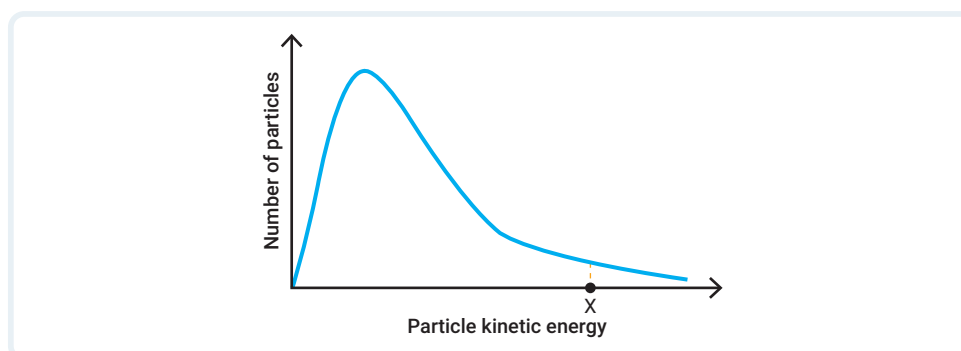


The effect of adding a catalyst to the system would be:

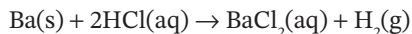
- A to decrease the size of II only.
- B to decrease the size of both I and II.
- C to increase the size of both I and II.
- D no change in either I or II.

SHORT RESPONSE

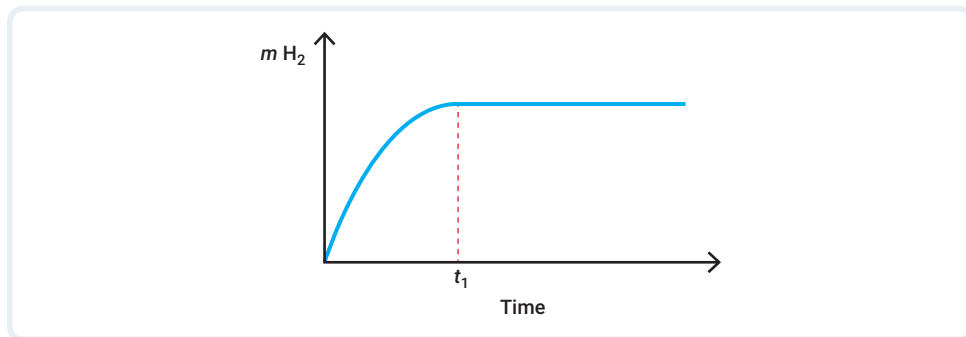
11. For a chemical reaction to proceed, particles must collide with sufficient energy and in the correct orientation. The following graph shows the range of average kinetic energies held by particles in a reaction mixture at a particular temperature.
- a Point X represents the activation energy for the reaction. **Explain** what this means.
 - b Redraw this graph on your own paper. On the axes, **sketch** a new graph to represent the range of particle energies at a higher temperature.
 - c Referring to the graph that you drew in part **b**, **explain** why even a small increase in temperature can result in a significant increase in reaction rate.



12. A series of experiments is conducted to investigate the reaction of barium with dilute hydrochloric acid. The equation that represents this reaction is:



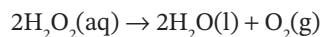
In the first experiment, 2.0 g of barium is dropped into a beaker containing 150 mL of 2.0 mol L^{-1} hydrochloric acid. The mass of hydrogen gas produced is measured at regular intervals and the results are shown on the following graph.



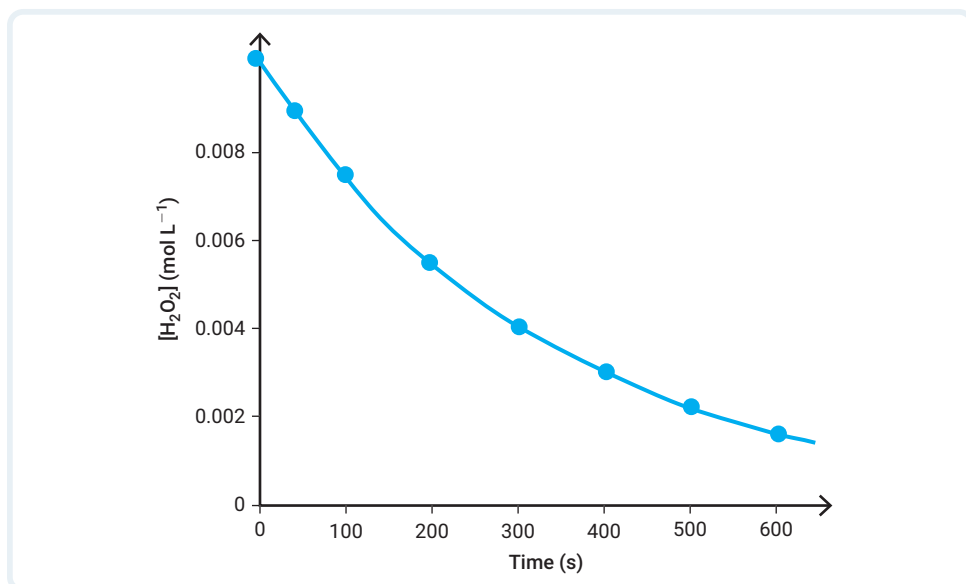
- Identify** the time when the reaction is proceeding at the greatest rate.
- Describe** what event occurs at time t_1 .
- The experiment is repeated using 1.0 g of powdered barium and the results are recorded. Copy the graph and **sketch** the results of the second experiment.
- Identify** two ways (other than the one used in this experiment) by which the reaction rate could be increased.

CROSS-CHAPTER QUESTION

13. Aqueous solutions of hydrogen peroxide decompose over time according to the equation:



The concentration of H_2O_2 changes with time during an experiment as shown by the following graph.



- a If the initial $[\text{H}_2\text{O}_2]$ was 0.010 mol L^{-1} and initial volume of the hydrogen peroxide solution was 200 mL, **calculate** the mass of oxygen released from the solution after 300 seconds.
- b **Calculate** the rate of reaction for the first 300 seconds using change in mass of oxygen.

DATA ANALYSIS

14. Analyse data

A series of experiments is undertaken to investigate the factors affecting the reaction rate of hydrochloric acid with marble, which is made from calcium carbonate (CaCO_3):

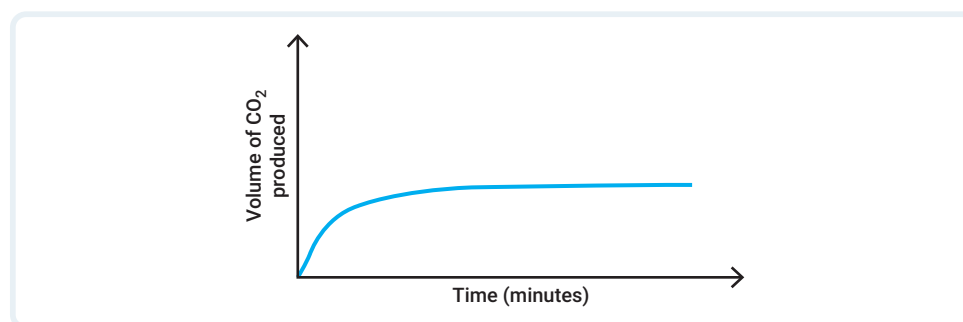


Flasks containing hydrochloric acid are set up under the conditions shown in Table 1.

TABLE 1 The reaction conditions for flasks A–D

Flask	Concentration of HCl (mol L^{-1})	Volume of HCl (mL)	Temperature of HCl ($^{\circ}\text{C}$)
A	0.15	50.0	20
B	0.20	50.0	20
C	0.20	50.0	30
D	0.15	100.0	20

Each experiment is initiated by adding 0.080 g of calcium carbonate. The volume of the carbon dioxide gas produced is measured at STP at set intervals. The result of the experiment for flask A is shown in the following graph.



- a Copy the graph and **sketch** the expected results for:
 - i flask B and label it 'B'
 - ii flask C and label it 'C'.

- b** To investigate the effect of the surface area of a solid reactant on the rate of the same reaction, a further three flasks are set up as shown in the following table. Each experiment is conducted as described except that the calcium carbonate is in the form shown in Table 2.

TABLE 2 The reaction conditions for flasks E–G

Flask	Concentration of HCl (mol L^{-1})	Volume of HCl (mL)	Temperature of HCl ($^{\circ}\text{C}$)	Form and mass of CaCO_3
E	0.15	100	20	Powder (0.080 g)
F	0.15	50	30	Single lump (0.080 g)
G	0.20	50	20	Powder (0.20 g)

Identify which flask in Table 2 provides the most informative results on the effect of surface area on the rate of reaction when analysed together with the results from Table 1. **Explain** your choice.

- c** The number of successful collisions between the reactant particles is less than 1% of the total number of collisions. **Explain** why the vast majority of collisions between reactant particles do not produce a reaction.
- 15.** In an experiment in which sodium hydrogen carbonate reacted with hydrochloric acid, the carbon dioxide gas (CO_2) was collected in a tube and the volume of gas produced over a period of time was recorded as shown below.

Time (s)	0	2	4	6	8	10	12	14	16	18
Volume of CO_2 (mL)	0	10	17	22	29	29	30	31	31.5	31.5

- a** Plot a graph showing the volume of CO_2 produced against time.
- b** From your graph, **state** the way in which the rate of the reaction between sodium hydrogen carbonate and hydrochloric acid varies with time. **Explain** why rate changes this way.
- c** **Identify** one way in which the rate of this reaction can be increased and **explain** why the increase occurs.

SCIENCE AS A HUMAN ENDEAVOUR

Syllabus dot point

- Appreciate that catalysts work in a variety of ways, and knowledge of the structure of enzyme molecules helps scientists to explain and predict how they are able to lower the activation energy for reactions.

Chemistry 2025 v1.1 General Senior Syllabus © QCAA 2024

Drivers of chemical reactions

Catalysts play a crucial role in various industries and natural processes by accelerating chemical reactions. By promoting more efficient reactions, catalysts contribute to greener and more sustainable practices.

Catalysts increase the rate of reactions by providing alternative pathways for reactions to occur that require a lower activation energy. They do this by:

- helping to orient the reactant particles in a way that increases the chances of successful collisions
- interacting with reactants to form intermediate compounds that require less energy to form the product
- helping to weaken bonds in the reactants so that less energy is required to break them and form products.

Consider the function of palladium–charcoal catalysts in producing alkanes from reactions involving alkenes and hydrogen. The palladium–charcoal catalyst helps to weaken bonds within each molecule and position reactants in a way that favours the production of products (**Figure 1**). By doing so, the catalyst decreases the activation energy of the overall reaction. (Me represents a CH_3 (methyl) group.)

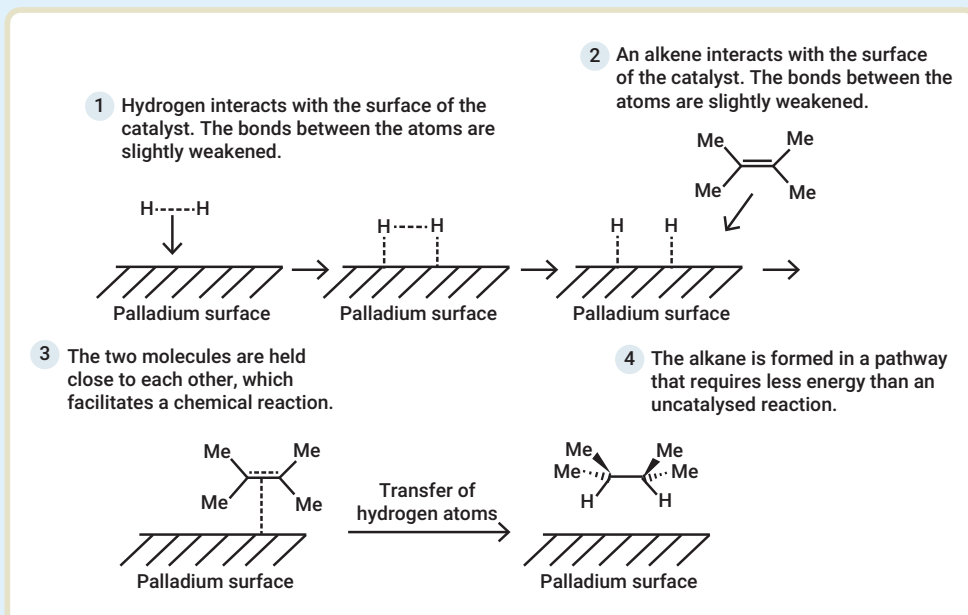


FIGURE 1 How palladium-charcoal catalysts function to catalyse reactions between alkenes and hydrogen to produce alkanes

Most catalysts used in chemistry are inorganic compounds, metals or metal oxides. However, there is a category of biological catalysts that are important in sustaining life.

Enzymes

Enzymes are biological catalysts, derived predominately from proteins. Enzymes interact with reactant molecules (substrates) at their active site. Owing to the uniqueness of their

active sites, enzymes are highly specific to particular substrates. The interaction between the enzyme and substrate forms the enzyme–substrate complex (Figure 2).

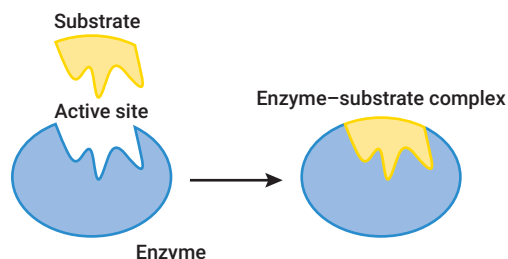


FIGURE 2 The interaction between an enzyme's active site and substrate to form the enzyme–substrate complex

Currently, there are two models to describe how enzymes function. In the lock-and-key model the substrate fits perfectly in the active site (Figure 3a). In the induced-fit model, it is believed that there is a slight change in the conformation in both the active site and substrate upon binding (Figure 3b).

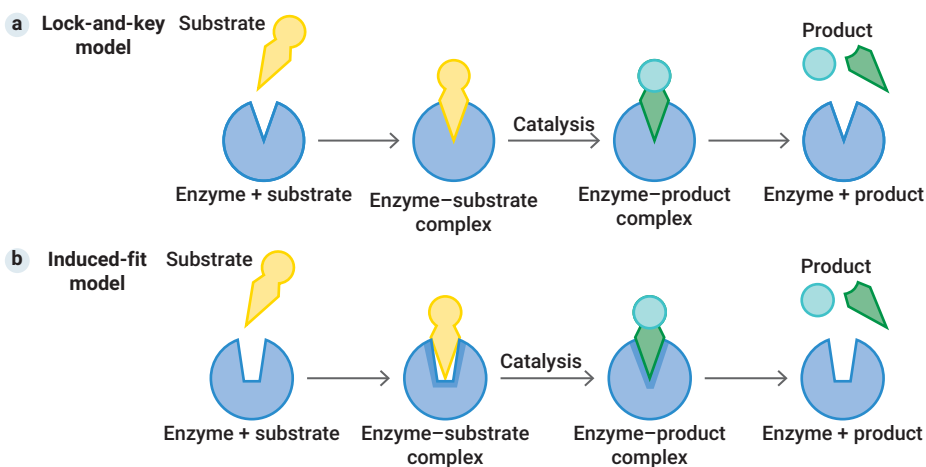


FIGURE 3 (a) The lock-and-key model and (b) the induced-fit model of enzyme function

The structure of the enzyme can help to determine how the enzyme functions. For example, different types of amino acids in the active site result in different interactions with the substrate. The amino acids could:

- form temporary covalent bonds with substrates to assist with bond making and breaking as well as stabilising the transition state
- donate or accept protons to optimise the pH environment for a reaction
- stabilise charges occurring in the transition state.

Understanding the structure of enzymes and their substrates is important for understanding how these reactions occur. Sir John Cornforth, a Sydney-born chemist was awarded the Nobel Prize in 1975 for identifying how the shape of enzymes affects the way they catalyse reactions. His work also helped show the processes involved in enzyme-catalysed reactions.

Understanding how catalysts and enzymes function is crucial for scientists, not only as a way to understand chemical reactions, but also to help determine ways to make reactions more efficient and effective.

ANSWERS

CHAPTER 1 ATOMIC STRUCTURE

LEARNING CHECK 1.1

DESCRIBING

- 1 Proton: positively charged; electron: negatively charged; neutron: uncharged (neutral)
- 2 Protons and neutrons are in the nucleus. Electrons are in an electron cloud surrounding the nucleus.

APPLYING

- 3 The atom is modelled as a positively charged nucleus surrounded by negatively charged electrons. The electrostatic attraction between the oppositely charged particles holds the atom together.
- 4 Only a tiny proportion of the alpha particles were deflected by the nuclei, so the nucleus must have been very small.
The alpha particles were large particles and a few were deflected back on their original path. The nuclei must have been very dense for this to occur.
The positively charged alpha particles that were passing close to the nucleus were still deflected. Therefore, the nuclei must also have been positively charged in order for this to occur.
- 5 Both models proposed the atom has a centre of positive charge; however, they differed in electron arrangement. Nagaoka's model proposed the electrons move in a fixed two-dimensional orbit around the nucleus (like the rings of Saturn) while Rutherford's model proposed the electrons were located in an electron cloud and moved randomly around the nucleus.

LEARNING CHECK 1.2

DESCRIBING

- 1 An element is a substance made up from one type of atom only
- 2 An element is represented by a unique symbol so that different elements can be distinguished. The atoms that make up an element have a specific number of protons, which is different for each element.
- 3 A group is a vertical column of elements and a period is a horizontal row of elements.
- 4 A neutral atom has no overall charge because the number of protons and electrons are equal, whereas an ion is a charged atom produced by the loss or gain of electrons.

APPLYING

- 5 Atomic number is the number of protons in the nucleus of an individual atom, whereas mass number is the number of protons and neutrons in the nucleus of an individual atom, which equates to the mass of an atom, because the mass of an electron is negligible in comparison with that of a proton or neutron.

- 6 Number of protons = atomic number. Number of electrons = atomic number (provided the atom is neutral and is not an ion). Therefore, number of neutrons = mass number – atomic number.

- 7 **a** Fluorine: 9 protons, 10 neutrons and 9 electrons
b Bromine: 35 protons, 45 neutrons and 35 electrons

- 8 **a** $^{27}_{13}\text{Al}$ **b** ^9_4Be

- 9 Answers are highlighted in bold italics in the table below.

Element	Atomic number	Mass number	No. of protons	No. of neutrons	No. of electrons
Hydrogen	1	1	1	0	1
Magnesium	12	24	12	12	12
Boron	5	11	5	6	5
Chlorine	17	35	17	18	17
Nickel	28	59	28	31	28

- 10 Potassium has an atomic number of 19, which means it has 19 protons. As this atom has only 18 electrons, it must be an ion with a +1 charge.

LEARNING CHECK 1.3

DESCRIBING

- 1 1st level: 2. 2nd level: 8. 3rd level: 18
- 2 The electron configuration shows the specific arrangement of electrons in the energy levels around the nucleus of an atom, which is unique to each element.
- 3 The Aufbau principle states that the electrons fill from the lowest energy level first and, once a level is at capacity, they enter the next lowest energy level, therefore showing the energy levels in which electrons will be located for the atoms and which is represented by the electron configuration.
- 4 The s orbital has the lowest energy, followed by the p orbital, which has higher energy than the s orbital but lower energy than the d orbital, which has the highest energy.
- 5 The different blocks of the periodic table correspond to the s, p, d and f subshells. Groups 1 and 2 are the s block, groups 13–18 are the p block, groups 3–12 are the d block, and elements 57–71 and 58–103 are the f block. The positioning of the different blocks also follows the order in which the subshells are filled (see Figure 1.3.3, page 17).

APPLYING

- 6 **a** **i** $1s^2 2s^2 2p^3$
ii $1s^2 2s^2 2p^6 3s^2 3p^6$
iii $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
iv $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
v $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

b i [He] $2s^2 2p^3$

iii [Ar] $4s^2 3d^1$



7 a Mg

b He

c Fe

8 a Sodium: Incorrect because there's a maximum of 8 electrons in the second shell, so configuration should be $1s^2 2s^2 2p^6 3s^1$.

b Calcium: 2 electrons enter the fourth shell before the 3p is complete, so configuration should be $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$.

9 Copper has the configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$, rather than the expected $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$. This is because the energy level of the completely filled 3d subshell is below that of the 4s, so the 4s subshell is left half-filled. For the other transition metals, with an unfilled 3d subshell, the 4s subshell fills first.

ANALYSING

10 The Aufbau principle of filling energy levels and sublevels from lowest to highest has been applied by filling energy levels 1 and 2, then 3s, 3p (as seen by [Ar]), then 4s before 3d. Pauli exclusion of no more than two electrons per orbital has been applied by the use of up and down arrows with no more than two arrows per orbital. Hund's rule has been applied by having only one electron in each of the last three orbitals rather than having two in one and none in the last.

CHAPTER EXAM

MULTIPLE CHOICE

1 D 3 C 5 D 7 D 9 A

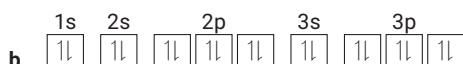
2 C 4 C 6 B 8 D 10 A

SHORT RESPONSE

- 11 a The group can be used to identify whether an atom is in the s, p, d or f block and so which outer orbitals are being filled.
- b The period can be used to determine which energy level is being filled, remembering that for the d block the energy level number is one less than the period, and for the f block the energy level number is two less than the period.

12 a Neon b Calcium c Sulfur d Chlorine

13 a $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$



DATA ANALYSIS

14 a i B

ii D

- b The nucleus is positively charged, very small and very dense.
- c Electrons exist in energy levels surrounding the nucleus. They cannot exist between the levels but can transition from one level to another if sufficient energy is supplied. Within each energy level, there are one or more subshells that contain one or more atomic orbitals.

15 $^{15}_7\text{N}$

CHAPTER 2 ISOTOPES

LEARNING CHECK 2.1

DESCRIBING

- Two atoms of the same element with the same number of protons but with different masses, resulting from their different number of neutrons
- Isotopes can be represented using four different formats:
 - ^A_ZM format, which gives atomic number, mass number and element symbol, e.g. $^{234}_{92}\text{U}$
 - Element name and mass number, e.g. uranium-234
 - Element symbol and mass number, e.g. U-234
 - ^AX format, which gives mass number and element symbol, e.g. ^{234}U

APPLYING

- Both isotopes have seven protons and seven electrons, but nitrogen-14 has seven neutrons, whereas nitrogen-13 has six neutrons.
- Isotopes of the same element have the same electron configuration, so they take part in chemical reactions in the same way.
- a A neutron needs to be added to the nucleus of Co-59.

b The Co-60 and Co-59 isotopes have different atomic masses and, because Co-60 is used for irradiating cancers, it must be radioactive, whereas Co-59 is not radioactive because it is not used for this purpose.

LEARNING CHECK 2.2

DESCRIBING

- The average mass of one atom of an element, relative to the mass of carbon-12
- First, the masses of each naturally occurring isotope are multiplied by their abundances. Then, the sum of these products is divided by the total of the abundances (usually 100).

APPLYING

- 3 RAM = $3 \times 12 = 36$
- 4 $\text{RAM} = \frac{(63 \times 69.17) + (65 \times 30.83)}{100} = 63.6$
- 5 Vanadium-51 is more abundant, because the average is closer to 51 than to 50.
- 6 Relative atomic mass of V = 50.94
- $$50.94 = \frac{(x \times 50) + ((100 - x) \times 51)}{100}$$
- $$5094 = 50x + 5100 - 51x$$
- $$x = 5100 - 5094$$
- $$= 6$$
- Percentage abundance of $^{50}\text{V} = x = 6.0\%$
- Percentage abundance of $^{51}\text{V} = (100 - x) = 94.0\%$

CHAPTER EXAM

MULTIPLE CHOICE

- 1 C 3 A 5 C 7 B 9 C
- 2 B 4 A 6 C 8 D 10 B

SHORT RESPONSE

- 11 a Two atoms of an element with the same number of protons but different numbers of neutrons
- b Isotopes have the same chemical properties because the atoms have the same electron configuration, but they have different physical properties, such as atomic mass, melting point and radioactivity. Elements are defined by the number of protons in the nucleus, the atomic number. However, atoms of the same element may have a different number of neutrons, thus forming an isotope of that element with some different physical properties.
- 12 Currently, there are 118 elements that have been discovered by scientists. However, elements can have isotopes, which are atoms of the same element with the same number of protons and different neutrons. Because of this, there are many more different types of atoms compared to the number of elements.

CROSS-CHAPTER QUESTION

- 13 a Five
- b Isotopes of the same element include:
- $^{31}_{15}\text{M}$ and $^{33}_{15}\text{M}$
 - $^{51}_{23}\text{M}$ and $^{52}_{23}\text{M}^+$
- c $^{31}_{15}\text{M}$ and $^{33}_{15}\text{M}$
- $^{127}_{52}\text{M}^-$ and $^{127}_{53}\text{M}$
- d $^{53}_{23}\text{M}^+$ and $^{58}_{28}\text{M}$ both have 30 neutrons

DATA ANALYSIS

- 14 $\text{RAM of Si} = \frac{(92.23 \times 28) + (4.68 \times 29) + (3.09 \times 30)}{100}$
- $$= \frac{2582.44 + 135.72 + 92.7}{100}$$
- $$= 28.11$$
- 15 RAM of Br = 79.90
- $$79.90 = \frac{(x \times 79) + ((100 - x) \times 81)}{100}$$
- $$7990 = 79x + 8100 - 81x$$
- $$= -2x + 8100$$
- $$-110 = -2x$$
- $$x = 55.0$$
- Percentage abundance $^{79}\text{Br} = 55.0\%$
- Percentage abundance $^{81}\text{Br} = 100 - 55.0 = 45.0\%$

CHAPTER 3 THE PERIODIC TABLE AND TRENDS

LEARNING CHECK 3.1

DESCRIBING

- 1 Mendeleev originally listed the elements in order of atomic mass and grouped them into rows and columns according to their properties. In the contemporary periodic table, the elements are now listed in order of atomic number.
- 2
- | Group | Common name |
|-------|-----------------------|
| 1 | Alkali metals |
| 2 | Alkaline earth metals |
| 17 | Halogens |
| 18 | Noble gases |
- 3 The elements in a group have similar properties, so knowing about the properties of the group allows predictions to be made about the behaviour of elements in chemical reactions.

APPLYING

- 4 Any two of fluorine, bromine, iodine and astatine
- 5 Magnesium would react with chlorine to form magnesium chloride. This is because beryllium is in group 2, like magnesium, and chlorine is in group 17, like fluorine. Therefore, the reactions would be very similar. Magnesium is below beryllium in group 2, so it is more reactive. Chlorine is lower than fluorine group 17, so it is less reactive. Therefore, it is not initially clear whether the overall reaction would be more or less vigorous.
- 6 Helium, neon, krypton, xenon and radon
- 7 Germanium

LEARNING CHECK 3.2

DESCRIBING

- 1 a 2 b 7 c 5 d 1 e 8
2 The energy required to remove a single electron from a neutral gaseous atom

APPLYING

- 3 a 1 b 4 c 1
4 Lithium, because atomic radius decreases across the period
5 Valency is the number bonds a single atom can form when forming a compound. Hydrogen has one electron so has a valency of 1. The valency can be determined from the electron configuration because it is equal to the number of electrons lost or gained by an atom to attain the electron configuration of the nearest noble gas.
6 a Oxygen b Beryllium c Fluorine d Carbon
7 Element with atomic number 17 is in group 17, 18 is in group 18, 19 is in group 1 and 20 is in group 2. Group 18 is a noble gas so it would have the highest ionisation energy, 1527 kJ mol^{-1} . Group 17 would have the next highest ionisation energy, 1257 kJ mol^{-1} as its outer shell is almost filled. Group 1 would have the lowest ionisation energy, 425 kJ mol^{-1} , as electrons are in the next energy level and further from the nucleus, so group 2 would have an ionisation energy of 596 kJ mol^{-1} . It has one more proton than group 1 and also a filled s subshell.
8 Across a period, there are more protons in the nucleus, and electrons are filling the same energy level, so attraction between nucleus and electrons is greater. Thus, as you move from left to right across the period, the atomic radius decreases. Down a group, there are more protons in the nucleus; however, electrons are in energy levels further from the nucleus, so the attraction between nucleus and outer electrons decreases so the atomic radius increases.
9 Although chlorine has more protons in the nucleus and is further to the right on the periodic table than oxygen, oxygen has one fewer energy levels than chlorine so its attraction for the outer electrons is stronger, thus its electronegativity is greater.

ANALYSING

- 10 a Element P is in group 1 and element Q is in group 14.
b For P, there is the largest relative increase in ionisation energy between first and second values, which means it must have one electron in its outer energy level. For Q, there is a significant increase in ionisation energy between the fourth and fifth values, so it must have four electrons in its outer energy level.

LEARNING CHECK 3.3

DESCRIBING

- 1 An element that has some of the properties of both a metal and a non-metal
2 Metallic character decreases from left to right across the periodic table and increases from top to bottom of the periodic table.
3 They react with water to form an alkaline solution.
4 It does not conduct electricity, it does not conduct heat, and it is not shiny.
5 Less vigorous: the reactivity of group 1 elements increases down the group, as the ionisation energy decreases and electrons can be more easily removed.

APPLYING

- 6 They all have one electron in their outer energy level and so lose that electron to form a positive ion, which has a noble gas electron configuration.
7 Electron affinity is the ability of an atom (in the gaseous state) to fully accept an electron and form a negative ion, whereas electronegativity is the ability of an atom to attract electrons from other atoms. The periodic trends in electron affinity and electronegativity are the same.
8 Fluorine is in group 17 of the periodic table just before the noble gas neon, which means it has a large number of protons in the nucleus and has only two energy levels containing electrons. Therefore, there is a very strong attraction of electrons from other elements.
9 Ionisation energy decreases down the periodic table. Group 1 elements react by losing their single valence electron and so this is done more easily, where outer energy level electrons are further from the nucleus. Thus, group 1 reactivity increases down the table.
However, group 17 reactivity decreases down the table, the reverse trend to ionisation energy. This is because the halogens react by gaining an additional electron to attain noble gas configuration rather than losing one.
10 Magnesium is less reactive than sodium. It is further to the right in the periodic table, thus is more electronegative, with a higher ionisation energy. Therefore, it requires more energy to lose an electron, which means it reacts more slowly than sodium.
11 Tennessine would have a melting point and boiling point greater than those of iodine and astatine – a boiling point of possibly 300°C . It will therefore be a solid at room temperature and is unlikely to vaporise. It is likely to be toxic. It may well be coloured and may possibly be a metalloid, with some slightly conductive properties, as the metallic character increases down group 17.

LEARNING CHECK 3.4

DESCRIBING

- 1 An amphoteric oxide is one that reacts with both acid and alkaline solutions
- 2 As we move from left to right, the oxides become more acidic and less alkaline.
- 3 It reacts with acid solutions to form water as one of the products. It does not react with alkaline solutions and dissolves sparingly in water to produce an alkaline solution.

APPLYING

- 4 a Alkaline b Acidic c Acidic
- 5 There is a greater proportion of alkaline oxides in period 4 than in period 3 because there are more metals – the transition metals are included in period 4. Therefore, the assumption is that they would form alkaline oxides.

CHAPTER EXAM

MULTIPLE CHOICE

1 C	3 B	5 C	7 C	9 B
2 D	4 D	6 A	8 A	10 C

SHORT RESPONSE

- 11 There is a greater attraction of outer electrons to the nucleus as number of protons increases, and no additional shielding of new electron shells, therefore outer electrons are harder to remove.

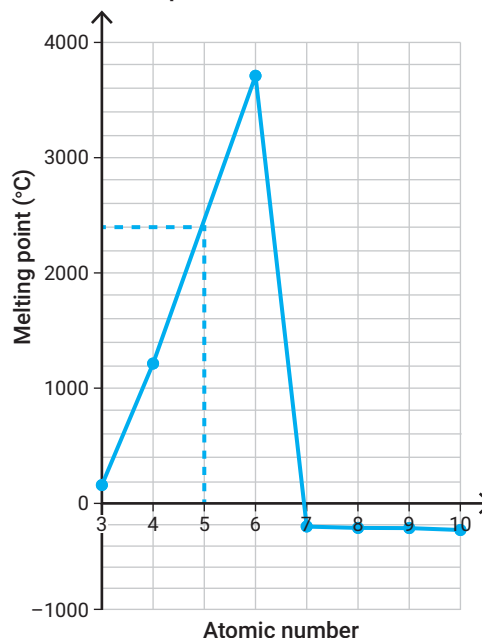
12	Radius of the K atom	>	Radius of the K ⁺ ion
	Number of valence electrons in a Ca atom	=	Number of valence electrons in a Ba atom
	Radius of an oxygen atom	<	Radius of a sulfur atom

CROSS-CHAPTER QUESTION

- 13 Both the alkali metals and halogens are highly reactive groups. The reactivity of the alkali metals increases as you go down the group, while the reactivity of the halogens decreases as you go down the group. This is because when the alkali metals react, they lose one electron from their outer energy level, which has the configuration s^1 , thus attaining the electron configuration of the adjacent noble gas. As you move down the group, the outer energy levels are further from the nucleus so the outer electron is more easily removed. The opposite occurs for the halogens, which have an outer electron configuration of p^5 so react to gain one electron to attain noble gas electron configuration. Electrons in lower energy levels are more attracted to the nucleus so the smaller elements in this group are more reactive.

DATA ANALYSIS

14 a Graph of melting point vs atomic number for period 2 elements



- i Boron has an approximate melting point 2400°C.
ii It has been assumed that melting point changes in a linear way from lithium to carbon.
 - c The data in both tables shows a consistent trend across the period. For both periods, melting point increases from group 1 to group 14, where it reaches a maximum, then decreases to group 18 (except for sulfur in group 16), where it reaches a minimum at group 18. The general trend is therefore melting point increases from group 1 to group 14, then decreases from group 14 to group 18.
- 15 a Fluorine, element number 9
- b
 - Relatively high nuclear charge given the small size of the atom, as it is to the right of the period
 - Only two energy levels of electrons with little shielding, as it is at the top of the periodic table
 - Strong attractiveness to outer electrons
 - c Data selected from three atoms in the same group, such as He, Ne, Ar, or Li, Na, K. There's an increased number of electron energy levels down the group. There's a decreased attraction of the nucleus to outer electrons due to greater shielding from the inner energy level electrons.
 - d The atoms with the lowest electronegativity are group 18 elements (noble gases). The next noble gas will be Kr. The difference in atomic number between the noble gases increases due to the d-block transition metals with the filling of the 3d subshell, which has energy between the 4s and 4p subshells.

CHAPTER 4 ANALYTICAL TECHNIQUES

LEARNING CHECK 4.1

DESCRIBING

- 1 A mass spectrum
- 2 They have different masses, so they are accelerated by different amounts by the electric field and so are moving at different speeds through the magnetic field. Therefore, they will be deflected by different amounts.

ANALYSING

- 3 a 90, 91, 92, 94, 96

b

Relative mass	Abundance
90	52
91	12
92	18
94	15
96	3

c
$$\text{RAM} = \frac{(90 \times 52) + (91 \times 12) + (92 \times 18) + (94 \times 15) + (96 \times 3)}{100}$$

$$= 91.3$$
 Element is zirconium.

LEARNING CHECK 4.2

DESCRIBING

- 1 Ground state is when all the electrons are in the lowest possible energy level. Excited state is when a higher energy level is filled and there is a gap in the lower energy levels.
- 2 When an atom is heated, the electrons will absorb a specific quanta of this extra energy and move to a higher energy level. An electron that has moved to a higher energy level will emit energy as it returns to the original lower energy level. The energy absorbed and emitted will be a specific wavelength, equal to the energy difference between the levels.
- 3 The Bohr model proposes that electrons are particles that orbit the nucleus in specific energy levels.
- 4 Bohr was able to quantitatively explain the emission spectrum of hydrogen. His calculations of the wavelengths for electron transitions between energy levels matched the values seen in the emission spectrum of hydrogen.
- 5 An atom has many different energy levels and, provided there is enough energy, electrons can move in different jumps between energy levels. Each line on the spectrum represents a different energy transition between different levels.

APPLYING

- 6 Any two elements in Table 4.2.1. Some elements produce flames of a specific colour because the wavelength of one of the electron's energy transitions corresponds to a colour in the visible region of the spectrum.
- 7 The electron configuration for every atom of the same element is identical, so the same wavelengths of light will be absorbed and emitted when heated.
- 8 Their electron configurations are different, so the energy levels are different and different, wavelengths of light will be emitted.

ANALYSING

- 9 The element is sodium – the emission lines match those for sodium in the standards supplied.

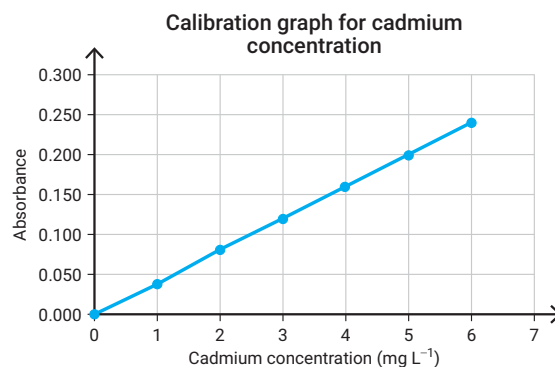
LEARNING CHECK 4.3

DESCRIBING

- 1 AAS measures the concentration of a particular metal ion. It does so by measuring a sample's absorption of light, as produced from a lamp that is made from the same metal being detected.
- 2 Measure the absorption of a series of known standard solutions. Plot a graph of absorption against concentration. Measure the absorption of the solution of unknown concentration and then use the graph to interpolate to find the concentration.
- 3 Absorption spectra show black lines on a coloured background because the electrons are absorbing that wavelength of light as they move to higher energy levels. Emission spectra show coloured lines on a black background because the electrons are emitting that wavelength of light as they move from higher to lower energy levels.
- 4 To ensure that only the required metal absorbs the light; although there may be a range of different metal ions in the sample, only the required metal will be measured. The lamp made from the element being tested can be excited and emit the specific wavelengths required.

ANALYSING

- 5 a



- b 2.7 mg L⁻¹

- 6 Atomic absorption enables scientists to determine the concentration of metal ions in sample solutions to a high level of accuracy. Examples can be used at student discretion.

CHAPTER EXAM

MULTIPLE CHOICE

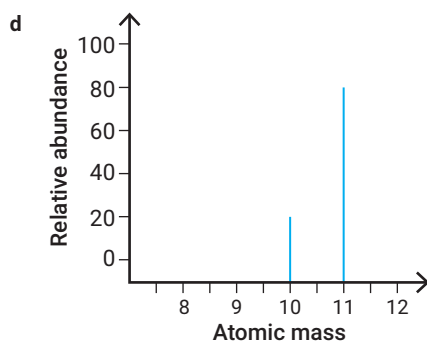
- 1 B 3 B 5 C 7 D 9 A
2 C 4 A 6 D 8 B 10 C

SHORT RESPONSE

- 11 a Helium – the wavelengths match those of the element in the sample.
b i The absorption spectrum would be a series of black lines on a coloured background because the wavelength of colour is being absorbed.
ii Although absorption and emission spectra are different in appearance, the wavelengths of the lines on the spectra will be the same as they both show the same transitions for electrons in the same element.
iii Mercury – the lines for the wavelengths provided are the same in both spectra.
- 12 Bohr calculated the energy values for the transitions of electrons in the hydrogen atom and these matched the values in the emission spectrum of hydrogen. He proposed electrons exist in specific energy levels around the nucleus and the energies are caused by electrons moving from higher to lower energy levels. The quantification of the energy values and match to the hydrogen spectrum led to the acceptance of his model.

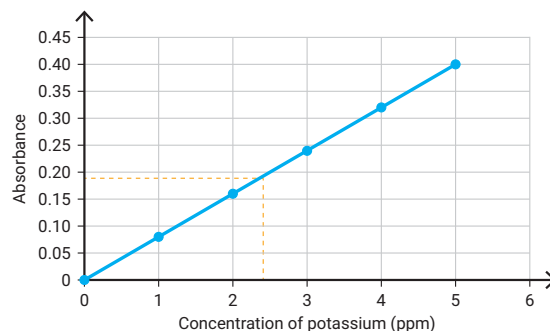
CROSS-CHAPTER QUESTION

- 13 a $1s^2 2s^2 2p^1$
b ${}^{11}_5\text{B}^+$
c $\text{RAM} = \frac{(20 \times 10) + (80 \times 11)}{100}$
 $= 10.8$

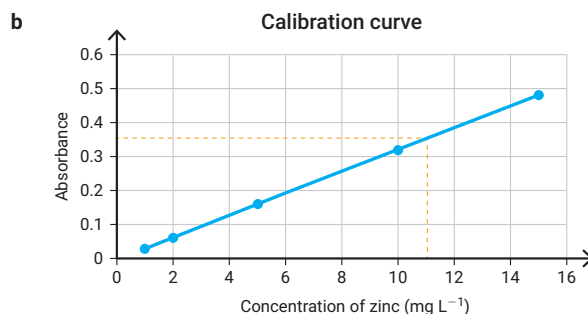


DATA ANALYSIS

- 14 a AAS can only measure the relative absorption of two different samples, so a comparison with different known standards is required.
b See orange line on graph – approximately 2.4 ppm of potassium



- c $\text{Concentration} = \frac{2.4 \times 100}{20} = 12 \text{ ppm}$
- 15 a The light emitted by the lamp would be the specific frequencies only emitted by zinc, so would not have been absorbed by any other metal.



- c See orange line on graph – 11 mg L^{-1}
d False – the concentration appears to be higher than 7.5 mg L^{-1} from this experiment.

CHAPTER 5 INTRODUCTION TO BONDING

LEARNING CHECK 5.1

DESCRIBING

- 1 Ionic, covalent and metallic bonding
- 2 Ionic bonding occurs between a metal and a non-metal. Covalent bonding occurs between two non-metals. Metallic bonding occurs between two metals.

APPLYING

- 3 Chemical bonds allow the the bonded atoms to form a lower energy, more chemically stable state with a full valence shell.

- 4 They have the same number of valence electrons.
- 5 a Metallic c Ionic e Ionic
b Covalent d Covalent

LEARNING CHECK 5.2

DESCRIBING

- 1 Electrostatic attraction between positive ions (atoms that have lost electrons) and negative ions (atoms that have gained electrons)
- 2 Zero: The total charge of the positive ions is exactly countered by the total charge of the negative ions.

APPLYING

- 3 Metals have a less than a half-full valence shell, so it is energetically favourable to lose electrons. Non-metals have a valence shell that is more than half-full, so it is energetically favourable to gain electrons.
- 4 a 
b 
c 
- 5 a 2+ b 3-
- 6 It has one valence electron in the first shell, so it can lose or gain an electron to form H^+ or H^- , respectively.
- 7 a Bromide Br^- d Potassium K^+
b Sulfide S^{2-} e Nitride N^{3-}
c Barium Ba^{2+}
- 8 There are no molecules formed, and the ionic bond is non-directional. Hence, an ionic compound is an array of positive and negative ions in fixed proportions. The formula gives the ratio of positive and negative ions so is an empirical formula.
- 9 a Potassium iodide d Lead(IV) oxide
b Barium chloride e Sodium phosphate
c Calcium hydride f Zinc nitrate
- 10 a Mg_3N_2 d NH_4OH
b Al_2O_3 e Cu_2CO_3
c HgS f $Zn_3(PO_4)_2$
- 11 a Sodium charge is 1+, bromide charge is 1-. Therefore, a ratio of 1:1 is required for an overall charge of 0.
b Aluminium charge is 3+, fluoride charge is 1-. Therefore, a ratio of 1:3 is required for an overall charge of 0.
- 12 a 3+ b 2+
13 a 1- b 3-

ANALYSING

- 14 a FeI_3
b $Ca_3(PO_4)_2$ or $Ba_3(PO_4)_2$

- c $CaSO_4$ or $BaSO_4$, CaS or BaS , LiI , $FePO_4$
d Ca^{2+} and S^{2-} or Ba^{2+} and I^-

LEARNING CHECK 5.3

DESCRIBING

- 1 They are loosely held, as the metals have low ionisation energies.
- 2 A series of positively charged ions arranged in a lattice, surrounded by delocalised electrons

APPLYING

- 3 They are attracted by the nuclei of the other atoms that are present and so are drawn into the space between the atoms, no longer being directly attached to any individual atom. The moving delocalised electrons allow electricity to flow through the metal, so a fork should not be put in a power socket.
- 4 a positive g valence
b ions h ionised
c delocalised i metal
d fixed j electrons
e free k attraction
f lattice

LEARNING CHECK 5.4

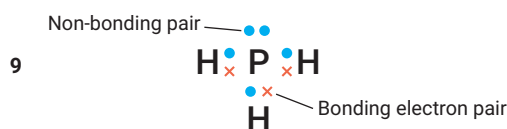
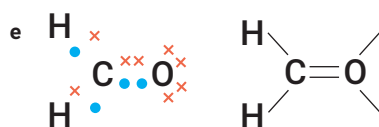
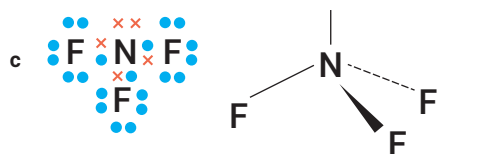
DESCRIBING

- 1 The element that is furthest to the left in the periodic table is always written first in a covalent compound.
- 2 a A pair of electrons shared between two atoms
b i 2 ii 4 iii 6
- 3 a 8 b 2 c 8
- 4 a OF_2 b BH_3
- 5 Carbon has 4 valence electrons, so shares one each with 4 hydrogen atoms, so that it has a total of 8 electrons in its valence shell.

APPLYING

- 6 a SO_2 c CCl_4 e $SiBr_4$
b N_2O_5 d NF_3
- 7 a Dinitrogen monoxide c Sulfur trioxide
b Nitrogen trichloride
d Dihydrogen monosulfide (also known as hydrogen sulfide and dihydrogen sulfide)
e Dinitrogen tetroxide





ANALYSING

10	Compound A	Compound B	Compound C
To which group in the periodic table does X belong?	15	14	16
Suggest an identity for X.	Nitrogen	Carbon	Oxygen
Suggest a possible molecule with this structure.	Ammonia NH ₃	Methane CH ₄	Water H ₂ O

CHAPTER EXAM

MULTIPLE CHOICE

- 1 D 3 B 5 D 7 C 9 B
2 D 4 B 6 B 8 D 10 C

SHORT RESPONSE

- 1 a Silver oxide d Ammonium bromide
b Dinitrogen trioxide e Magnesium hydroxide
c Sulfur tetrafluoride f Nitrogen monoxide
12 a BCl₃ c SiO₂ e K₂S
b CaCO₃ d N₂O₄ f PbO₂

CROSS-CHAPTER QUESTION

- 13 a Similarities: Magnesium is in both group 2 and period 3.
Differences: Group 2 elements include Be, Mg, Ca, Sr, Ba and Ra. Period 3 elements include Na, Mg, Al, Si, P, S, Cl and Ar. Group 2 elements have relatively low first ionisation energies. The first ionisation energy generally decreases down the group due to increasing atomic size and shielding effect. Period 3 elements show a more

varied trend in first ionisation energy. The general trend is an increase in ionisation energy from left to right.

- b Metals with lower first ionisation energies are more likely to take part in ionic bonding because they readily lose electrons to form cations.

DATA ANALYSIS

- 14 Sodium would most easily take part in bonding with chlorine as it has the lowest ionisation energy and only one valence electron.
15 a Q and R c PQ, P₂R, SO₂, SR
b P and S

CHAPTER 6 COMPOUNDS AND MIXTURES

LEARNING CHECK 6.1

DESCRIBING

- 1 a Pure substances don't have to be one element – they can also be compounds that contain different elements.
b Pure substances are made up of only one *type of substance*, which can be elements or compounds.
2 Mixtures can be separated by physical processes, whereas pure substances cannot.
Pure substances have a more uniform composition, whereas mixtures have more variable composition.

APPLYING

- 3 Due to the consistent properties of the substance across all samples, substance A could be classified as a pure substance.
4 Beach sand is an example of a mixture of various substances. As a result, it can be easily separated by physical techniques, using different properties such as particle size and solubility of components. In contrast, pure gold is an example of a pure substance. Since it is only made up of one element, there are only gold atoms, which can only be separated by turning the sample into a gas. This would require a large amount of energy.

LEARNING CHECK 6.2

DESCRIBING

- 1 In homogenous mixtures, the components are uniformly distributed, whereas in heterogenous mixtures, the components are not uniformly distributed.
2 A

APPLYING

- 3 Smoke is made up of many different substances that are not evenly distributed. Therefore, it would be considered a heterogenous mixture.

- 4 To separate the components of the mixture, you could perform:
 - filtration to separate the tea leaves from the liquid tea mixture
 - evaporation to separate the water from the sugar.
- 5 The beaker on the left, in which the light is able to pass through, is a homogenous mixture, whereas the beaker on the right, in which the light is clearly visible, is a heterogenous mixture.
- 6 Particle size can be used to separate out the sand by filtration, and boiling point can be used to separate water and dissolved magnesium chloride.

LEARNING CHECK 6.3

DESCRIBING

- 1 A pure substance has a constant composition; therefore, its properties would be consistent throughout. Mixtures are composed of two or more pure substances physically combined. Therefore, the properties of mixtures depend on the types of substances and the ratio of each.
- 2 A pure substance has a constant boiling point regardless of the quantity of the sample, whereas the boiling point of a mixture could vary, especially as some of the components evaporate.

APPLYING

- 3 You could melt both samples and record the melting temperatures. Pure water has a melting point of 0°C , whereas the sugar–water mixture would have a different melting point. You could place both samples in a glass of pure water. The pure water ice cube would float but the sugar–water cube would either sink or float lower because it would be more dense.

ANALYSING

- 4 a The melting point decreases as the percentage of gold decreases.
- b Since the melting point of the sample is different (lower) from the melting point of pure gold, this suggests that the sample is a mixture.

CHAPTER EXAM

MULTIPLE CHOICE

1 C	3 C	5 A	7 A	9 B
2 A	4 D	6 B	8 B	10 D

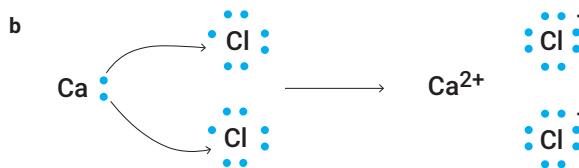
SHORT RESPONSE

- 11 When melted, sodium chloride consists only of sodium chloride but in liquid form. When dissolved in water, a mixture is formed that consists of both water and sodium chloride.
- 12 Copper by itself is a pure substance, and therefore would have a consistent melting point. Although it contains copper, brass also contains other metals. As a result, the

properties of brass would be affected by the composition of the different metals and would therefore have a different melting point from that of pure copper.

CROSS-CHAPTER QUESTION

- 13 a Ionic compound



- c When calcium chloride dissolves in water, the resultant particles are evenly distributed throughout the mixture.

DATA ANALYSIS

- 14 a As the percentage of ethanol decreases, the density increases.
- b The density of pure ethanol would be less than that of water. This is because the density increases as there is a greater water-to-ethanol ratio, suggesting that water is more dense than ethanol.
- 15 a The boiling point of butanol is 118°C ; the boiling point of ethanol is 78°C .
- b The boiling point of a 50:50 mix of butanol and ethanol is approximately 90°C . The graph shows that the boiling point changes as the composition of the mixture changes. This suggests that the boiling point of mixtures varies depending on the composition of the mixture, and is different from the boiling point of each of the pure substances found in the mixture.

CHAPTER 7 BONDING AND PROPERTIES

LEARNING CHECK 7.1

DESCRIBING

- 1 A type of bonding that involves the electrostatic forces of attraction between positively charged cations and negatively charged anions; results in an ionic compound with a three-dimensional lattice structure
- 2 Any three of high melting point, hard but brittle, does not conduct electricity in solid state, does conduct electricity in molten state

APPLYING

- 3 High melting point: There is strong attraction between ions, so a lot of energy is needed to move the ions out of their fixed position in the lattice.
Hard but brittle: Ions are held rigidly in place in the lattice, but applying stress brings ions of the same charge close together so the crystal can shatter.

Does not conduct electricity in solid state Electrical conduction depends on the presence of mobile charged particles, such as electrons or ions. In the solid state, the ions are fixed in the crystal lattice, so they are unable to move.

Does conduct electricity in the molten state: In the molten state, the ions are free to move so can transfer an electric charge.

- 4 Each ion in an ionic compound is surrounded by oppositely charged ions, so each ion can be attracted to another oppositely charged ion in any direction.
- 5 When an ionic compound is dissolved in water, the water molecules surround and separate the individual ions, breaking the ionic lattice. This process is called dissociation. Once dissociated, the ions are free to move independently in the solution. The ions act as mobile charge carriers, enabling the flow of electric current.

ANALYSING

- 6 The ionic radii of halogens increase down the group. The ionic radius of Na^+ is much smaller than that of Pb^{2+} ; because the Na^+ ions and Cl^- ions are closer to each other, the attraction between them is much stronger than the attraction of Pb^{2+} ions to I^- ions, which are further apart. This means the bonding in NaCl is much stronger than PbI_2 , so much more thermal energy is required to melt NaCl than PbI_2 .

LEARNING CHECK 7.2

DESCRIBING

- 1 Any three of metallic lustre, good conductors of heat and electricity, malleable/ductile
- 2
 - a Resistance to being scratched
 - b The ability of a substance to be drawn into a wire
 - c The ability of a substance to be hammered or pressed into shape
- 3 Any three of density, hardness, tensile strength, reactivity, boiling point and melting point

APPLYING

- 4
 - a Metals can be drawn out into a wire because the mobile delocalised valence electrons can move to offset the positive charge build-up caused by bringing metal ions close together.
 - b Metals are good conductors of heat, so a metal spoon that has been sitting in a cup of hot water will have easily conducted the heat from the water into itself and will be hot.
 - c Lightning is a discharge of electricity from the atmosphere to the ground. The umbrella frame, being metal, is a good conductor of electricity as the metallic lattice structure allows the delocalised electrons to move freely throughout the metal. This provides a pathway for the electric current to flow through the umbrella and into your body. This can cause serious injury or even death due to electric shock.

- 5
 - a Copper, being a metal, is a good conductor of electricity and can be used in electrical wiring.
 - b Aluminium is a particularly good conductor of heat so it can be used to make a saucepan where the efficient transfer of heat from the heat source to the food in the saucepan is required.
 - c Tungsten, being a metal, conducts electricity and its ductility means that it can be drawn into a thin wire. This is useful in a light filament because a very thin wire has a high electrical resistance, meaning that when electricity is passed through it, the wire becomes so hot that it glows.
 - d Iron is a particularly hard metal, so it is very difficult to deform the metal lattice, making it a very useful material in construction.
 - e Gold and silver exhibit a high lustre, meaning that the delocalised electrons at the surface of these metals reflect most of the light that hits their surface. The relative non-reactivity of gold and silver make them ideal as jewellery metals.

ANALYSING

- 6
 - a Similarity: They are all metals.
Differences: They all have a different atomic mass.
Trend: As atomic mass of the metal increases, thermal and electrical conductivities also increase.
 - b As atomic mass increases from sodium to copper, the number of delocalised electrons increases, meaning there are more delocalised electrons to conduct electricity and heat.
 - c As atomic mass increases, melting point increases. More importantly, as atomic mass increases, the number of delocalised electrons increases. This means that the metals that are higher in their respective groups or those that are to the left-hand side of the periodic table will have fewer delocalised electrons. The more delocalised electrons that are present, the more electrostatic attractions will be present, so the thermal energy required to melt the substance will be higher.

LEARNING CHECK 7.3

DESCRIBING

- 1 Covalent network solids: very high melting and boiling points, very hard brittle solids, chemically unreactive
Covalent molecular substances: low melting and boiling points, solids are generally soft, variable reactivity
- 2 Group 14 elements such as carbon and silicon are most likely to form giant covalent network substances because of their ability to form four covalent bonds, which means that they can form three-dimensional networks of covalently bonded atoms.

APPLYING

- 3 Covalent molecular substances exist as discrete molecules with intramolecular covalent bonding. The molecules are held together by weak intermolecular forces. This explains why covalent molecular substances have low melting points; the solids are usually soft and they exhibit variable chemical reactivity.

However, the atoms in giant covalent network substances form three-dimensional networks of covalently bonded atoms in which the covalent bonding extends indefinitely throughout the whole crystal. This explains the very high melting points of these substances as well as their hardness, and their non-reactivity.

- 4 In graphite, the carbon atoms are arranged in flat parallel layers. Within each layer, each carbon atom is covalently bonded to three other carbon atoms, which means only three of carbon's four electrons are involved in covalent bonding. The fourth electron is delocalised, which means it is free to move along the layer occupied by its 'parent' atom. This explains the ability of graphite to conduct electricity.

In diamond, each carbon atom is bonded to four other carbon atoms, so all the electrons are involved in forming covalent bonds. There are no free electrons to conduct electricity.

- 5 In graphite, the carbon atoms are arranged in flat parallel layers. Within each layer, each carbon atom is covalently bonded to three other carbon atoms, so each layer is a two-dimensional network of carbon atoms. Weak intermolecular forces hold the layers together. These sheets can easily be rubbed off, giving graphite a soft, slippery feel.

In diamond, each carbon atom is bonded to four other carbon atoms. It is very difficult to dislodge these atoms from the lattice. This makes diamond a very hard substance.

- 6 Similarities: Both graphite and fullerenes consist of carbon atoms covalently bonded to three other carbon atoms, while the fourth electron of each carbon atom is delocalised.

Differences: Unlike graphite, in which the carbon atoms form a hexagonal arrangement, the carbon atoms in fullerenes are arranged in hexagonal and pentagonal rings, which prevent them from forming flat sheets. As such, fullerenes, like graphite, are soft due to the weak forces of attraction between molecules but, unlike graphite, do not conduct electricity and have relatively low melting points.

ANALYSING

- 7 a C has the highest melting and boiling points.
b A, B and D would all be covalent molecular substances. Based on the melting and boiling points, the size of the molecules from smallest to largest would be A, B then D.
c Substance A
d Substance B

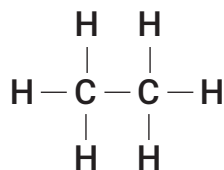
LEARNING CHECK 7.4

DESCRIBING

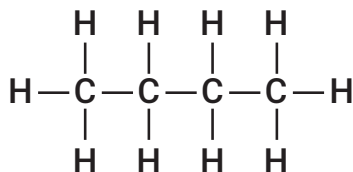
- 1 a C_nH_{2n+2} b C_nH_{2n} c C_6H_6
2 Saturated compounds contain carbon atoms that are bonded to other carbon atoms by single bonds only. Unsaturated compounds contain carbon atoms that are bonded to other carbon atoms by double or triple bonds.
3 Similarities: They are all types of hydrocarbons that consist only of hydrogen and carbon atoms.
Differences: Alkanes are saturated hydrocarbons, they contain only single covalent bonds between carbon atoms and they are relatively unreactive. Alkenes contain at least one double bond and are more reactive than alkanes. Benzene is an aromatic hydrocarbon with a ring structure consisting of six carbon atoms bonded together in a hexagonal ring with double bonds and is highly stable.
4 Benzene's stability is a result of the cyclic nature of the delocalised electrons within the benzene ring.

APPLYING

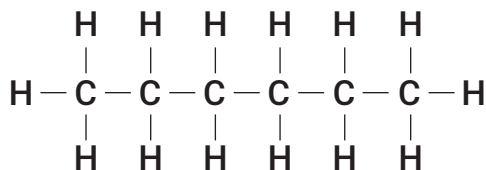
- 5 a C_6H_{12} is an alkene (hexene). Its formula matches that of an alkene C_nH_{2n}
b C_6H_{14} is an alkane (hexane). Its formula matches that of an alkane C_nH_{2n+2}
c C_9H_{18} is an alkene (nonene). Its formula matches that of an alkene C_nH_{2n}
6 a Ethane



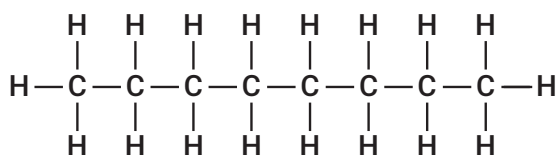
- b Butane



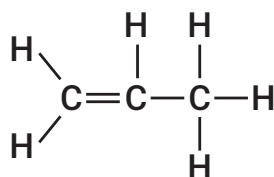
- c Hexane



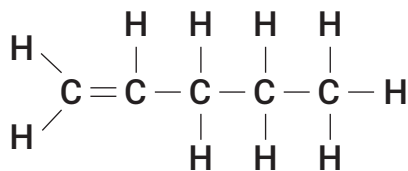
- d Octane



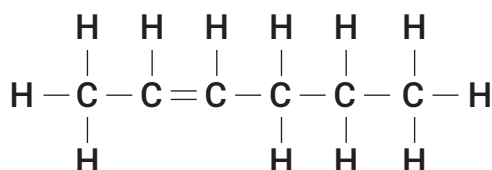
e 1-Propene



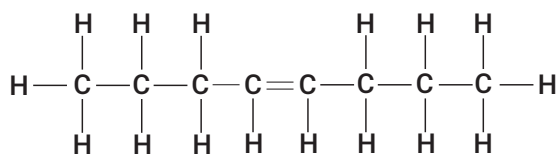
f 1-Pentene



g 2-Hexene



h 4-Octene



- 7 a Butane (alkane) e 1-Pentene (alkene)
 b Nonane (alkane) f 2-Butene (alkene)
 c Butane (alkane) g 2-Pentene (alkene)
 d Hexane (alkane)
- 8 a 5-Hexene should be named 1-Hexene. The carbon chain is numbered from the side that gives the C atom from which the double bond originates the smallest number
 b 4-Heptene should be 3-Heptene.

ANALYSING

- 9 Bottle A is water, because it is soluble in water, non-flammable and does not react with bromine water. Bottle B is hexane, because it is insoluble in water, flammable and reacts with bromine water. Bottle C is pentane, which is insoluble in water, flammable and has no reaction with bromine water. Bottle D is ethanol by a process of elimination.

LEARNING CHECK 7.5

APPLYING

- 1 Ionic bonding occurs between a metal and a non-metal. It involves the transfer of electrons from one atom (typically a metal) to another atom (typically a non-metal), resulting in the formation of oppositely charged ions. These ions are held together by electrostatic forces of attraction, forming an ionic compound.

Metallic bonding occurs between atoms of metal elements. It involves the delocalisation of valence electrons among a lattice of metal ions. These delocalised electrons form a 'sea' of mobile charge carriers that are free to move throughout the metal lattice, holding the positive metal ions together via electrostatic attraction.

- 2 Covalent molecules are formed when atoms share pairs of electrons to achieve a stable electron configuration. It occurs between non-metal atoms. Covalent molecules typically have low melting and boiling points and are often gases, liquids or soft solids at room temperature. They do not conduct electricity.

Giant covalent networks are three-dimensional networks of atoms bonded together by strong covalent bonds. Each atom forms single covalent bonds with multiple neighbouring atoms. They have very high melting and boiling points and are typically very hard. They do not conduct electricity except for graphite.

Allotropes of carbon are different forms of pure carbon that exist in various structural arrangements, each with unique properties. For example, fullerenes are arranged in spherical structures with unique properties.

ANALYSING

- 3 a i Solid X: delocalised electrons
 ii Molten XY_2 : cations (positive ions) and anions (negative ions)
- b XY_2 is an ionic compound. Ionic compounds have high melting points due to the strong electrostatic forces of attraction between the positively and negatively charged ions that make up the crystal lattice structure.
- c When Y_2 melts, covalent bonds are broken. Intermolecular forces holding the molecules together will also need to be overcome to allow it to melt.
- 4 a A: metallic; B: covalent network; C: covalent molecular; D: ionic; E: covalent network
- b Sodium chloride is an ionic compound composed of sodium and chloride ions arranged in a three-dimensional lattice structure. In the solid state, sodium chloride does not conduct electricity because the ions are held in fixed positions and there are no mobile charge carriers. When it is melted to form a molten liquid state, the ionic bonds are broken, and the ions become mobile and can carry an electric current.

Copper is a metallic element with a crystalline structure consisting of positively charged copper ions surrounded by a 'sea' of delocalised electrons. In the solid state, the delocalised electrons are free to move and will conduct electricity. When molten, the delocalised electrons remain present and mobile, still allowing for some flow of electric current as the lattice structure has broken down and the atoms become disordered.

- c Covalent molecular compounds have low melting points due to the weak intermolecular forces between molecules. They require less energy to break these forces, resulting in lower melting points. They are non-conductors of electricity as the electrons are not mobile and are shared between the atoms. There are no free-moving charged particles to carry electric current. Generally, covalent molecular compounds are soft, so will flatten with the 'hammer' test.

Giant covalent network compounds have high melting points, as the atoms are bonded together by strong covalent bonds, forming a network structure throughout the substance. Breaking these bonds requires a significant amount of energy. They are poor conductors of electricity as there are no charged particles to carry electric current. Giant covalent network compounds are hard and brittle, so will shatter when hit with a hammer.

CHAPTER EXAM

MULTIPLE CHOICE

- 1 D 3 A 5 C 7 C 9 B
2 C 4 B 6 C 8 D 10 A

SHORT RESPONSE

- 11 While phosphorus trichloride is a covalent molecular substance – low melting point, non-conductor – calcium chloride is an ionic substance, with a high melting point, and it conducts electricity when molten.
- 12 a It does not conduct electricity.
b The melting and boiling points are too low. The liquid does not conduct electricity.
c Covalent molecular compound because it has a low boiling point and does not conduct electricity.

CROSS-CHAPTER QUESTION

- 13 a F
b i EG_3 ii Covalent

DATA ANALYSIS

Group	Substance
Metallic	U, X
Ionic	V, Y
Covalent molecular	W
Covalent network	Z

Metallic substances: Substances U and X both have a lustrous silvery appearance, are insoluble in water and kerosene, and conduct electricity as both a solid and molten liquid.

Ionic substances: Substances V and Y are both crystals that are soluble in water and conduct electricity in aqueous solutions.

Covalent molecular substances: Substance W has waxy crystals and a low melting point, is soluble in kerosene and does not conduct electricity.

Covalent network substances: Substance Z has a very high melting point, is insoluble and does not conduct electricity.

CHAPTER 8 CHEMICAL REACTIONS

LEARNING CHECK 8.1

DESCRIBING

- 1 a The production of new substances
b The change of state of a substance between a solid, liquid or a gas
c Reactions that occur without the ongoing addition of energy

APPLYING

- 2 a Iron → oxygen → iron(III) oxide
Reactants: iron and oxygen; product: iron(III) oxide
b Octane → oxygen → carbon dioxide → water
Reactants: octane and oxygen;
products: carbon dioxide and water
- 3 There is a reaction occurring between water and baking soda as it produces bubbles of gas. This is an indicator that a chemical reaction has occurred.

LEARNING CHECK 8.2

DESCRIBING

- 1 Chemical equations are used to summarise information about chemical reactions. These equations communicate information concisely about reactants, products and numbers of atoms of each element present.
- 2 1 Write the unbalanced equation.
2 Count the number of atoms on each side of the equation.
3 Add coefficients to make the numbers of atoms equal on both sides (coefficient × subscript = number of atoms)
4 If necessary, reduce coefficients to lowest possible ratio (never change subscripts).
5 Update and double check atom balance until complete (remember the law of conservation of mass).
6 Add states for each of the reactants and products.

APPLYING

- 3 a $\text{Fe(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2\text{(aq)} + \text{H}_2\text{(g)}$
b $2\text{NaHCO}_3\text{(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + 2\text{CO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$
c $2\text{CuFeS}_2\text{(s)} + 2\text{SiO}_2\text{(s)} + 4\text{O}_2\text{(g)} \rightarrow \text{Cu}_2\text{S(s)} + 2\text{FeSiO}_3\text{(s)} + 3\text{SO}_2\text{(g)}$

- d $2\text{PbS(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{PbO(s)} + 2\text{SO}_2\text{(g)}$
- e $\text{C}_3\text{H}_8\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O(g)}$
- 4 a $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O(l)}$. Note that the water will be gaseous after the reaction until it condenses, so the equation could also be written with $2\text{H}_2\text{O(g)}$ as the product.
- b Balanced
- c $\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} \rightarrow 2\text{C}_2\text{H}_5\text{OH(l)} + 2\text{CO}_2\text{(g)}$
- d Balanced
- e Balanced
- 5 a $\text{C}_6\text{H}_{12}\text{(l)} + 9\text{O}_2\text{(g)} \rightarrow 6\text{CO}_2\text{(g)} + 6\text{H}_2\text{O(g)}$
- b $4\text{Fe(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Fe}_2\text{O}_3\text{(s)}$
- c $\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} + 6\text{O}_2\text{(g)} \rightarrow 6\text{CO}_2\text{(g)} + 6\text{H}_2\text{O(l)}$
- d $3\text{H}_2\text{(g)} + \text{N}_2\text{(g)} \rightarrow 2\text{NH}_3\text{(g)}$
- e $\text{Mg(OH)}_2\text{(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + 2\text{H}_2\text{O(l)}$
- f $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{ZnSO}_4\text{(aq)} + \text{H}_2\text{(g)}$

LEARNING CHECK 8.3

DESCRIBING

- 1 Use a pop test to test for hydrogen gas. Put a lit splint or burning match near the mouth of the test tube; if it pops, hydrogen gas is present. If oxygen gas is present, inserting a glowing splint into the mouth of the test tube will cause it to relight. If carbon dioxide gas is present, bubbling the gas into limewater will cause the limewater to go cloudy.

APPLYING

- 2 a Double-displacement reaction
- b Double-displacement reaction
- c Combustion reaction
- d Single-displacement reaction
- e Decomposition reaction
- f Combination reaction
- g Double-displacement reaction
- h Single-displacement reaction
- i Combustion reaction
- j Combination reaction
- k Single-displacement reaction
- l Double-displacement reaction
- m Decomposition reaction
- n Decomposition reaction
- o Double-displacement reaction, also an acid–base reaction
- p Combination reaction, also redox reaction

CHAPTER EXAM

MULTIPLE CHOICE

1 D	3 C	5 C	7 D	9 A
2 C	4 C	6 C	8 C	10 C

SHORT RESPONSE

- 11 The coefficients in the chemical equation should be the lowest ratio, therefore the balanced equation should be:
 $\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$
- 12 a $\text{N}_2\text{(g)} + 3\text{H}_2\text{(g)} \rightarrow 2\text{NH}_3\text{(g)}$
- b $2\text{FeSO}_4\text{(s)} \rightarrow \text{Fe}_2\text{O}_3\text{(s)} + \text{SO}_2\text{(g)} + \text{SO}_3\text{(g)}$
- c $\text{C}_3\text{H}_8\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O(g)}$
- d $\text{C}_2\text{H}_5\text{OH(l)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)} + 3\text{H}_2\text{O(g)}$
- e $6\text{CO}_2\text{(g)} + 6\text{H}_2\text{O(l)} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} + 6\text{O}_2\text{(g)}$
- f $2\text{Cu(NO}_3)_2\text{(s)} \rightarrow 2\text{CuO(s)} + 4\text{NO}_2\text{(g)} + \text{O}_2\text{(g)}$
- g $2\text{NaOH(aq)} + \text{MgCl}_2\text{(aq)} \rightarrow \text{Mg(OH)}_2\text{(s)} + 2\text{NaCl(aq)}$

CROSS-CHAPTER QUESTION

- 13 a Ionic compounds: calcium chloride, sodium bicarbonate, calcium carbonate, sodium chloride; covalent compound: water and carbon dioxide
- b Ionic compounds have high melting and boiling points, are good conductors of electricity in the molten and aqueous states, and are hard but brittle. Covalent compounds have low melting and boiling points, are non-conductors of electricity, and form solids that are generally soft.
- c $\text{CaCl}_2\text{(aq)} + 2\text{NaHCO}_3\text{(aq)} \rightarrow \text{CaCO}_3\text{(s)} + 2\text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$
- d Two clear aqueous solutions react together producing a solid. This formation of a solid is an indicator of a chemical reaction.
- e Double-displacement reaction or more specifically a precipitation reaction has occurred.

DATA ANALYSIS

- 14 Substance A is undergoing a phase change; going from a solid to a liquid. Substance B has taken part in a chemical reaction as a new odour was produced along with bubbles. These are both indicators of a chemical reaction.
- 15 a i Trial 3 of the zinc experiment is the anomaly.
- ii It may have been a smaller piece of zinc metal used in the reaction.
- b Magnesium: 10.3°C
 Zinc: 8.65°C
 Iron: 5.1°C
- c They could observe the reaction of each metal with water or oxygen gas.

CHAPTER 9 THE MOLE CONCEPT AND THE LAW OF CONSERVATION OF MASS

LEARNING CHECK 9.1

DESCRIBING

- 6.02×10^{23} – the number of atoms in exactly 12 g of the carbon-12 isotope
 - The amount of a substance containing N_A particles of that substance
 - One-twelfth of the mass of one atom of carbon-12

APPLYING

- The relationship between the number of moles (n) of a substance and the number of particles (atoms, ions or molecules) is given by:

$$\text{number of moles } (n) = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$
- 6.64 mol
 - 0.0171 mol
 - 0.2000 mol
 - 0.0137 mol
 - 55 mol
 - 0.0199 mol
 - 0.011 mol
 - 5.00 mol
 - 1.50 mol
- 1.5×10^2 molecules
 - 1.99×10^{23} atoms
 - 9.6×10^{23} ions
 - 2.1×10^{22} particles
 - 6×10^{22} molecules
 - 3.6×10^{24} atoms
 - 3×10^{23} ions
 - 4.9×10^{22} molecules
 - 1.26×10^{24} molecules
- 3 mol
 - 3 mol
 - 9 mol
 - 5.42×10^{24} atoms
- 498 mol
 - 3987 mol
 - 8970 mol
- 10.0 mol
 - 0.40 mol
 - 45.1 mol

LEARNING CHECK 9.2

DESCRIBING

- The relative molecular mass (atomic mass units) is calculated by adding the relative atomic masses of all the component atoms of the molecule. The mass of 1 mol of a substance is called the molar mass (M). The mass of 1 mol of a particular element or compound is numerically equal to its relative atomic, molecular or formula mass, but the units of molar mass are grams per mole (g/mol or g mol^{-1}).
- Empirical formula is the chemical formula showing the simplest ratio of elements in a compound, whereas the molecular formula specifies the total number of each of the atoms that are contained in the molecule.
- The percentage by mass of each of the different elements in the compound

APPLYING

- 123.88 amu
 - 44.01 amu
 - 63.02 amu
 - 180.18 amu
- 40.99 amu
 - 261.35 amu
 - 105.99 amu
 - 159.61 amu
- 34.02 g mol^{-1}
 - $184.11 \text{ g mol}^{-1}$
 - $100.09 \text{ g mol}^{-1}$
 - 72.17 g mol^{-1}
- 23.8 g
 - 1125 g
 - 646 g
 - 148.8 g
 - 95.9 g
- 0.5549 mol
 - 0.1665 mol
 - 0.3324 mol
 - 0.0480 mol
 - 0.0972 mol
- 74.83% C and 25.17% H
 - 39.34% Na and 60.66% Cl
 - 43.51% Ca, 26.07% C and 30.42% N
 - 22.57% N, 6.51% H, 19.35% C and 51.56% O
- COCl_2
 - C_2H_5
 - NO_2
 - Na_2PO_3
- OCl_2
- i CCl_4
 - ii $\text{C}_4\text{H}_{10}\text{O}$
 - iii $\text{C}_2\text{H}_5\text{NO}_2$
- $\text{C}_4\text{H}_{10}\text{O}$
- $\text{C}_5\text{H}_7\text{N}$
 - $\text{C}_{10}\text{H}_{14}\text{N}_2$
- 294.34 g
 - 0.0211 mol
 - 72.11 g
 - 8.75×10^{18} molecules
- 7.53×10^{22} molecules
 - 1.51×10^{23} atoms
- 3552 mol
 - 2.14×10^{27} water molecules

LEARNING CHECK 9.3

DESCRIBING

- The relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers
- Mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or physical transformations.
- Chemical reactions must obey the law of the conservation of mass. This means that, in chemical reactions, the combined mass of the products must equal the combined mass of the reactants. Atoms cannot be created or destroyed in a chemical reaction, so the number of individual atoms must be the same on both sides of a chemical equation (balanced in terms of the number of atoms).

APPLYING

4	Fe ₂ O ₃	CO	Fe	CO ₂
	1	3	2	3
	1.5	4.5	3	4.5
	0.2	0.6	0.4	0.6
	0.03	0.09	0.06	0.09

LEARNING CHECK 9.4

DESCRIBING

- The reactant that is totally consumed when the chemical reaction reaches completion
- The amount (mass) of a product that is produced from a complete reaction of a limiting reactant
- The experimental yield is usually less than the theoretical yield because:
 - competing reactions may not have been considered in the balanced equation we have written
 - the mechanism of the reaction may differ because the experimental conditions are different from what is required for the assumed reaction
 - the conditions required for the reaction to take place, such as temperature or pressure, may not be perfect
 - the reaction mixture may contain impurities, which influence the reaction.

APPLYING

- 4 a Oxygen b Ethanol c 1 : 2
- 5 a $2\text{C}_4\text{H}_{10}(\text{g}) + 13\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 10\text{H}_2\text{O}(\text{g})$
- b Butane d 18 g CO₂
- c $\frac{13}{2}$ or 6.5 mol O₂ e 9.3 g H₂O
- f 59%

ANALYSING

- 6 a $(\text{CH}_3)_2\text{NNH}_2(\text{l}) + 2\text{N}_2\text{O}_4(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g}) + 3\text{N}_2(\text{g})$
- b No. of moles of $(\text{CH}_3)_2\text{NNH}_2 = 2495$ mol
No. of moles of $\text{N}_2\text{O}_4 = 4999$ mol
From mole ratios of $(\text{CH}_3)_2\text{NNH}_2$ and $2\text{N}_2\text{O}_4$ (2495 : 4999 = 1 : 1.002), we determine that N_2O_4 is slightly in excess, so $(\text{CH}_3)_2\text{NNH}_2$ is the limiting reactant.
Yield of $\text{N}_2 = 3 \times 2496 \text{ mol} \times 28.02 \text{ g mol}^{-1}$
= 210 kg
- 7 a 10.24 g b 49.5 g
- 8 a i 14 g CaO ii 11 g CO₂
- b 75% yield of CaO
- 9 a i 5.42 g ZnCl₂ ii 0.080 g H₂
- b 96%
- 10 a $2\text{C}_8\text{H}_{18}(\text{l}) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{g})$
- b i 160 kg ii 141 kg

- c i 75.0% yield
- ii Hydrocarbon combustion involves competing reactions that lead to production of CO. These reactions, called 'incomplete combustion', reduce the CO₂ produced according to the balanced equation. It's also possible that some CO₂ or reactant was lost during the experiment, resulting in a lower calculated yield.

- 11 a $\text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) \rightarrow 2\text{C}_2\text{H}_5\text{OH}(\text{aq}) + 2\text{CO}_2(\text{g})$
- b 23 g
- c 60% experimental yield. It's possible that there were competing reactions involving glucose that reduced the amount of glucose available for fermentation. Also, the conditions for fermentation were not favourable for yeast, so it was not able to allow for 100% yield.
- 12 a i 2.25 mol O₂ ii 4.5 mol MgO
- b i 0.75 mol Fe iii 4 mol Fe
- ii 1.5 mol H₂
- c i 35.75 mol O₂ iii 27.5 mol H₂O
- ii 22 mol CO₂

CHAPTER EXAM

MULTIPLE CHOICE

- 1 D 3 C 5 D 7 C 9 D
- 2 B 4 C 6 A 8 C 10 C

SHORT RESPONSE

- 11 a 180.18 b 6 moles c 6.02×10^{23} molecules
- 12 81.3%

CROSS-CHAPTER QUESTION

- 13 a i Silver nitrate: AgNO₃; iron(III) chloride: FeCl₃; silver chloride: AgCl; iron(III) nitrate: Fe(NO₃)₃
- ii All of the compounds in this reaction are ionic.
- iii Ionic compounds have high melting points, will conduct electricity in the molten state and in solution, are hard but brittle, are mostly soluble in water and exist in a giant lattice of repeating ions.
- b $3\text{AgNO}_3(\text{aq}) + \text{FeCl}_3(\text{aq}) \rightarrow 3\text{AgCl}(\text{s}) + \text{Fe}(\text{NO}_3)_3(\text{aq})$
- c AgNO₃
- d 0.147 mol AgCl
- e 21.1 g AgCl
- f 37 g Fe(NO₃)₃

DATA ANALYSIS

- 14 a CuO
- b The results were inaccurate because the mass of Cu is 3.175 g (23.675 – 20.500 = 3.175 g), which is greater than

the mass of the compound, 2.975 g ($23.475 - 20.500 = 2.975$ g). Since the mass of the compound is less than the original copper, the results are inaccurate.

- 15 a** If there is too little magnesium for the amount of oxygen, it suggests that not all the magnesium has reacted to form magnesium oxide. This could be due to inadequate heating during the experiment. The presence of magnesium nitride can also contribute to a lower magnesium-to-oxygen ratio. The formation of magnesium nitride results in less magnesium available to form magnesium oxide. Another possibility is the loss of magnesium during the experiment, such as splattering during heating or spilling powder after heating.
- b** If there is too much magnesium for the amount of oxygen, it suggests that the amount of magnesium used exceeded the stoichiometric amount required to react with the available oxygen. This could be due to an error in measuring the amount of magnesium or inaccuracies in the experimental procedure. The presence of unreacted magnesium in the crucible can also contribute to a higher magnesium-to-oxygen ratio. Soot formation on the crucible due to use of the safety flame while heating can also lead to errors in measuring the final mass of magnesium oxide.

CHAPTER 10 EXOTHERMIC AND ENDOTHERMIC REACTIONS

LEARNING CHECK 10.1

DESCRIBING

- 1 a** A measure of the average kinetic energy of the particles that constitute a solid, liquid or gas
- b** Temperature ($^{\circ}\text{C}$) = temperature (K) $- 273$. Absolute zero is the lowest temperature that is theoretically possible, at which the motion of particles, which constitutes heat, is at a minimum. It equals -273.15°C .
- c** Kinetic energy is the energy possessed by moving particles.

APPLYING

- 2 a** 287 K **b** 262 K **c** -67°C
- 3** The graph representing gas X is at the lowest temperature.

LEARNING CHECK 10.2

DESCRIBING

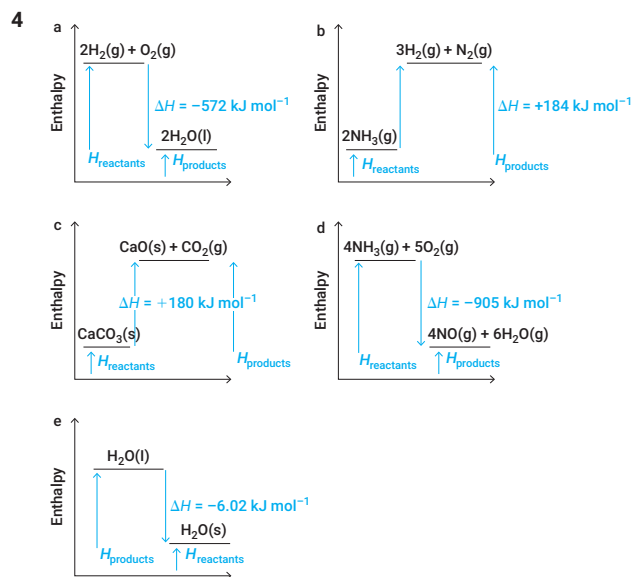
- 1 a** A reaction in which energy is absorbed from the surroundings, causing a temperature decrease
- b** A reaction in which energy is released to the surroundings, causing a temperature increase
- c** The total energy possessed by a chemical substance, which changes as bonds are broken or formed during a chemical reaction

- d** The amount of energy required to break 1 mole of the stated bond to give separated atoms, usually expressed in units of kJ mol^{-1}

- 2 a** True **b** False

APPLYING

- 3 a** $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$ $\Delta H = -572 \text{ kJ mol}^{-1}$
- b** $2\text{NH}_3(\text{g}) \rightarrow 3\text{H}_2(\text{g}) + \text{N}_2(\text{g})$ $\Delta H = +184 \text{ kJ mol}^{-1}$
- c** $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ $\Delta H = +180 \text{ kJ mol}^{-1}$
- d** $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$ $\Delta H = -905 \text{ kJ mol}^{-1}$
- e** $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{s})$ $\Delta H = -6.02 \text{ kJ mol}^{-1}$



- 5 a** Y is highest in energy.
b X is lowest in energy.
c X is most stable.
d Y is least stable.
- 6 a** The reactants are less stable and the products are more stable. The enthalpy change is negative.
b The reactants are less stable and the products are more stable. The enthalpy change is negative.
- 7** Reactants: $945 (\text{N}_2) + 1308 (3 \times \text{H}_2) = 2253 \text{ kJ mol}^{-1}$
 Products: $-2346 (3 \times \text{NH}_3) \text{ kJ mol}^{-1}$
 $\Delta H = 2253 - 2346 = -93 \text{ kJ mol}^{-1}$
- 8** Reactants: $4004 (\text{C}_3\text{H}_8) + 1743 (3\frac{1}{2} \times \text{O}_2) = 5747 \text{ kJ mol}^{-1}$
 Products: $3216 (3 \times \text{CO}) + 3704 (4 \times \text{H}_2\text{O}) = 6920 \text{ kJ mol}^{-1}$
 $\Delta H = 5747 - 6920 = -1173 \text{ kJ mol}^{-1}$

LEARNING CHECK 10.3

DESCRIBING

- 1 a The heat gained or lost per mole of reactant: enthalpy change = (enthalpy of products) – (enthalpy of reactants)
- b The heat energy required to raise the temperature of 1 gram of a substance by 1 kelvin

APPLYING

- 2 16 720 J or 16.72 kJ 4 a 1476 J or 1.48 kJ
- 3 74°C b 39.1 g

LEARNING CHECK 10.4

DESCRIBING

- 1 a An exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that produces energy
- b The heat released when a fuel undergoes complete combustion at standard atmospheric pressure
- c Atmospheric pressure at sea level, described as 1 atm, or 101.325 kPa
- d The heat evolved when an acid and a base react to form a salt and water

APPLYING

- 2 33.7 kJ
- 3 a -30.4 kJ g⁻¹
- b The main cause of discrepancy will derive from the loss of heat from the spirit burner to the surroundings. In other words, not all the heat released by the burning butanol will be absorbed by the water in the calorimeter. This may be due to air currents disturbing the flame as well as poor insulation of the calorimeter leading to its own heat loss to the surroundings, which yields a lower measured temperature rise of the water. This will yield a lower measured estimate for ΔH_c for the butanol sample.
- 4 -41.2 kJ mol⁻¹

LEARNING CHECK 10.5

DESCRIBING

- 1 An enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs.
- 2 -511 kJ 4 +51 kJ
- 3 -385.4 kJ 5 -1789 kJ

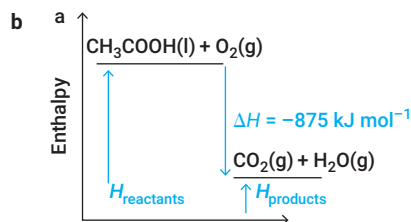
CHAPTER EXAM

MULTIPLE CHOICE

- 1 A 3 C 5 D 7 A 9 B
- 2 C 4 B 6 C 8 A 10 A

SHORT RESPONSE

- 11 a $2\text{Cu}(\text{NO}_3)_2(\text{s}) \rightarrow 2\text{CuO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$
- b Endothermic, because the reaction requires strong heating for it to proceed. The products have more stored energy in their bonds than the reactants do.
- 12 a $\text{CH}_3\text{COOH}(\text{l}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \quad \Delta H = -875 \text{ kJ}$



CROSS-CHAPTER QUESTION

- 13 a
- b Ammonia is a covalent molecular substance.
- c Is a gas at room temperature, it has a low melting and boiling point
- d $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
- e This reaction is exothermic.
- f The products are most stable in this reaction as they have the lowest energy.
- g The number of moles of N_2 is 7.94 mol. If 1 mol of N_2 produces 92 kJ, then 7.94 mol of N_2 will produce 730 kJ.

DATA ANALYSIS

- 14 a i $\text{C}_4\text{H}_{10}(\text{l}) + 6\frac{1}{2}\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O}(\text{g})$
 $\Delta H = -2877 \text{ kJ mol}^{-1}$
- ii 11.1 kJ
- iii Molar heat of combustion = $-1.79 \times 10^3 \text{ kJ mol}^{-1}$
- iv 37.4%
- v Measurement error can arise from inaccuracies in measuring masses, and variability in measuring temperatures. Instrumental uncertainty can arise from limited precision of thermometers, uncertainty in the calibration of instruments and limitations of the calorimeter design, such as heat loss to the surroundings. The nature of the methodology can also introduce uncertainties and limitations. For example, if the combustion reaction is not carried out under controlled conditions, the enthalpy change may not accurately reflect the true heat released. Assumptions made in the calculation of enthalpy change, such as ignoring heat loss to the surroundings, will introduce errors. Incomplete combustion may be occurring.

- vi There are both systematic and random errors due to the reasons outlined in part v.
- b i $2.79 \times 10^2 \text{ kJ/100 g}$
- ii A biscuit is not a pure substance; it is a mixture, so does not have a chemical formula or a molar mass.
- c $4.46 \times 10^2 \text{ kJ/100 g}$
- d To minimise measurement error, multiple trials should be conducted and the results averaged. Calibrated instruments and precise measurement techniques can also help to reduce errors. To address instrumental uncertainty, use instruments with appropriate precision. To minimise uncertainties related to methodology, the experiment must be designed carefully to control variables as much as possible. The insulation of the calorimeter can be increased to minimise heat loss during the experiment. The experiment could be extended by comparing the results obtained from other calorimetry experiments with theoretical predictions or results from other analytical techniques.

CHAPTER 11 INTERMOLECULAR FORCES

LEARNING CHECK 11.1

DESCRIBING

- a A pair of electrons shared by two atoms to form a covalent bond

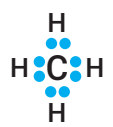
b A valence pair of electrons present in a covalent substance that is not involved in the covalent bond

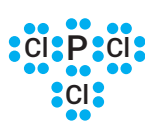
c Electrons in the outermost shell of an atom in its ground state

APPLYING

- a  Linear

b  Linear

c  Tetrahedral

d  Trigonal pyramidal

e  Tetrahedral

f  Trigonal planar



CO_2 is linear, whereas SO_2 is bent because of the lone pair of electrons on the sulfur atom.

- Ethyne contains a $\text{C}\equiv\text{C}$ triple bond and two $\text{C}-\text{H}$ bonds. The $\text{C}\equiv\text{C}$ triple bond is counted as one bonding pair, giving a total of three bonding pairs. This gives a linear shape.



LEARNING CHECK 11.2

DESCRIBING

- a A permanent build-up of negative charge at one end, and positive charge at another end, of a covalent bond or molecule

b A covalent bond that has uneven sharing of electrons

c A covalent bond that has relatively equal sharing of electrons

APPLYING

- a $\text{O}-\text{H}$: $\text{O } \delta^-$, $\text{H } \delta^+$
 $\text{Cl}-\text{H}$: $\text{Cl } \delta^-$, $\text{H } \delta^+$
 $\text{C}-\text{H}$: very small difference, $\text{C } \delta^-$, $\text{H } \delta^+$
 $\text{P}-\text{H}$: no difference
 $\text{N}-\text{O}$: $\text{O } \delta^-$, $\text{N } \delta^+$
 $\text{N}-\text{Cl}$: very small difference $\text{N } \delta^+$, $\text{Cl } \delta^-$
 $\text{N}-\text{H}$: $\text{N } \delta^-$, $\text{H } \delta^+$
 $\text{S}-\text{H}$: $\text{S } \delta^-$, $\text{H } \delta^+$
 $\text{C}-\text{S}$: no difference
 $\text{F}-\text{O}$: $\text{F } \delta^-$, $\text{O } \delta^+$
 $\text{Cl}-\text{O}$: $\text{O } \delta^-$, $\text{Cl } \delta^+$
 $\text{C}-\text{O}$: $\text{O } \delta^-$, $\text{C } \delta^+$
 $\text{P}-\text{O}$: $\text{O } \delta^-$, $\text{P } \delta^+$
 $\text{C}-\text{Cl}$: $\text{Cl } \delta^-$, $\text{C } \delta^+$

b Highly polar: $\text{O}-\text{H}$, $\text{N}-\text{H}$, $\text{Cl}-\text{H}$, $\text{C}-\text{O}$, $\text{P}-\text{O}$, $\text{C}-\text{Cl}$
Slightly polar: $\text{N}-\text{O}$, $\text{N}-\text{Cl}$, $\text{S}-\text{H}$, $\text{F}-\text{O}$, $\text{Cl}-\text{O}$, $\text{C}-\text{H}$
Non-polar: $\text{P}-\text{H}$, $\text{C}-\text{S}$
- a Polar c Polar e Non-polar
b Polar d Polar

LEARNING CHECK 11.3

DESCRIBING

- Metallic, ionic and covalent
- Dipole-dipole, hydrogen bonding and dispersion forces
- Nitrogen, oxygen or fluorine when attached to a hydrogen atom

APPLYING

- 4 Dispersion, dipole–dipole, hydrogen bonding
- Dispersion forces arise from temporary fluctuations in the electron distribution within molecules, leading to temporary induced dipoles.
- Dipole–dipole attractions occur between polar molecules where the permanent dipoles lead to stronger attractions between molecules with opposite charges.
- Hydrogen bonding is an enhanced dipole–dipole attraction as the hydrogen is bonded to the highly electronegative atoms oxygen, nitrogen and fluorine. This large difference leads to the stronger hydrogen bonding.
- 5 Substances dissolve best in solvents that have the same intermolecular forces as them.
- 6 Dihydrogen sulfide exhibits dipole–dipole intermolecular forces. Water exhibits hydrogen bonding, a much stronger intermolecular force. This requires a much greater thermal energy to overcome, meaning that water has a much higher boiling point.

ANALYSING

- 7 In general, the trend in boiling points of group 17 hydrides increases down the group (with the exception of HF). This is due to the increase in strength of the dispersion forces as a result of an increase in atomic size. Hydrogen bonding in HF contributes significantly to its higher boiling point compared to the other hydrides in the group.

CHAPTER EXAM

MULTIPLE CHOICE

1 A	3 B	5 B	7 B	9 A
2 B	4 A	6 D	8 C	10 C

SHORT RESPONSE

- 11 F—P or F—H or P—H
- 12 Both molecules have dispersion forces but CCl_4 contains more electrons, so its dispersion forces will be stronger and more energy will be required to move from liquid to gaseous state, therefore the boiling point will be higher.

CROSS-CHAPTER QUESTION

- 13 a Silicon tetrafluoride is a covalent compound and sodium chloride is an ionic compound.
- b Silicon tetrafluoride is a non-polar covalent compound so it will only have weak dispersion forces between the molecules, whereas sodium chloride is composed of sodium cations and chloride anions, resulting in strong electrostatic forces between them in a three-dimensional lattice structure. These forces are much stronger than the dispersion forces between silicon tetrafluoride, resulting in a higher boiling point.

DATA ANALYSIS

- 14 a Unknown molecule A is octane and unknown molecule B is hexane.
- b Octane is a larger molecule than hexane, so will have stronger dispersion forces between the molecules, resulting in a higher boiling point.
- 15 a X: dispersion forces, dipole–dipole attractions and hydrogen bonds
Y: dispersion forces, dipole–dipole attractions and hydrogen bonds
Z: dispersion forces and dipole–dipole attractions
- b Similarities: They are both organic compounds, both can form hydrogen bonds between molecules.
Differences: X has a —OH attached to one of the carbon atoms, whereas Y has a —COOH attached. Y has an extra oxygen that is double bonded to the carbon.
The boiling point of Y is higher than that of X as the —COOH gives it a more polar structure with the addition of the extra oxygen, which will strengthen the forces between the molecules.
- c Z will have a lower boiling point than X and Y. There are no hydrogen bonds between the molecules of Z because it does not have an —OH group as X and Y do.

CHAPTER 12 CHROMATOGRAPHY

LEARNING CHECK 12.1

DESCRIBING

- 1 The mobile phase dissolves the mixture and carries the components along, while the stationary phase is used to attract the different components and hold onto them.
- 2 These techniques can determine the composition of a mixture and whether a substance is pure.
- 3 Both use a liquid mobile phase. Thin-layer chromatography has a solid stationary phase, whereas in paper chromatography the water trapped in the paper is the liquid stationary phase. They can separate similar types of mixtures, they work on the same chemical basis; however, while paper chromatography is simpler, TLC is faster, produces better separations, is more sensitive and can be used with a wider range of solvents.

APPLYING

- 4 A large quantity of material would soon clog up a TLC plate.
- 5 a To provide known samples with which to compare samples from police
b The sample contains the commercial drug Y because it produces a spot with an R_f value of $\frac{14.5}{20} = 0.725$, which is the same as the R_f value for the drug Y.

$$6 \quad G = \frac{8}{16} = 0.50$$

$$P = \frac{12}{16} = 0.75$$

$$R = \frac{4}{16} = 0.25$$

$$Y = \frac{10.5}{16} = 0.66$$

- 7 a 10, 4, 2, 3, 5, 6, 7, 8, 9, 1

- b Any three of:

- A pencil should be used, not a gel pen.
- The spot should not be too large.
- The origin should be above the solvent.
- The component that absorbs most strongly will travel the shortest distance.

LEARNING CHECK 12.2

DESCRIBING

- 1 a The substance in which the solute dissolves or the greater part of a solution
- b The substance that is dissolved or the smaller component of a solution
- c The intermolecular attraction between hydrogen in —OH, —NH or —FH and the lone pairs on nitrogen, oxygen or fluorine on an adjacent molecule
- d A covalent bond that has a separation of charge
- e A covalent bond that has little or no separation of charge
- 2 a In a non-polar stationary phase on a TLC plate, hydrogen-bonding substances would be insoluble, so would spend most of their time in the mobile phase. They would appear at the top of the plate close to the solvent front. Dipole–dipole substances would appear next, followed by the dispersion-only substances which would be the lowest.
- b In a polar stationary phase, the situation described in part a would be reversed.
- 3 The cellulose fibres in paper strongly attract water molecules. The water molecules bonded to the surface of the paper fibres are the liquid stationary phase, as the paper acts as the support for these molecules. The solvent is a mobile liquid.
- 4 An R_f value close to 1 indicates the component is more strongly attracted to the mobile phase than to the stationary phase and will continue to travel with the mobile phase while not adhering to the stationary phase.

APPLYING

- 5 While each spot represents a particular component, it is possible that very small quantities of other substances

travel with each component and will spend slightly different amounts of time in the stationary phase

- 6 Changing the polarity of the stationary phase can make chromatography more effective, but great care needs to be taken to choose a stationary phase that does not chemically interfere with the components to be separated.
- 7 a Black jelly bean: 5 spots so 5 dyes
Green jelly bean: 2 spots so 2 dyes
Pink jelly bean: 1 spot so 1 dye
Blue jelly bean: 1 spot so 1 dye
- b The red dye was most attracted to the stationary phase because it travelled the least distance from the origin.
- c The blue and yellow dyes must be either non-polar or only slightly polar molecules as they were not strongly attracted to the highly polar stationary phase.
- 8 a Aspartic acid is more attracted to the stationary phase than leucine, since it has travelled a shorter distance up the stationary phase. (This suggests that it is a more polar molecule than leucine.)
- b $R_f = \frac{3.3}{20} = 0.17$
- c The lemon juice tested contained aspartic acid, since it showed a spot that was at the same height and therefore had the same R_f value in the same solvent as aspartic acid. The second spot was another amino acid, but this was not leucine, as it was not at the same height as the spot for leucine.

CHAPTER EXAM

MULTIPLE CHOICE

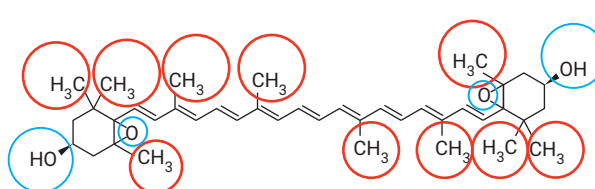
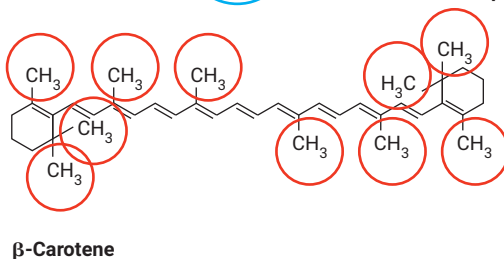
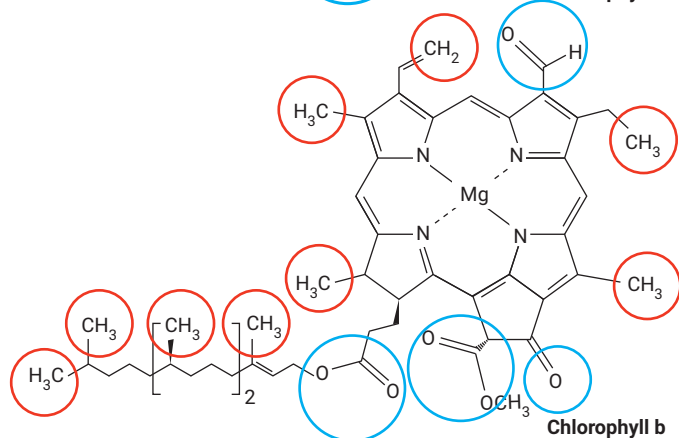
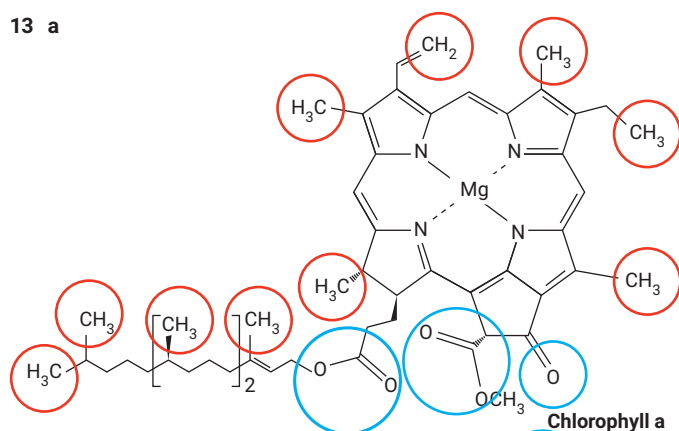
1 C	3 A	5 C	7 D	9 B
2 C	4 A	6 A	8 A	10 B

SHORT RESPONSE

- 11 a The attraction to the surface of a material; the opposite of 'desorption'
- b The action of the substance moving from the stationary phase into the mobile phase
- c The phase that carries solutes through the stationary phase
- d The substance that is fixed in place during chromatography, to which the solute adsorbs
- e Retardation factor; the distance the solute moves compared to the distance the solvent has moved
- 12 The different types of polar groups, the amount of charged and polar chemical groups present, molecular weight, geometry, the positions and numbers of carbon–carbon double bonds

CROSS-CHAPTER QUESTION

13 a



○ Polar
○ Non-polar

Chlorophyll a: 10 non-polar groups, 3 polar groups;
 chlorophyll b: 9 non-polar groups, 4 polar groups;
 β-carotene: 10 non-polar groups; violaxanthin: 10
 non-polar groups, 4 polar groups

- b** Violaxanthin (most polar), chlorophyll b, chlorophyll a,
 β-carotene (non-polar)

Violaxanthin must be the most polar because it has
 two highly polar alcohol groups and two polar ether
 groups. Chlorophyll b must be the next most polar as it
 has an aldehyde and ketone group as well as two ester
 groups. β-Carotene must be the least polar as it has no
 polar groups. This means chlorophyll a with two ester

groups and one ketone group must be less polar than
 chlorophyll b but more polar than β-carotene.

- c** Stationary phase – silica gel surface of TLC plate
 Mobile phase – solvent mixture of hexane, acetone,
 trichloromethane

Stationary phase is polar as silica gel has many —OH
 groups

Mobile phase will be predominately non-polar as it is
 composed of 3 parts non-polar hexane, 1 part slightly
 polar acetone and 1 part polar trichloromethane.

- d** W = violaxanthin, X = chlorophyll b,
 Y = chlorophyll a, Z = β-carotene

As the stationary phase is highly polar, the most polar pigment, violaxanthin, will attach strongly and only move slowly with the mobile phase, which is only very slightly polar. The non-polar pigment, β -carotene, will travel faster with the non-polar solvent and be the last to attach. The other two pigments will attach to the stationary phase, with chlorophyll b slightly below chlorophyll a because it is slightly more polar.

- e Solvent distance = 10 cm

$$W: R_f = \frac{1.5}{10} = 0.15$$

$$X: R_f = \frac{4.2}{10} = 0.42$$

$$Y: R_f = \frac{5.9}{10} = 0.59$$

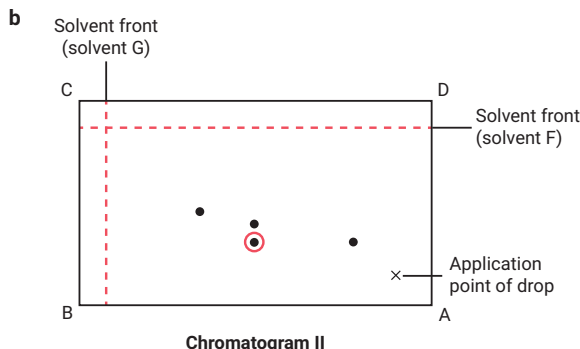
$$Z: R_f = \frac{9.7}{10} = 0.97$$

DATA ANALYSIS

- 14 Use a thin-layer polar plate such as one coated by silica. The mobile phase, ethyl ethanoate is moderately polar (Table 12.2.1) so the aspirin, which is polar, would dissolve and attach to the plate. Run the chromatogram using a sample of pure aspirin, a sample of pure salicylic acid, a sample of pure ethanoic anhydride and a sample of the prepared aspirin product.

Check the chromatogram under UV light. If the prepared sample is pure, it should have only one spot with an R_f value the same as that of the pure aspirin sample. If there is more than one spot, then the sample is impure. If the impurity is one of the reactants, it can be identified by comparing the R_f of the pure reactants with that of the product sample spot.

- 15 a Tyrosine



- c In chromatogram I, which used solvent G, alanine and threonine appeared as the same spot, so were not separated, because they had the same R_f value (0.51). In chromatogram II, which used solvent F, alanine and threonine had different R_f values and were separate.

CHAPTER 13 GASES

LEARNING CHECK 13.1

DESCRIBING

- a A state of matter in which particles flow easily to take the shape and volume of their container

b Equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles

c The volume occupied by 1 mole of a gas at known temperature and pressure
- 100 kPa, 273 K (0°C)

APPLYING

- The pressure must have increased. It is the same amount of gas at the same temperature but a smaller volume, so the only way the volume can be decreased is to increase the pressure.
- a 2.16 L

b i 0.02 mol ii 0.64 g

c i 112 L ii 5.2 L

iii There are more moles of H_2 in 10 g than there are moles of CO_2 . The greater the number of moles of gas present, the greater the volume occupied by the gas.
- At STP, the same volume of gases contain the same number of particles. Therefore, the number of particles of helium and oxygen would be the same.

LEARNING CHECK 13.2

DESCRIBING

- Gases consist of particles that move in continual, random straight-line motion.
 - The average distance between gas molecules is very large compared to the size of the molecule. Intermolecular forces between molecules are negligible.
 - All collisions of gas molecules are perfectly elastic collisions, which means there is no net energy loss during these collisions.
 - Pressure is due to collisions of the molecules with the walls of the container. Temperature is a measure of the average kinetic energy of the molecules.

APPLYING

- At higher temperatures, the gas particles move faster. This causes the particles to collide more frequently with the walls of the balloon, which results in greater pressure on the inside of the balloon, causing it to burst.
- Hydrogen gas, H_2 , has a molar mass of 2.02 g mol^{-1} , whereas oxygen gas, O_2 , has a molar mass of 32 g mol^{-1} . Hydrogen is much lighter, so to produce the same pressure as oxygen, the molecules move much faster.

- 4 Molar mass of $\text{HCl} = 36.5 \text{ g mol}^{-1}$, molar mass of $\text{NH}_3 = 17 \text{ g mol}^{-1}$. The white 'smoke' will appear closer to where the HCl is introduced.

NH_3 , being a lighter molecule, will travel faster and, therefore, further.

- 5 a 4.11 L c 283 mL
b 1.4 atm d 25 kPa
6 a 4.2 mL c 53 mL
b i 2.1 L
ii $140 \text{ K} = -132^\circ\text{C}$
7 a 14 L b 0.74 g c 336 kPa
8 1.12 L
9 Volume of oxygen and carbon dioxide = 18.9 L
10 300 g

CHAPTER EXAM

MULTIPLE CHOICE

- 1 B 3 A 5 C 7 A 9 B
2 B 4 A 6 B 8 A 10 D

SHORT RESPONSE

- 11 28 mL 12 1.39 L

CROSS-CHAPTER QUESTION

- 13 a 2.313 mol b 210 L

DATA ANALYSIS

- 14 $7.7 \times 10^2 \text{ kPa}$
15 a Dependent variable: pressure (atm)
Independent variable: volume (L)
b Boyle's law
c Temperature

CHAPTER 14 AQUEOUS SOLUTIONS AND MOLARITY

LEARNING CHECK 14.1

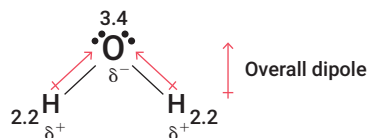
DESCRIBING

- 1 a The mass per unit volume of a substance
b The tension of the surface film of a liquid, due to the attraction between particles in the surface layer
c The pressure of a substance in the gaseous state above its liquid
d The ability of a liquid to flow through narrow spaces in opposition to gravity
e Water molecules present in a fixed ratio that form an essential part of the crystal structure of some compounds

- 2 Water freezes at 0°C and boils at 100°C . It can act as a solvent over this entire temperature range, dissolving and transporting materials across a cell or over oceans. It is the only substance that exists naturally in its three states of matter: solid, liquid and gas. And, unlike most other substances, solid water is less dense than the liquid form, which explains ice's capacity to float on liquid water. Boiling point, density in solid and liquid phases, surface tension and ability to act as a solvent are among water's unique properties.
- 3 Water's surface tension is high because hydrogen bonding occurs in all directions and is relatively strong. It holds the molecules together.
- 4 Vapour pressure of water is relatively low at room temperature because many of the water molecules at this temperature do not have enough energy to overcome the strong hydrogen bonding intermolecular forces and escape as a gas. This means that few water molecules exist in the gas phase to build vapour pressure.

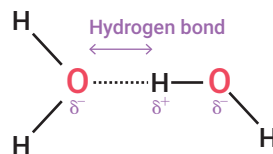
APPLYING

- 5 Water expands when it freezes. If the bottle is completely full, it can shatter, due to the expansion of the ice as it forms.
- 6 The water molecule has polar covalent bonds ($\text{O}-\text{H}$) within the molecule. According to VSEPR theory, the lone pairs on oxygen give the water molecule a bent shape. This, combined with the polar covalent bonds within the molecule, creates an overall dipole.



The polar water molecule, therefore, experiences intermolecular forces. Because the water molecule has $\text{O}-\text{H}$ bonds within the water molecule, the intermolecular forces experienced between water molecules is strong hydrogen bonding.

- 7 a 100°C
b Anomaly. Water has a much higher than expected boiling point given the trend shown by the other substances. Its boiling point is unusually high.
c Strong intermolecular forces (hydrogen bonds) hold the water molecules together and it takes a lot of heat energy to overcome these forces and separate the molecules in the liquid.



ANALYSING

- 8 This would be effective because it would allow the water to be removed from the crystal with each heating. If water was still present in the sample, the mass would decrease with each subsequent heating and weighing. Once the mass stops changing after heating events, it can be assumed that the sample is anhydrous.
- 9 The smaller the capillary tube, the higher the water will travel due to adhesive forces between the water molecules and the sides of the capillary tubes and cohesive forces between water molecules. Water moves up through the xylem cells in the trunk and stems through capillary action. So, at the top of the tree, where the gravitational potential energy is highest, the tubes are at their narrowest.
- 10 $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

LEARNING CHECK 14.2

DESCRIBING

- 1 B
- 2 Water will dissolve a wide range of substances, so many materials can dissolve and be transported in bodies of water and in living organisms.

APPLYING

- 3 Polar water molecules are attracted to and surround the sodium and chloride ions so they dissolve.
- 4 The solute is the sodium chloride that dissolved in the solvent. The solvent is water in an aqueous solution.
- 5 The attraction holding together the solute particles (the barium and sulfate ions) is too strong to break – even in a polar solvent such as water. Therefore, dissolution cannot occur.

LEARNING CHECK 14.3

DESCRIBING

- 1 A 0.15 M solution of sodium chloride solution is *less* concentrated than a 0.20 M potassium chloride solution.
- 2 Molarity is the unit of measurement for concentration, whereas concentration describes the amount of solute in a given quantity of solvent.
- 3 A concentrated solution has a large proportion of solute compared to solvent, whereas strength, in terms of acids and bases, describes the ability of the substance to dissociate or ionise.

APPLYING

- 4 a 2 M c 2.1×10^{-1} M
b 5.2 M d 5.4×10^{-4} M
- 5 a i 0.750 mol L⁻¹ ii 89.3 g L⁻¹
b i 0.010 mol L⁻¹ ii 1.50 g L⁻¹

- 6 12 g of sugar is present.

7 Acetic acid concentration is $\frac{2.1}{0.25} = 8.4$ M.

Molecular weight of KCl = 75 g.

Potassium chloride concentration is $\frac{1.42}{75} \times 2 = 0.038$ M.

So the acetic acid solution is more concentrated.

LEARNING CHECK 14.4

DESCRIBING

1 $c_1V_1 = c_2V_2$

APPLYING

- 2 a 0.040 mol Ba(NO₃)₂ c 0.064 mol CuCl₂
b 0.075 mol Na₂SO₄
- 3 a 10 g b 10.7 g c 8.6 g
- 4 a 0.18 M b 0.0045 M c 0.67 M
- 5 a 0.053 g b 0.51 g
- 6 0.040 L = 40 mL
- 7 $n(\text{NaOH}) = 0.0450$ mol
 $n(\text{Cu}(\text{OH})_2) = \frac{1}{2}n(\text{NaOH}) = 0.0225$ mol
 $m = 2.2$ g.
- 8 $\text{ZnCl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s}) + 2\text{NaCl}(\text{aq})$
 $n(\text{ZnCl}_2) = 0.0050$ mol
 $n(\text{Zn}(\text{OH})_2) = 0.0050$ mol
 $m = 0.50$ g
- 9 a 0.0000 468 mol MgCl₂
b 0.0000 468 mol Mg²⁺
c 0.0000 936 mol Cl⁻
- 10 a 1.7 M c 0.24 M e 1.07 M
b 0.125 M d 2.07 M
- 11 $n(\text{AgNO}_3) = 0.00375$ mol
 $n(\text{AgI}) = 0.00375$ mol
 $m = 0.88$ g

LEARNING CHECK 14.5

DESCRIBING

- 1 Saturated, unsaturated, supersaturated

APPLYING

- 2 a The solution is unsaturated.
b Saturated solution
c There would be insufficient water to dissolve all of the crystals.
d Crystals would form as the solution would be supersaturated.

CHAPTER EXAM

MULTIPLE CHOICE

- 1 C 3 C 5 A 7 D 9 D
2 B 4 C 6 C 8 D 10 D

SHORT RESPONSE

- 11 a 0.0242 M c 0.004 56 M
b 0.0019 M
12 a 0.40 mol c 0.00375 mol
b 0.0025 mol

CROSS-CHAPTER QUESTION

- 13 a $\text{Na}_2\text{CO}_3(\text{aq}) + \text{MgCl}_2(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{MgCO}_3(\text{s})$
b 0.006 mol
c 0.006 mol
d 0.5

DATA ANALYSIS

- 14 a As temperature increases, solubility increases
b 28.5 g/100 g water
c 1.79 M
15 a Reactant. The trend in the graph shows that the concentration of HCl decreased over time, which means it was being consumed during the reaction.
b 50 mol

CHAPTER 15 IDENTIFYING IONS IN SOLUTION

LEARNING CHECK 15.1

DESCRIBING

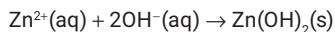
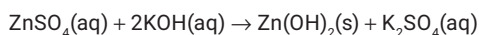
- 1 a An insoluble solid formed from a chemical reaction between two solutions
b An ion that remains in solution before and after a chemical reaction, so it does not take part in the reaction
c When the ions in an ionic compound separate from one another to form a solution in water
d A solution that contains ions, and will therefore conduct electricity
e A balanced equation showing all ions present in a chemical reaction, both those reacting and spectator ions
f A simple balanced equation showing the formulas of all compounds taking part in a chemical reaction, showing no ions

- g A balanced equation that shows only the species that are taking part in the chemical reaction; spectator ions are not included
h The solution that remains on top of the precipitate, composed of spectator ions in aqueous solution

APPLYING

- 2 a Soluble c Soluble e Soluble
b Insoluble d Insoluble
3 a $\text{NaCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{AgCl}(\text{s})$
 $\text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{AgCl}(\text{s})$
 $\text{Cl}^-(\text{aq}) + \text{Ag}^+(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
b $\text{CuSO}_4(\text{aq}) + 2\text{KOH}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + \text{K}_2\text{SO}_4(\text{aq})$
 $\text{Cu}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{K}^+(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + 2\text{K}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
 $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
c $\text{Na}_2\text{CO}_3(\text{aq}) + \text{FeSO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{FeCO}_3(\text{s})$
 $2\text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{FeCO}_3(\text{s})$
 $\text{Fe}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{FeCO}_3(\text{s})$
d $\text{Zn}(\text{NO}_3)_2(\text{aq}) + 2\text{NH}_4\text{OH}(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s}) + 2\text{NH}_4\text{NO}_3(\text{aq})$
 $\text{Zn}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{NH}_4^+(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s}) + 2\text{NH}_4^+ + 2\text{NO}_3^-(\text{aq})$
 $\text{Zn}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s})$
e $\text{K}_2\text{CO}_3(\text{aq}) + \text{CaCl}_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{KCl}(\text{aq})$
 $2\text{K}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) + \text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{K}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$
 $\text{CO}_3^{2-}(\text{aq}) + \text{Ca}^{2+}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$
4 a Precipitate = zinc carbonate; therefore, sodium carbonate, lithium carbonate and ammonium carbonate are suitable.
b Precipitate = lead sulfate; therefore, any group 1 sulfate is suitable, e.g. sodium sulfate, potassium sulfate.
c Precipitate = barium sulfate; therefore, barium bromide, barium chloride, and barium iodide are suitable.
5 a Magnesium chloride and sodium carbonate
 $\text{MgCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{MgCO}_3(\text{s}) + 2\text{NaCl}(\text{aq})$
b Sodium sulfate and lead(II) nitrate
 $\text{Na}_2\text{SO}_4(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow 2\text{NaNO}_3(\text{aq}) + \text{PbSO}_4(\text{s})$
c Silver nitrate and potassium bromide
 $\text{AgNO}_3(\text{aq}) + \text{KBr}(\text{aq}) \rightarrow \text{AgBr}(\text{s}) + \text{KNO}_3(\text{aq})$

6 Zinc hydroxide



- 7 a $\text{Cu}(\text{OH})_2$, copper(II) hydroxide
b None produced
c PbSO_4 , lead(II) sulfate

ANALYSING

8 B

LEARNING CHECK 15.2

DESCRIBING

- 1 a A mixture of a liquid and an insoluble solid, in which the solid distributes evenly throughout the liquid and gives a cloudy appearance
b An aqueous solution of calcium hydroxide ($\text{Ca}(\text{OH})_2$).
2 An acid is added to the sample to determine if there is any fizzing (indicating CO_2 gas is given off). The CO_2 gas is bubbled through the limewater, which produces a precipitate and turns the limewater cloudy. This, in combination, indicates that carbonates were in the original sample.

APPLYING

- 3 a No precipitation
c $2\text{NH}_4\text{Br}(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{PbBr}_2(\text{s}) + 2\text{NH}_4\text{NO}_3(\text{aq})$
 $2\text{Br}^{-}(\text{aq}) + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{PbBr}_2(\text{s})$
c No precipitation
4 a Lead(II) nitrate and sodium sulfate
b Iron(II) nitrate and ammonium carbonate
c Magnesium nitrate and potassium hydroxide
5 Add some ammonium chloride solution – if lead, a precipitate of lead chloride will form; if calcium, no precipitate will form.
6 a React both with lithium hydroxide. Lead(II) nitrate will form a precipitate when mixed (lithium hydroxide), barium nitrate will not.
b React both with sodium iodide. Copper(II) sulfate will produce a precipitate (copper(II) iodide), iron(II) sulfate will not.
7 Barium – they both formed a white precipitate, but only barium produces the pale green flame colour, while calcium would produce a red flame colour.
8 A: lead; B: iron(III); C: barium; D: possibly calcium because none of the other ions listed produce those results. Conduct a flame test and if then flame is red, the presence of calcium is confirmed
9 Answers will vary. Generally, adding Cl^{-} to lead ions will produce a white precipitate, which does not dissolve in ammonia solution. This could be used as the basis for a test for lead.

CHAPTER EXAM

MULTIPLE CHOICE

- 1 D 3 B 5 A 7 D 9 A
2 C 4 A 6 A 8 B 10 A

SHORT RESPONSE

- 11 a $\text{PbCl}_2(\text{aq}) + \text{FeSO}_4(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + \text{FeCl}_2(\text{aq})$
b Lead(II) sulfate
c $\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + \text{Fe}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$
d $\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$
12 The carbonate anion is present, as the substance produced a gas with hydrochloric acid, indicating that CO_2 formed. Furthermore, a precipitate of BaCO_3 forms when barium nitrate reacts with sodium carbonate; however, no precipitate forms when the barium solution reacts with sodium chloride.

CROSS-CHAPTER QUESTION

- 13 a Most ionic solids have a higher solubility at higher temperatures.
b Magnesium carbonate is insoluble so sodium carbonate would be a suitable addition to remove the magnesium ions without interfering with the sodium ions.
c $\text{Mg}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{MgCO}_3(\text{s})$

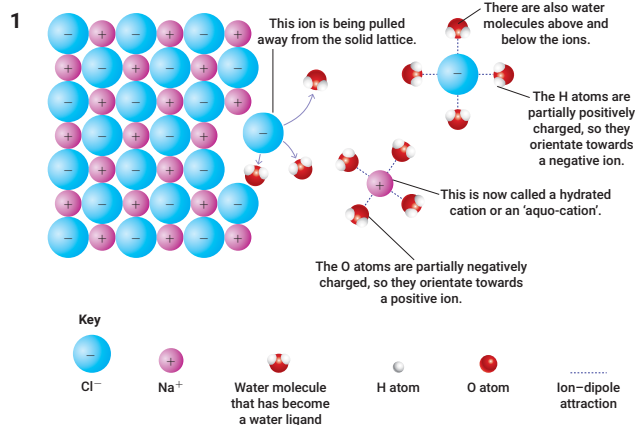
DATA ANALYSIS

- 14 a X: iron(II) iodide; Y: sodium chloride
b FeCO_3
c When barium chloride is mixed with silver nitrate or potassium carbonate, a precipitate will form.
15 a I – lead(II) chloride and III – barium nitrate
b Soluble
c Barium sulfate and ammonium nitrate
d For I, when lead chloride is mixed with ammonium sulfate, the insoluble precipitate lead sulfate is formed. However, for II, when copper chloride is mixed with ammonium sulfate, no precipitate is formed and as a result all products are soluble.
e I: precipitate formed
II: no precipitate formed
III: precipitate formed
IV: precipitate formed
V: no precipitate formed
VI: no precipitate formed

CHAPTER 16 SOLUBILITY

LEARNING CHECK 16.1

DESCRIBING



2 Ion-dipole interactions

3 D

APPLYING

- 4 HBr is polar, so it can form dipole-dipole attractions with water molecules. Methane is non-polar so there are no intermolecular bonds formed with water molecules.
- 5 When sugar is added to a cup of coffee, sugar becomes the solute and water the solvent. Since coffee is mainly made up of polar water molecules, the bonds between the solute sugar molecules break and individual sugar molecules are able to form hydrogen bonds with water and therefore dissolves.
- 6
- The ionic solid splits into positive and negative ions, which form ion-dipole bonds with the water.
 - $\text{KBr(s)} \rightarrow \text{K}^+(\text{aq}) + \text{Br}^-(\text{aq})$
 - Electrolyte: The presence of dissociated ions means that it will conduct electricity.
- 7 $\text{CH}_3\text{CH}_2\text{OH(l)} \rightarrow \text{CH}_3\text{CH}_2\text{OH(aq)}$. Hydrogen bonds between solute ethanol molecules break and new hydrogen bonds form between ethanol molecules and solvent water molecules.

ANALYSING

- 8 Cerium sulfate has stronger forces of attraction between its ions than between water and the individual ions; therefore, there is not enough energy for dissociation to occur.
- 9
- | | |
|---------------------|-------------|
| a Most are soluble. | d Insoluble |
| b Insoluble | e Soluble |
| c Soluble | f Insoluble |
- 10
- | | |
|-----------|-------------|
| a Soluble | c Insoluble |
| b Soluble | d Soluble |

e Soluble	g Insoluble
f Soluble	h Soluble
- 11
- | | |
|-------------------|-------------------|
| a Non-electrolyte | d Non-electrolyte |
| b Electrolyte | e Electrolyte |
| c Electrolyte | |

INTERPRETING

- 12
- | |
|---|
| a Ionic |
| b Molecular |
| c Sodium chloride, copper sulfate |
| d Ethanol, unknown |
| e Sodium chloride dissociates into ions, so it is capable of carrying a charge in solution. |

LEARNING CHECK 16.2

DESCRIBING

- 1
- | |
|---|
| a A measure of the average kinetic energy of the particles that constitute a solid, liquid or gas |
| b A graph showing the variation of the solubility of a substance in water with temperature |

APPLYING

- 2 In general, an increase in temperature results in increasing solubility for solids however a decrease in solubility for gases. Greater kinetic energy causes solid structure to be broken down more easily, so the higher the temperature, the more soluble an ionic solid is. For gases, an increase in kinetic energy causes the forces of interaction between gas molecules and water to break, leading to a decrease in solubility of gases.
- 3 An increase in temperature increases the kinetic energy of particles, so there are more collisions with solute particles and ion-dipole bonds with water are formed faster. Also, solutes with more kinetic energy are more unstable structures so they break apart more easily.

ANALYSING

- 4 Raising sea temperatures can reduce the solubility of oxygen and carbon dioxide in water, which can affect the survival of marine animals and plants.

INTERPRETING

- 5 carbon dioxide and oxygen both decrease in solubility with an increase in temperature, and oxygen is generally less soluble than carbon dioxide.
- 6 Heat mostly increases the solubility of ionic substances, but at different rates. For example, the solubility of NaCl only slightly increases, whereas that of $\text{Pb}(\text{NO}_3)_2$ increases dramatically. The application of heat does not always increase solubility – it may decrease the solubility of a substance such as $\text{Ce}_2(\text{SO}_4)_3$.

LEARNING CHECK 16.3

DESCRIBING

- 1 The maximum amount of solute that can be dissolved in 100 mL water at a given temperature

APPLYING

- 2 a 13.5 g b 46 g c 45 g
3 Unsaturated
4 KCl
5 15.5 g

ANALYSING

- 6 At 70°C, 47 g dissolves in 100 mL, so 94 g dissolves in 200 mL. At 30°C, 16 g dissolves in 100 mL, so 32 g dissolves in 200 mL. 94 g – 32 g = 66 g is precipitated.
7 a 30 g b 40 g c 80 g
8 a At 30°C, there would be 36 g of KCl and 36 g of NaCl.
b The solutions would be saturated because the curve represents where the solutions are saturated.
9 a 30 g b More than 23 g c 16 g
10 a 3.0 M b 4.6 M c 2.0 M

INTERPRETING

- 11 a Answers will vary – generally diagrams should show KCl ionising to form K^+ and Cl^- ions. The negative dipole of the water (oxygen side) forms ion–dipole attractions to the K^+ ion and the positive dipole of the water molecule (hydrogen side) forms ion–dipole attractions to the Cl^- ion until the ions are hydrated.
b $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$
c 26 g
d Supersaturated, as there is still excess solute dissolved in solution
e i 43°C
ii 45 g (at 60°C) – 30 g (at 10°C) = 15 g precipitates
12 a The solubility of all three substances increases as temperature increases.
b The solubility of potassium nitrate increases the most from approximately 14 g per 100 g water at 0°C to over 200 g per 100 g water.
c According to the solubility curve, approximately 39 g of NaCl will dissolve at 80°C in 100 g of water. Therefore half of that value, 19.5 g would be able to dissolve in 50 g of water.
d Sodium chloride
e Lead(II) nitrate is consistently more soluble than potassium nitrate below 40°C. Sodium chloride is more soluble than potassium nitrate until about 23°C but then is less soluble until 40°C
f –63–64°C

- g At 23°C, sodium chloride and potassium nitrate have similar solubilities (~35 g/100 g water), whereas the solubility of lead nitrate is higher (~60 g/100 g water).

CHAPTER EXAM

MULTIPLE CHOICE

- 1 D 3 A 5 B 7 C 9 B
2 A 4 C 6 A 8 C 10 C

SHORT RESPONSE

- 11 At 10°C, sodium chloride is more soluble than potassium nitrate. However, at 80°C potassium nitrate is more soluble than sodium chloride.
12 The student could look up the solubility curve to find out how much should dissolve at the final temperature. They can add more than this amount and heat the solution to dissolve more solute. The student can then leave the solution undisturbed to cool. The cooled solution should be supersaturated once cool.

CROSS-CHAPTER QUESTION

- 13 a 33 g b 33 g – 17 g = 16 g
c $Na_2CO_3(s) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$. The gas formed is carbon dioxide.
d The limewater will turn cloudy as a precipitate is formed when the CO_2 reacts with the limewater solution.
e $n(Na_2CO_3) = 0.14 \text{ mol}$
Ratio of Na_2CO_3 to CO_2 is 1:1, therefore 0.14 mol of CO_2 was formed.

DATA ANALYSIS

- 14 a 205 g in 200 mL water = 82 g/100 g water = ~45°C
b 50°C = 84 g/100 g water = 63 g/75 g water
30°C = 77 g/100 g water = 57.8 g/75 g water
63 g – 57.8 g = 5.2 g
c 5.5 M = 5.5 mol L⁻¹
 $m((NH_4)_2SO_4) = 72.7 \text{ g/100 mL}$
At 40°C, 80 g/100 mL dissolves therefore the solution is unsaturated.
d 90–93 g
15 a 105 g/100 mL = 63 g/60 mL
b Unsaturated. At 20°C, solubility of potassium nitrate is 367.5 g/350 g water. The mass of potassium nitrate dissolved is less than 367.5 g.
c 0.84 M
d Based on the data, an increase in temperature results in an increase in solubility of potassium nitrate.

CHAPTER 17 REACTIONS OF ACIDS AND BASES ANSWERS

LEARNING CHECK 17.1

DESCRIBING

- 1 Acid + base $\rightarrow$ salt + water

APPLYING

- 2 a $\text{HCl(aq)} + \text{KOH(aq)} \rightarrow \text{KCl(aq)} + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)}$
- b $\text{H}_2\text{SO}_4(\text{aq}) + \text{Mg(OH)}_2(\text{s}) \rightarrow \text{MgSO}_4(\text{aq}) + 2\text{H}_2\text{O(l)}$
 $2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Mg(OH)}_2(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{Mg(OH)}_2(\text{s}) \rightarrow \text{H}_2\text{O(l)} + \text{Mg}^{2+}(\text{aq})$
- c $\text{HNO}_3(\text{aq}) + \text{LiOH(aq)} \rightarrow \text{LiNO}_3(\text{aq}) + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Li}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Li}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)}$
- d $2\text{HF(aq)} + \text{Ba(OH)}_2(\text{aq}) \rightarrow \text{BaF}_2(\text{aq}) + 2\text{H}_2\text{O(l)}$
 $2\text{H}^+(\text{aq}) + 2\text{F}^-(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2\text{F}^-(\text{aq}) + 2\text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)}$

LEARNING CHECK 17.2

DESCRIBING

- 1 a When an acid ionises it produces an H^+ in aqueous solution.
- b A salt is an ionic compound formed from cations and anions.
- c A neutralisation reaction occurs when an acid and a base react to form water and a salt.
- d The acid-carbonate reaction occurs when an acid reacts with a carbonate to form a salt, carbon dioxide and water.
- e The acid-metal reaction occurs when an acid reacts with a metal to form hydrogen gas and a salt.
- f The pop test is an experimental test for hydrogen, where a small explosion of hydrogen in a test tube causes a pop sound.

APPLYING

- 2 a $2\text{HNO}_3(\text{aq}) + \text{MgCO}_3(\text{s}) \rightarrow \text{Mg(NO}_3)_2(\text{aq}) + \text{H}_2\text{O(l)} + \text{CO}_2(\text{g})$
- b $\text{H}_3\text{PO}_4(\text{aq}) + 3\text{KOH(aq)} \rightarrow \text{K}_3\text{PO}_4(\text{aq}) + 3\text{H}_2\text{O(l)}$
- c $2\text{CH}_3\text{COOH(aq)} + \text{Mg(s)} \rightarrow \text{Mg(CH}_3\text{COO)}_2(\text{aq}) + \text{H}_2(\text{g})$
- 3 a Hydrogen gas
- b Water

- c Sodium chloride, carbon dioxide

- 4 a $\text{HNO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$
- b $\text{HCl(aq)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
- c $\text{H}_3\text{PO}_4(\text{aq}) \rightarrow 3\text{H}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
- 5 a $\text{HNO}_3(\text{aq}) + \text{KOH(aq)} \rightarrow \text{KNO}_3(\text{aq}) + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{H}_2\text{O(l)}$
 $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)}$
- b $2\text{H}_3\text{PO}_4(\text{aq}) + 3\text{CaCO}_3(\text{s}) \rightarrow \text{Ca}_3(\text{PO}_4)_2(\text{aq}) + 3\text{H}_2\text{O(l)} + 3\text{CO}_2(\text{g})$
 $6\text{H}^+(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) + 3\text{CaCO}_3(\text{s}) \rightarrow 3\text{Ca}^{2+}(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) + 3\text{H}_2\text{O(l)} + 3\text{CO}_2(\text{g})$
 $6\text{H}^+(\text{aq}) + 3\text{CaCO}_3(\text{s}) \rightarrow 3\text{H}_2\text{O(l)} + 3\text{CO}_2(\text{g}) + 3\text{Ca}^{2+}(\text{aq})$
 $2\text{H}^+(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow \text{H}_2\text{O(l)} + \text{CO}_2(\text{g}) + \text{Ca}^{2+}(\text{aq})$
- c $2\text{HCl(aq)} + \text{Zn(s)} \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + 2\text{Cl}^-(\text{aq}) + \text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + \text{H}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + \text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$

CHAPTER EXAM

MULTIPLE CHOICE

- 1 C 3 C 5 A 7 D 9 B
- 2 C 4 D 6 B 8 A 10 D

SHORT RESPONSE

- 11 $2\text{HNO}_3(\text{aq}) + \text{Ca(HCO}_3)_2(\text{aq}) \rightarrow \text{Ca(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O(l)} + 2\text{CO}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{Ca}^{2+}(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{H}_2\text{O(l)} + 2\text{CO}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)} + 2\text{CO}_2(\text{g})$
- 12 a $2\text{HNO}_3(\text{aq}) + \text{Mg(s)} \rightarrow \text{Mg(NO}_3)_2(\text{aq}) + \text{H}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{H}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + \text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2(\text{g})$
- b $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaHCO}_3(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O(l)} + 2\text{CO}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{Na}^+(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{H}_2\text{O(l)} + 2\text{CO}_2(\text{g})$
 $2\text{H}^+(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)} + 2\text{CO}_2(\text{g})$
- c $2\text{H}_3\text{PO}_4(\text{aq}) + 3\text{Mg(s)} \rightarrow \text{Mg}_3(\text{PO}_4)_2(\text{aq}) + 3\text{H}_2(\text{g})$
 $6\text{H}^+(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) + 3\text{Mg(s)} \rightarrow 3\text{Mg}^{2+}(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) + 3\text{H}_2(\text{g})$
 $6\text{H}^+(\text{aq}) + 3\text{Mg(s)} \rightarrow 3\text{Mg}^{2+}(\text{aq}) + 3\text{H}_2(\text{g})$
- d $2\text{KOH(aq)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O(l)}$
 $2\text{K}^+(\text{aq}) + 2\text{OH}^-(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow 2\text{K}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{H}_2\text{O(l)}$
 $2\text{OH}^-(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)}$

CROSS-CHAPTER QUESTION

- 13 a $\text{CH}_3\text{COOH}(\text{aq}) + \text{NaHCO}_3(\text{aq}) \rightarrow \text{CH}_3\text{COONa}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- b The carbon dioxide gas reacted with the $\text{Ca}(\text{OH})_2$ in the limewater to form a precipitate (milky).
- c $\text{Ca}(\text{OH})_2(\text{aq}) + \text{CO}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
 CaCO_3 is the precipitate that turned the limewater milky.

DATA ANALYSIS

- 14 a Yellow
- b The indicator changed colour from yellow to blue as the sodium hydroxide was added.
- c Green
- d 3 mL
- e Red
- 15 a The two substances are similar in appearance (white crystalline powder), but a gas forms when unknown B reacts with vinegar, which does not happen with unknown A.
- b Unknown B is the hydrogen carbonate because it formed a gas, which is presumably carbon dioxide. Unknown B is therefore most likely to be sodium hydrogen carbonate.
- c Carbon dioxide

CHAPTER 18 pH

LEARNING CHECK 18.1

DESCRIBING

- 1 a A substance that produces hydrogen ions, H^+ , in aqueous solution
- b A substance that contains either the oxide (O^{2-}) or the hydroxide ion (OH^-), or produces a hydroxide ion in aqueous solution
- c A measure of the acidity or alkalinity of a solution, based on the concentration of hydrogen ions in a solution
- d A scale that measures the acidity and alkalinity of a solution, based on the concentration of hydrogen ions in a solution
- e A substance that changes colour in solution depending on how acidic or basic the solution is
- 2 Acidic substances have a pH between 0 and 6, basic substances have a pH between 8 and 14, and neutral substances have a pH of 7.
- 3 There is not necessarily a relationship between concentration of an acid or a base and pH. The pH depends on the concentration of H^+ ions or OH^- ions.

APPLYING

- 4 a 0–3 b 6–7 c 8–10 d 11–14
- 5 B and D

LEARNING CHECK 18.2

DESCRIBING

- 1 An Arrhenius acid is a substance that produces H^+ ions in solution, whereas an Arrhenius base is a substance that produces OH^- ions in solution.

APPLYING

- 2 This statement is not accurate but it does capture a general trend. Substances that contain hydrogen atoms bonded to electronegative elements have the potential to act as acids. This is because these hydrogen atoms can dissociate as hydrogen ions in aqueous solutions. However, not all substances that contain hydrogen act as acids (e.g. methane (CH_4)). It contains hydrogen atoms, but they do not dissociate to form hydrogen ions because of the strong covalent bonds between the carbon and the hydrogen.
- 3 a KOH is acting as a base.
- b HClO_4 is acting as an acid.
- c $\text{Ba}(\text{OH})_2$ is acting as a base.

LEARNING CHECK 18.3

DESCRIBING

- 1 a A substance in which almost all of the acid molecules ionise to produce hydrogen ions, H^+ , in aqueous solution
- b A substance in which relatively few of the acid molecules ionise, resulting in only a small number of hydrogen ions, H^+ , in aqueous solution
- c A substance in which almost all the particles dissociate to produce a relatively large number of hydroxide ions, OH^- , in aqueous solution
- d A substance in which only a few of the particles ionise, resulting in only a relatively small number of hydroxide ions, OH^- , in aqueous solution
- 2 Strong acids: hydrochloric acid, nitric acid and sulfuric acid. Strong bases: sodium hydroxide and potassium hydroxide.
- 3 Weak acids: acetic (ethanoic) and carbonic (potassium hydrogen phthalate) acids. Weak bases: ammonia and ethanamine.

APPLYING

- 4 It can be considered a weak acid because only a few of the molecules ionise in water. It can also be considered to be concentrated if it contains a large amount of solute dissolved in a solvent.

ANALYSING

- 5 a The stronger the acid or base is, the greater its electrical conductivity. Nitric acid is a strong acid and acetic acid is a weak acid.
- b $\text{Mg}(\text{OH})_2$ should have a high electrical conductivity. However, the low value is due to its very low solubility.

CHAPTER EXAM

MULTIPLE CHOICE

- 1 C 3 D 5 D 7 C 9 A
2 A 4 B 6 A 8 A 10 B

SHORT RESPONSE

- 11 The bond that enables it to release hydrogen ions is the O—H bond. This bond is a covalent bond.
12 The NH_4^+ ion produced is itself a weak acid. If the HCl and NH_3 have been neutralised, all that is left are the NH_4^+ and Cl^- ions. The Cl^- ion is neutral but the NH_4^+ ion has a pH of about 4.

CROSS-CHAPTER QUESTION

- 13 a 0.064 mol (0.0638 mol to 3 sig. fig.)
b 0.255 mol L^{-1}
c The resulting solution will have a high conductivity and a relatively high pH.

DATA ANALYSIS

- 14 a As the hydrogen ion concentration increases, pH decreases.
b $\text{HCl(aq)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
15 Sample A is most likely taken from an area with high wastewater irrigation because it has the lowest pH of the three samples.

CHAPTER 19 RATES OF REACTIONS

LEARNING CHECK 19.1

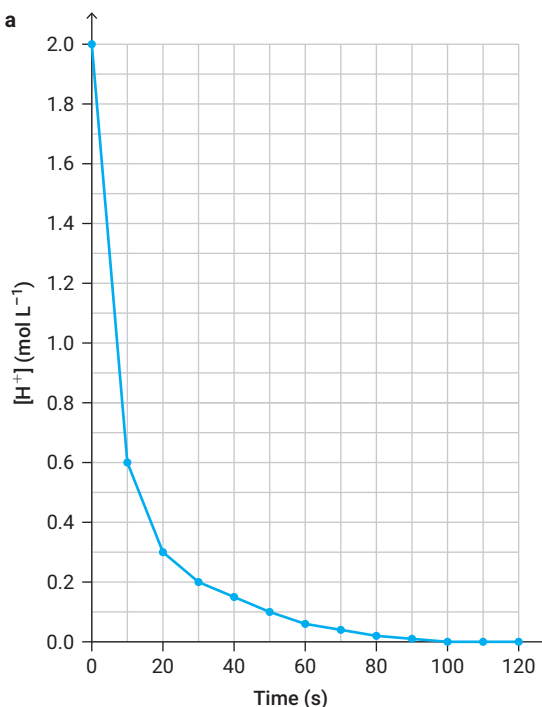
DESCRIBING

- 1 The change in a quantity (such as mass, concentration or volume) over change in time
2 For example, you could measure the changing variable (such as mass, concentration or volume) for a series of measured regular time periods.
Alternatively, you could measure the time taken for a particular change in the variable to occur (such as the time taken for 100 mL of gas to be produced).
3 At the beginning of the reaction, when the concentration of the reactants is highest
4 Measure a change in quantity and divide the magnitude of the change by the time taken for it to occur.
5 The greater the gradient – the steeper the slope – the faster the reaction

APPLYING

- 6 $\text{Rate} = \frac{20}{4} = 5 \text{ g min}^{-1}$
7 Time = 0 s, rate = 11 mL s^{-1}
Time = 1 s, rate = 3.5 mL s^{-1}
Time = 5 s, rate = 0.25 mL s^{-1}

8 a



- b i For change in mass of hydrogen gas:

$$\text{Overall rate} = \frac{2.00 - 0.00}{100} = 2.00 \times 10^{-2} \text{ g s}^{-1}$$

- ii For change in $[\text{H}^+]$:

$$\text{Overall rate} = \frac{0.00 - 2.00}{100} = -2.00 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}$$

- c Although the graphs are the opposite of each other, they show the same pattern of changes where the reaction rate is highest at the beginning, then the reaction slows and ceases at the same time. Although the rate using mass of hydrogen is positive and the rate using $[\text{H}^+]$ is negative, the calculated overall rates both have a value of 2.00×10^{-2} . The difference in sign is due to one measuring an increasing amount of substance while the other measures a decreasing amount of substance. The answer to the question is 'No, the rate of this reaction does not vary with the quantity being measured, provided all factors are the same'.
d Use a different reaction and perform experiments to measure the rate using two different quantities. Compare the rates and see if they are the same.

LEARNING CHECK 19.2

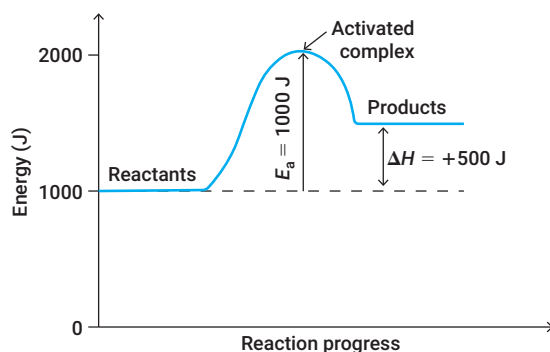
DESCRIBING

- 1 A reaction will occur when two particles collide with sufficient energy, and in the appropriate orientation.
2 a The minimum energy needed for a reaction to occur

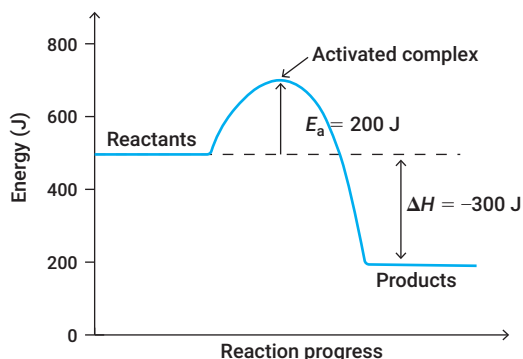
- b The intermediate state of a chemical reaction where bonds are breaking and forming
- 3 Answers will vary. The diagram should look similar to Figure 19.2.7.
- 4 a If the collision is too gentle, then there is not enough energy for the reactant bonds to be broken, and the reaction will not occur.
- b Certain sections of the molecules reacting may need to come into contact with each other for the reaction to occur. Hence, if the molecules are not oriented correctly, the reaction cannot occur.
- 5 If the reactants have weak and/or few bonds, then the amount of energy needed to break them will be low, so the activation energy will also be low. Therefore, a larger number of particles will have enough energy for successful collisions and overcome the activation energy barrier. This means more particles undergoing successful collisions means a faster reaction rate, i.e. weak and/or few bonds → low activation energy → faster reaction; strong and/or many bonds → high activation energy → slower reaction.

APPLYING

6 a



b



LEARNING CHECK 19.3

DESCRIBING

- 1 Concentration, gas pressure, surface area, temperature, catalysts

- 2 A substance that changes the rate of a reaction, but which is not used up in the reaction
- 3 The diagram should be similar to Figure 19.3.5.

APPLYING

- 4 With a greater number of smaller-sized pieces, more of the solid is in contact with the other reactant, so there are more collisions and reaction rate will increase.
- 5 Temperature is a measure of kinetic energy, so the average kinetic energy will be greater at 40°C than at 10°C.
- 6 Both pressure and concentration are a measure of the number of particles per unit volume.
- 7 a Greater concentration or pressure means more reactant particles, so a collision between reactant molecules is more likely, so the reaction rate will be faster.
- b The higher the temperature, the faster the molecules move, so the greater the frequency of collisions and the greater the energy of the collisions, and the more likely the collisions are to be successful.
- c The catalyst provides an alternative pathway for the reaction, with lower activation energy. Hence, a greater proportion of collisions will be successful and the rate of the reaction will be higher.
- d Smaller particle size means that a greater amount of a reactant is in contact with the other reactant. This means that more collisions can occur and the reaction can be faster.
- 8 Ben's tea would have been sweeter. Stirring would promote mixing of the sugar and water particles, and therefore increase the rate of the dissolving process.

ANALYSING

- 9 Tube 5 would have the fastest reaction because it has a high concentration, high temperature and small particle size, which all combine to result in a faster reaction.
- 10 a Curve for reaction B has a steeper slope than A at the beginning, so this reaction occurs faster. Temperature affects the rate at which a reaction occurs.
- b Temperature for reaction B is higher because the reaction occurs at a faster rate and there are more moles of CO (product) formed in the same time as in reaction A.

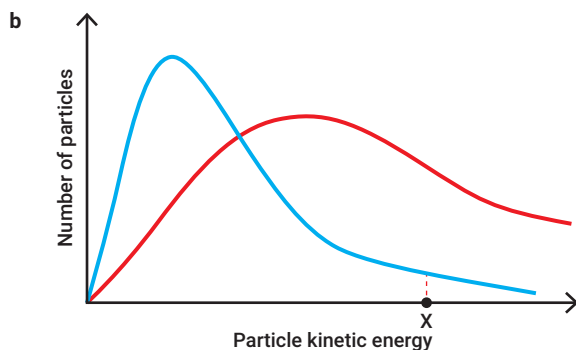
CHAPTER EXAM

MULTIPLE CHOICE

1 C	3 C	5 D	7 B	9 A
2 B	4 B	6 C	8 C	10 B

SHORT RESPONSE

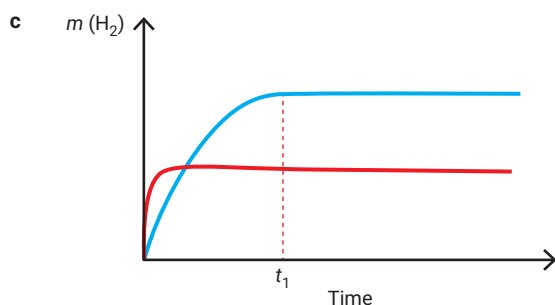
- 11 a Minimum amount of energy required for a reaction to occur



- c Significantly more particles have an energy greater than the activation energy.

- 12 a At the start

- b No more hydrogen gas is being produced so the reaction is over.



- d Any two of: higher temperature, higher concentration of HCl, use of a catalyst

CROSS-CHAPTER QUESTION

- 13 a Initially, $[\text{H}_2\text{O}_2] = 0.010 \text{ mol L}^{-1}$, $V = 200 \text{ mL}$

$$n(\text{H}_2\text{O}_2) = cV = 0.010 \times 0.2 = 0.002 \text{ mol}$$

$$\text{After } 300 \text{ s, } [\text{H}_2\text{O}_2] = 0.004 \text{ mol L}^{-1}, V = 200 \text{ mL}$$

$$n(\text{H}_2\text{O}_2) = cV = 0.004 \times 0.2 = 0.0008 \text{ mol}$$

$$n(\text{H}_2\text{O}_2)_{\text{reacted}} = 0.002 - 0.0008 = 0.0012 \text{ mol}$$

$$n(\text{O}_2)_{\text{produced}} = \frac{0.0012}{2} = 0.0006 \text{ mol}$$

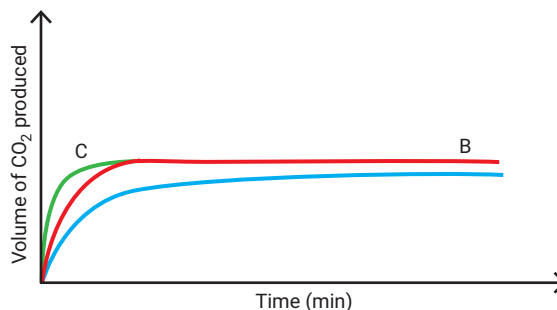
$$m(\text{O}_2) = n \times M = 0.0006 \times 32 = 0.0192 \text{ g}$$

- b Initial mass $\text{O}_2 = 0 \text{ g}$

$$\text{Rate of reaction} = \frac{0.0192 - 0}{300} = 6.4 \times 10^{-5} \text{ g s}^{-1}$$

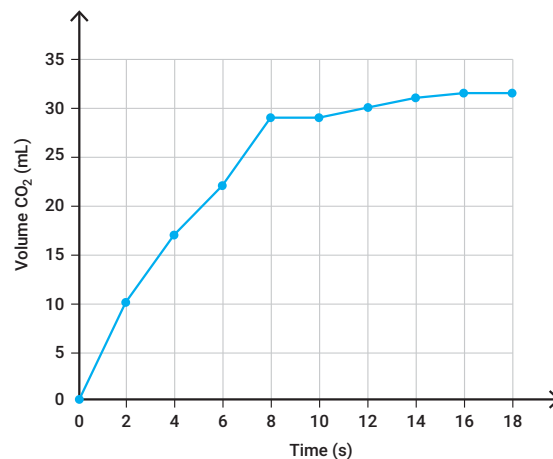
DATA ANALYSIS

- 14 a i–ii



- b Flask E: It has the same mass as A, but powder instead of lump, same temperature.
- c They have an energy level lower than the required activation energy for that reaction.

- 15 a



- b Reaction rate decreases as the reaction proceeds, which can be seen by the decrease in slope of the graph. The rate decreases because the reactants are being used up and there are fewer reactant particles to collide to produce the CO_2 product.
- c Any of:
- grinding up the sodium hydrogen carbonate, producing a greater surface area
 - increasing the concentration of hydrochloric acid so there are more particles available for collision
 - increasing the temperature, causing more particles to have enough energy to overcome activation energy barrier so more successful collisions occur between particles.

GLOSSARY

A

absolute uncertainty the magnitude of the difference between the observed/measured value and the true value

absolute zero precisely 0 K on the kelvin scale, which is a thermodynamic (absolute) temperature scale and -273.15°C on the Celsius scale

absorbance value a measurement of the relative amount of radiation that is absorbed by a sample

absorption spectrum a spectrum showing dark lines or bands due to specific wavelengths being absorbed as electrons move from ground state energy levels to higher energy levels

accepted value the value of a substance or quantity that is universally agreed as being a best estimate due to multiple and highly accurate measurements

accuracy the degree to which the result of a measurement, calculation or specification conforms to the correct value or a standard

acid a substance that produces hydrogen ions (H^+) in aqueous solution

actinoid a metallic element of atomic number 89–103 that along with the lanthanides make up the f block of the periodic table

activated complex the intermediate state of a chemical reaction in which bonds are breaking and forming

activation energy (E_a) the minimum energy required for a chemical reaction to occur

addition reaction a reaction in which atoms are added to a molecule by breaking a double or triple bond

adhesive force a force of attraction between unlike molecules

adsorb to be attracted to the surface of a material; the opposite of desorb

alkali/base a substance that reacts with acidic substances to form water as one of the products of the reaction

allotrope one of two or more forms of the same element, with different physical properties, e.g. diamond and graphite are both different forms of carbon

alloy a mixture of two or more metals, which has properties of each of the constituent metals

alpha particle a nucleus of a helium atom that is emitted by some radioactive substances as alpha radiation

amphoteric a substance that reacts with both alkaline and acidic solutions, so it is classified as neither alkali nor acid

anhydrous containing no water, particularly of ionic solids

anion a non-metallic atom that has become negatively charged by gaining one or more electrons

aromatic hydrocarbon a hydrocarbon containing one or more benzene rings

atom the smallest stable unit of matter that can exist by itself

atomic mass unit (amu) one-twelfth of the mass of an atom of carbon-12

atomic number (Z) the number of protons in one atom of an element – each element has a unique atomic number

atomic orbital a region of three-dimensional space around an atom's nucleus containing a maximum of two electrons

atomic radius a measure of the size of an atom – half the distance between the nuclei of two neighbouring atoms of the same element – the distances are very small and are measured in picometres (pm) ($1\text{ pm} = 1 \times 10^{-12}\text{ m}$)

atomise heating a substance until it exists as a vapour of gaseous atoms

Aufbau principle electrons around an atom will fill the lowest available energy level first, before filling higher energy levels

Avogadro constant (N_A) the number of units of one mole of substance, approximated as 6.022×10^{23}

Avogadro's hypothesis equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles

B

base a substance that either contains the oxide ion (O^{2-}) or the hydroxide ion (OH^-), or produces hydroxide ions in an aqueous solution

beta particle an electron or positron (which has the same mass as an electron but the opposite charge) emitted when an unstable nucleus decays

Bohr model a model of the atom that considers electrons as particles orbiting the nucleus in specific energy levels

boiling point the temperature at which bubbles of gaseous vapour are formed within the liquid

bonding pair a pair of electrons shared between two atoms

brittle a substance that will shatter, rather than bend, when a force is applied

C

calibration curve a graph of absorbance values of a series of samples of known concentrations used to determine the unknown concentration of a sample

calorimetry the process of measuring the amount of heat released or absorbed during a chemical reaction

capillary action the ability of a liquid to flow through narrow spaces in opposition to gravity

carbon dating a method used to determine the age of a once-living object by comparing the ratio of carbon-12 to carbon-14 in the object

catalyst a substance that affects the rate of certain reactions by providing an alternative reaction pathway

cation a metallic atom that has become positively charged by losing one or more electrons

chemical bond a force of attraction between atoms that holds them together to form compounds or molecules

chemical bond energy the amount of energy required to break 1 mole of the stated bond to give separated atoms, usually expressed in units of kJ mol^{-1} ; also known as bond enthalpy

chemical property a property of a substance that relates to the ways in which it bonds or reacts with other substances, e.g. acidity and flammability

chemical reaction the formation of new substances and the associated energy transformations

chromatography a group of techniques that separate substances based upon differential distribution between a stationary phase and a mobile phase

cohesive force a force of attraction between molecules in a substance that helps to maintain a certain shape of the liquid

collision theory for a reaction to occur, the particles must collide with sufficient energy and in the required orientation

combination reaction a reaction in which two or more reactants combine to form a single product

combustion an exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that releases energy

combustion reaction an exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that produces energy

complete ionic equation shows all of the soluble ions (cations and anions) present in a solution as well as the formula of any precipitate formed

compound a substance consisting of atoms of two or more elements chemically bonded together in a fixed ratio

concentrated solution a large amount of solute dissolved in a given volume of solvent

concentration a measure of the amount of solute dissolved in a given amount of solvent

confounding variable a variable that is related to the independent and dependent variables

covalent bonding bonding between two or more atoms where electron pairs are shared between the valence shells

covalent molecular substance a substance composed of discrete molecules whose atoms are held together by intramolecular bonds

D

decomposition reaction a reaction in which a single substance breaks down into smaller, simpler substances

delocalised electrons electrons that are not held in a fixed orbit around one nucleus, but are shared by many atoms

density the mass per unit volume of a substance

dependent variable the variable that changes due to changes to the independent variable

desorb the action of a substance moving from the stationary phase to the mobile phase

dilute solution a small amount of solute dissolved in a given volume of solvent

dilution adding water to a solution to reduce the concentration

dipole oppositely charged poles at each end of a bond or molecule

dipole–dipole attraction the attraction between molecules with permanent dipoles

dispersion force a weak intermolecular force that arises from electrostatic attractions between instantaneous dipoles in neighbouring molecules

dissociated when the ions from an ionic compound separate from each other and form a solution in water

distillation a separation technique of a liquid mixture based on boiling point; the mixture is heated and the vapour produced is condensed

double covalent bond a covalent bond consisting of two shared pairs of electrons

double-displacement reaction a reaction in which a positive ion from one ionic compound exchanges with the positive ion of another ionic compound

ductile can be drawn out to form a thin wire

E

elastic collision a collision with no net loss of energy

electrolyte a solution that can conduct electricity because of the presence of ions

electrolytic solution a solution that contains ions, so it will conduct electricity

electron a subatomic particle with a negative charge that is present in all atoms

electron affinity the ability of an atom in the gaseous state to accept an electron and form a negative ion

electron configuration the specific arrangement of electrons within an individual atom

electron energy levels a region of an atom where electrons have the same energy

electronegativity the ability of an atom to attract electrons, while being bonded to another atom

electrostatic attraction a force that occurs between particles with opposite charges

element a substance consisting of only one type of atom

emission spectrum a spectrum showing coloured lines or bands due to specific wavelengths of electromagnetic radiation being emitted as excited electrons move from higher energy levels back to their ground state levels

empirical formula a chemical formula showing the simplest ratio of elements in a compound rather than the total number of atoms in the compound

endothermic reaction a reaction in which energy is absorbed from the surroundings, causing temperature to decrease

energy profile diagram a diagram showing relative enthalpies (energies) during the progress of a reaction

enthalpy (H) the total energy content of a chemical substance

enthalpy change (ΔH) the heat gained or lost per mole of reactant in a chemical reaction:
enthalpy change = (enthalpy of products) – (enthalpy of reactants)

enzyme a biological catalyst

error the difference between a measured value and true value

excess reactant the reactant that it is not totally consumed when a chemical reaction is complete

excited state an altered electron configuration from the ground state that occurs when electrons absorb additional energy

exothermic reaction a reaction in which energy is released to the surroundings, causing temperature to increase

experimental yield the actual yield (mass) of a product obtained for a reaction conducted in the laboratory

extraneous variable any variable that is not directly related to the experiment but could affect the results of the experiment

F

first ionisation energy the energy required to remove a single electron from an atom in the gaseous state

formula unit a chemical formula showing the overall ratio of positive and negative ions in an ionic compound

G

gas a substance (element, compound or mixture) that exists in a state with no defined volume or shape

gas pressure the force per unit area exerted by gas particles as they collide with the walls of their container

giant covalent network substance a compound made up of carbon or silicon whose structure is a three-dimensional network of covalently bonded atoms

gradient the steepness of the line on a graph, which can be used as a measure of the rate of change of the quantity measured on the y-axis (the dependent variable) relative to that on the x-axis (the independent variable)

gravity filtration a technique that uses gravity to separate a solid–liquid mixture by passing the mixture through filter paper

ground state the lowest energy electron configuration of an atom

group a vertical column on the periodic table

H

heat the thermal energy transferred from a hotter system to a cooler system

heat of combustion (ΔH_c) the heat released when a fuel undergoes complete combustion at standard atmospheric pressure

heterogeneous mixture a mixture in which substances are non-uniformly distributed throughout

homogeneous mixture a mixture in which substances are uniformly distributed throughout

Hund's rule of maximum multiplicity every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied

hydrated ionic crystals with a specific number of water molecules in their chemical formulas

hydrocarbon a compound containing only hydrogen and carbon

hydrogen bond when a H atom that is bonded to an O, N or F atom in one molecule is attracted to the lone pair of electrons of N, O or F of an adjacent molecule

hydrophobic 'water-hating', not attracted to the water molecule

I

ideal gas a gas that perfectly obeys all the proposals of the kinetic theory of gases and the gas laws

independent variable the variable that is purposely changed or manipulated in an experiment

insoluble incapable of being dissolved or very poorly soluble in a solvent

instrumental uncertainty the inherent limitations and potential errors associated with the measuring instruments or tools used in scientific experiments or observations

intermolecular bonding bonds that exist between molecules; much weaker than intramolecular bonds but important in determining the physical properties of the molecule such as melting and boiling points and solubility

intermolecular force an attractive force that causes molecules to aggregate to form liquids and solids

interpolation to read or construct a new data point that has not been measured but is within the range of measured data

intramolecular bonding bonds that occur between atoms within a molecule

ion an atom that has become charged through the gain or loss of one or more electrons

ionic bond the mutual, non-directional, electrostatic attraction of positively and negatively charged ions, which holds them together in a lattice in a compound

ionic compound a compound consisting of atoms that have lost and gained electrons to form charged ions that are electrostatically attracted to one another

ionic radius the distance between the centre of the nucleus and the outermost electron shell in an ion

ion-dipole bond a bond formed by attraction between a charged ion and the dipole of a polar molecule

isotopes two or more atoms of the same element with the same number of protons but with different masses, resulting from their different number of neutrons

isotopic composition a list of the different isotopes present in a sample of an element and the relative amount of each isotope that is present

IUPAC the International Union of Pure and Applied Chemistry

K

kinetic energy the energy possessed by moving particles

kinetic theory a model used to explain the behaviour of gases based on their movement

L

lanthanoid a metallic element of atomic number 57–71 that along with the actinoids make up the f block of the periodic table

latent heat of fusion the energy required to change a specific quantity of a substance from a solid to a liquid

latent heat of vaporisation the energy required to change a specific quantity of a substance from liquid to a gas

lattice a regular repeating three-dimensional arrangement of ions, which results in the formation of crystals

Lewis structure a representation of the electron arrangement of atoms, using dots or crosses to represent electrons in their shells; also referred to as a Lewis dot structure, electron dot structure or electron dot formula

limewater an aqueous solution of calcium hydroxide ($\text{Ca}(\text{OH})_2$)

limiting reactant the reactant that it is totally consumed when a chemical reaction is complete

line of best fit a straight line through data points in a graph that best expresses the relationship shown in a scatterplot

line spectrum a pattern of lines showing the wavelengths that have been absorbed or emitted

logbook a complete, permanent record of how an experiment or research project was conducted; it shows what was done at every step along the way

M

malleable can be repeatedly deformed, or hammered, without shattering

mass number (A) the number of protons and neutrons in the nucleus of an atom

mass spectrometry a technique used to determine the atomic or molecular composition of a sample and its isotopic composition

mass spectrum a graph of the information from a mass spectrometer, showing the number of particles of different mass (or, more precisely, mass-to-charge ratio) that are present in a sample and the relative abundance of each

maximum trendline a trendline with the greatest gradient that fits within the data within the uncertainty values

Maxwell-Boltzmann distribution a frequency distribution of the energy value of a sample of particles at a particular temperature

mean the average value of a set of values

meniscus the curved shape formed by the surface of a liquid in contact with a cylinder or tube

metallic bond the electrostatic force of attraction between cations and the delocalised electrons that are free to move within the regular lattice structure of a metal

minimum trendline a trendline with the smallest gradient that fits within the data within the uncertainty values

mixture a combination of two or more substances

mobile phase the fluid used to carry the components of a mixture up the stationary phase so they can separate

molar mass (M) mass of 1 mol of a substance (equals mass of $N_A = 6.022 \times 10^{23}$ atoms, molecules or formula units of a substance)

molar volume of a gas the volume occupied by 1 mole of gas at a known temperature and pressure

molarity the number of moles of solute per litre of solution

mole (n) the amount of a substance (g) containing $N_A = 6.022 \times 10^{23}$ particles (atoms, ions or molecules) of that substance

molecular formula a chemical formula that gives the total number of atoms of each element in a molecule

molecule a group of atoms covalently bonded together, representing the smallest unit of an element or a compound that can take part in a chemical reaction

monatomic ion an ion consisting of one atom

monochromator a filter used to ensure that only radiation of a specific wavelength is shone at a sample

multiple covalent bond a bond formed when two atoms share more than one pair of electrons; a double bond consists of two shared pairs of electrons and a triple bond of three shared pairs

N

nanoparticle a particle with a diameter less than 100 nm that behaves as a whole unit

net ionic equation represents only the chemical species that are involved in the precipitate reaction, with spectator ions eliminated

neutral neither acidic nor basic

neutral species equation a balanced chemical equation that gives the complete neutral chemical formulas of reactants and products without breaking them down into individual ions

neutralisation the reaction between an acid and a base to form a salt and water

neutron an uncharged particle that exists in the nuclei of all atoms except some hydrogen atoms

noble gas an element from group 18 on the periodic table

noble gas electron configuration an arrangement of electrons that is identical to one of the elements in group 18, known as the noble gases

non-bonding or lone pair electrons a pair of electrons located on an individual atom, which are not shared

non-electrolyte a substance that does not form ions when it dissolves in water, so the resulting solution does not conduct electricity

non-polar describes bonds or molecules that have relatively equal sharing of electrons

non-spontaneous chemical reaction a reaction that can only occur with the input of additional energy such as heat, light or electricity

nucleus the positively charged centre of an atom, containing positively charged protons and electrically neutral neutrons, to which the orbiting electrons are attracted

O

organic chemistry the study of carbon-based compounds

oxide a compound that forms when an element combines with oxygen

P

pascal the SI unit of gas pressure, equivalent to a force of 1 newton per square metre; symbol Pa

Pauli exclusion principle no two electrons can occupy the same quantum state; hence, two electrons in the same orbital will have opposing spins

Pearson correlation coefficient (R) a statistical measure that quantifies the direction and strength of a relationship between two variables

percentage composition the percentage by mass of each of the different elements in the compound

percentage error the difference between a measurement result and an accepted value, expressed as a percentage of the accepted value

percentage uncertainty a measure of the uncertainty of a measurement compared with the size of the measurement, given as a percentage

period a horizontal row on the periodic table

periodic table of elements a table listing the known elements in order of atomic number, based on their chemical and structural properties, so that patterns and trends are easily recognisable

pH a scale used to measure the acidity of a solution: pH = 1 is strongly acidic; pH = 7 is neutral; pH = 14 is strongly alkaline

phase change the transformation of a substance between solid, liquid or gaseous forms

photon a particle representing a specific quantity of energy

physical property a property of a substance that does not involve a chemical reaction, e.g. melting point and electrical conductivity

polar describes bonds or molecules that have uneven sharing of electrons

polarity a measure of the unequal sharing of electrons in a covalent bond; the result of the difference in electronegativity of bonded atoms

polyatomic ion a group of atoms that chemically bond together and have an overall charge, which can be positive or negative; they can take part in ionic bonding

pop test an audible test to determine the presence of hydrogen gas

precipitate an insoluble solid produced when two different solutions react

precipitation reaction a reaction between two solutions that produces a solid

precision the closeness of several independent measurements of the same quantity to each other

primary data data collected directly by a person or group

protein a compound consisting of chains of amino acid molecules; proteins are essential for life

proton a positively charged particle that exists in the nuclei of all atoms

pure substance composed of only one substance

Q

qualitative analysis a technique that determines the identity of a sample, rather than the quantity

qualitative data information that is not numerical in nature

quantitative analysis a technique that measures the amount of a substance present in a sample

quantitative data numerical information

R

radioactive decay the process by which an unstable nucleus changes into a more stable nucleus by releasing small particles and/or electromagnetic radiation

radioisotope a radioactive isotope of an element that is unstable and emits radiation as it decays

radiotherapy the use of radiation to kill cancer cells and shrink tumours

random error a variation that affects a measurement in a random way so that successive measured values may reflect small changes from each other

range the difference between the maximum and minimum values of a measured confidence interval

rate how quickly one quantity (a dependent quantity) changes compared to another (an independent quantity)

reactivity the likelihood of an element or substance undergoing a chemical reaction

redox reaction short for reduction–oxidation reaction, a type of reaction that involves transfer of electrons between two species; one reactant is oxidised (loses electrons) and the other reactant is reduced (gains electrons)

relative atomic mass (A_r) the average mass of one atom of an element, relative to the mass of a carbon-12 atom

relative formula mass (M_r) the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 atom is exactly 12

relative molecular mass (M_r) the mass of a molecular substance relative to the mass of a carbon-12 atom

reliability the extent to which the results of assessments are consistent, replicable and free from error

research question a question that directs the scientific inquiry activity; it focuses the research investigation or student experiment, informing the direction of the research, and guiding all stages of inquiry, analysis, interpretation and evaluation

retardation factor (R_f) the distance travelled by the sample from its origin divided by the distance moved by the solvent from the origin

S

salt a substance formed from the positive ion (cation) of a base and the negative ion (anion) of an acid

saturated compound an organic compound with only single bonds between carbon atoms

saturated solution a solution in which no more solute can dissolve in a given quantity of solvent at a given temperature

secondary data data that is collected by someone else
sieving a technique that separates particles in a mixture based on particle size

single-displacement reaction a reaction in which an element within a compound is displaced by another element

solubility the amount of solute that can dissolve in 100 g water (solvent) at a particular temperature (100 g water = 100 mL water)

solubility curve a graph showing the variation of the solubility of a substance in water with temperature

soluble capable of being dissolved in a solvent

solute a substance that is dissolved in a solvent to form a solution

solution a homogeneous mixture with a solute dissolved in a solvent, where the particles of the solute cannot be observed

solvent a substance (usually a liquid) in which another substance is dissolved

specific heat capacity (c) the heat energy required to raise the temperature of 1 gram of a substance by 1 kelvin

spectator ion an ion that remains in solution before and after a chemical reaction has taken place; it is not involved in the reaction

spectroscope a device used to separate light into its component wavelengths

spectroscopy a method of analysis performed by subjecting a sample to electromagnetic radiation and observing the frequencies absorbed or emitted

spontaneous chemical reaction a reaction that occurs without the input of additional energy such as heat, light or electricity

standard atmospheric pressure atmospheric pressure at sea level, described as 1 atm, or 100 kPa

stationary phase the immobile surface that different components of a mixture can adhere to when carried along by the mobile phase

stoichiometric ratio the exact proportion of reactants and products in a chemical reaction, based on the balanced chemical equation

stoichiometry the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers

sublimation the process of a solid turning into a gas when heated, without forming a liquid

subshell an energy sublevel made up of multiple orbitals that contain electrons with the same amount of energy

substitution reaction a reaction in which an atom or group of atoms takes the place of hydrogen on a parent carbon atom in a hydrocarbon molecule

successive ionisation energies the energies required to remove the electrons from an atom, in sequence, starting with the outermost

supernate the solution above a precipitate; composed of spectator ions in aqueous solution

supersaturated solution an unstable solution that has a higher amount of solute than the saturated solution at a given temperature

surface area the total area of all surfaces

surface tension the tension of the surface film of a liquid, due to the attraction between particles in the surface layer

suspension a mixture of a solution and an insoluble solid; the solid distributes evenly throughout the solution

systematic error an error that acts to give a consistent offset in data, e.g. consistently above or consistently below

T

temperature a measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

theoretical yield the amount (mass) of a product that is produced from a complete reaction of the limiting reactant

transition metal a metallic element from the d block of the periodic table

trendline a line that represents the general direction or pattern of data points

triple covalent bond a covalent bond consisting of three shared pairs of electrons

U

uncertainty the range of values for a measurement result, taking account of the likely values that could be attributed to the measurement result given the measurement equipment, procedure and environment

unsaturated compound an organic compound that has at least one double bond or triple bond between carbon atoms

unsaturated solution a solution with less than the maximum amount of solute dissolved in a given quantity of solvent at a given temperature

V

valence shell the outermost electron energy level of an atom

valency a measure of the number of chemical bonds an atom can form with other atoms

validity the extent to which the experiment measures what it is intended to measure

vapour pressure the pressure of a substance in the gas phase (vapour) that exists above a sample of liquid of the same substance in a closed container

W

water of crystallisation water molecules present in a fixed ratio that form an essential part of the crystal structure of some ionic compounds

wavelength a property of light related to the length of the wave, which can give light properties such as colour

INDEX

- absolute zero (0 K), 319, 322
- absorbance value, 79
- absorption spectroscopy, 77–8
 - atomic, 78–82
- absorption spectrum, 77
- acid–base (neutralisation) reactions, 180, 417–19
- acid–carbonate reactions, 421–5
- acid–metal reactions, 420
- acids, 58, 432
 - definitions of, 432–3
 - recognising strong and weak, 438–42
- acids, reactions of
 - acid–base, 180, 417–19
 - acid–carbonate, 421–5
 - acid–metal, 420
 - representing, 416
- actinoids, 16
- activated complex, 458
- activation energy, 460
 - and nature of reactants, 461
 - and rate of reaction, 461
- adhesive forces, 347
- adsorb, 290
- alkali metals, 55
- alkaline, 58
- alkanes, 148–9
 - naming, 149–50
 - properties of, 150–1
- alkenes, 151–2
 - naming, 152–3
 - properties of, 153–4
- allotropes, 145
- alloys, 93
- alpha particles, 6
- ammonia (NH_3), 271, 439
- amphoteric oxide, 58
- analytical techniques
 - absorption spectroscopy, 77–82
 - emission spectroscopy, 71–7
 - mass spectrometry, 68–71
- anhydrous, 347
- anions, 94–5, 371, 380–1, 383
- aqueous solutions and molarity
 - concentration, 350–4
 - dilutions and reacting quantities, 354–61
 - saturated solutions, 361–2
- solutes and solvents, 349–50
- structure and properties of water, 342–9
- aromatic hydrocarbons, 154
- Arrhenius model, 436–8
- Arrhenius, Svante, 436
- atmospheric pressure, 319
- atom
 - calculating valency, 46–7
 - electron configuration, 13–22
 - overview, 6–8
 - subatomic particles, 8–9
- atomic absorption spectroscopy, 78–82
- atomic mass unit (amu), 192
- atomic number, 9, 10
- atomic orbitals, 15–16
- atomic radius, 45–6
- atomic structure
 - atom, 6–9
 - electron configuration, 13–22
 - elements, 9–10
 - hydrogen emission spectrum, 75–7
 - periodic table of elements, 10
 - protons, neutrons and electrons, 10–12
- atomise, 79
- Aufbau principle, 14
- Avogadro constant (N_A), 192–3
 - to mole, 193
- Avogadro's hypothesis, 315
- balanced chemical equations, 173–5
 - combustion of methane, 175–7
- base, reactions of
 - acid–base, 180, 417–19
 - acid–carbonate, 421–5
 - acid–metal, 420
 - representing, 416
- bases
 - definitions of, 432–3
 - recognising strong and weak, 438–42
- benzene, 154
 - properties of, 154–5
- biofuels, 258–9
- blood plasma, 447–8
- Bohr model, 75
- Bohr, Niels, 13
- boiling point and intermolecular forces, 277–8
 - of water, 344–5
- bond angles, 266
- bonding
 - chemical, 92–3
 - comparing different substances, 157–60
 - covalent, 102–7
 - covalent compounds, 143–8
 - hydrocarbons, 148–57
 - importance, 277–80
 - ionic, 94–101
 - ionic compounds, 136–8
 - metallic, 101–2
 - metals, 139–42
- bonding pair, 102
- Boyle, Robert, 9, 320
- Boyle's law, 320–2
- branched-chain compounds, 150
- brittle, 136
- bromine water, 154
- Buseck, Peter, 166
- calcium carbonate (CaCO_3), 178, 353, 421, 378, 453
- calculation of reacting quantities, 359–61
- calibration curve, 79
- calorimetry, 238, 240–1
- capillary action, 291
 - of water, 345–7
- carbon
 - allotropes, 146–8
 - fullerenes, 147
 - graphite, 146–7
- carbon dioxide (CO_2), 105, 120, 144, 171, 175, 178, 197, 208, 270–1
 - and Hess's law, 247
 - from combustion reaction, 241
 - limewater test, 378, 421
 - reaction with water, 375
 - solubility curve, 399
- carbonates, test for, 378–82
- catalysts, 470, 483–4
- cations, 94–5, 371, 381
 - in metallic bonding, 101–2
 - in metals, 140
- cellulose, 299–300
- Celsius temperature scale, 229

- Chadwick, James, 8
- change of state. *see* phase change
- Charles, Jacques, 322
- Charles' law, 322–4
- chemical bond energy, 235
- chemical bonding, 92–3
- chemical equations, 171
- chemical properties (elements), 31
- chemical reactions, 170
- activation energy, 460
 - balanced chemical equations, 173–7
 - chemical equations, 171
 - collision theory, 458, 462
 - energy profile diagram, 459
 - Maxwell–Boltzmann distribution, 460–1
 - phase change, 171–2
 - physical and chemical change, 170–2
 - rate, 452–62
 - spontaneity, 171
 - types, 177–83
- chromatogram, 292
- chromatography
- investigation of paper, 293–4
 - investigation of thin-layer, 294–7
 - paper and thin-layer, 290–3
 - separating components, 297–302
- chromium, electron configurations for, 21–2
- cohesive forces, 347
- collision theory, 458, 462
- combination reactions, 177
- combined gas law, 324–5
- combustion, 151
- of methane, 175–7
 - reactions, 178, 241
- complete ionic equations, 372
- compounds and mixtures
- chemical bonding, 92–3
 - classifying substances, 118–19
 - heterogeneous and homogeneous mixtures, 120–5
 - pure substances and mixtures, 126–7
- concentrated solution, 350–4, 434
- concentration, 452
- rate of reaction, 462–5
 - of solution, 350–4, 434
- conductivity, 439
- of metals, 139–40
- copper, electron configurations for, 21–2
- covalent bonding, 102–3
- multiple covalent bond, 104–5
 - names and formulas, 105
 - periodic table, 103–4
 - polarity, 269
 - representation, 103
- covalent compounds, 143
- allotropes of carbon, 146–8
 - giant covalent network substances, 145–6
 - molecular substances, 143–4
- covalent molecular substances, 143–4
- Curie, Marie, 114
- Dalton, John, 6
- decomposition reactions, 178
- deflection, 69
- delocalised electrons, 101, 139
- density, 344
- density of ice, 345
- deposition, 171
- desorb, 290
- dilute solution, 434
- dilutions, 356–9
- dipole–dipole attractions, 276, 344
- dispersion forces, 275–6, 344
- dissociated, 375
- distillation, 122
- double covalent bond, 104
- double-displacement reactions, 180
- acid–base (neutralisation), 180
 - gas-forming, 181
 - precipitation, 180
- ductility, 54
- of metals, 140–1
- elastic collisions, 318
- electrolytes, 391
- electrolytic solution, 375
- electron affinity, 54
- electron configuration
- atomic orbitals and subshells, 15–16
 - atoms, 13–14
 - chromium and copper, 21–2
 - determining, 14
 - elements in period 4, 19–20
 - ions, 20–1
 - orbital notation, 18–19
 - periodic table, 16–17
 - simple way of writing, 17
 - stable, 44
- electron energy levels, 13
- electronegativity, 52–3, 268–70
- electrons, 6, 10–12
- cyclic delocalisation of, 154
 - movement of, 71–2
- electrostatic attraction, 9, 95
- elements, 9–10
- chemical bonding, 92–3
 - period 3, 57–8
 - period 4, 19–20
- emission spectroscopy, 71
- analysing emission spectra, 74–5
 - flame tests, 73–4
 - hydrogen emission spectrum and atomic structure, 75–7
 - movement of electrons, 71–2
- empirical formula, 95, 201
- determining, 202–4
- endothermic reactions
- energy and temperature, 228–30
 - energy change calculation, 240–6
 - enthalpy change calculation 247–50
 - Hess's law, 247–50
 - law of conservation of energy, 230–7
 - temperature and enthalpy, 238–40
- energy
- law of conservation, 230–7
 - temperature, 228–30
- energy change
- calculation, 240–1
 - combustion reactions, 241
 - heat of combustion (ΔH_c°), 242–5
 - heat of neutralisation, 245–6
- energy profile diagram, 459
- enthalpy (H), 232, 459
- change calculations, 247–51
 - temperature, 238–40
- enthalpy change (ΔH), 232
- level diagrams, 233
- ethanamine (ethylamine, C_2H_7N), 439
- ethanoic (acetic) acid, 437
- evaporation, 279
- excess reactant(s), 212
- excited state, 72
- exothermic reactions, 231
- energy and temperature, 228–30
 - energy change calculation, 240–6
 - enthalpy change calculation, 247–51
 - Hess's law, 247–51
 - law of conservation of energy, 230–7
 - temperature and enthalpy, 238–40

experimental yield, 213
 first ionisation energy, 47–50
 flame tests, 73–4
 fluorine, 55
 formula unit, 95
 fossil fuels, 258–9
 fractional yield, 213
 Fuller, Buckminster, 147
 fullerenes, 147
 gas chromatography, 290
 gas-forming reactions, 181
 gas laws
 Boyle's law, 320–2
 Charles' law, 322–4
 combined gas law, 324–5
 ideal gas equation, 325–7
 gas pressure, 462
 gases, 314–15
 kinetic theory, 317–28
 relationship between number of moles of gas and volume, 315–17
 solubility, 399–401
 Geiger, Hans, 6
 Geiger–Marsden experiment, 7
 giant covalent network substances, 145–6
 gradient, 454
 graphite, 146–7
 gravity filtration, 122
 ground state, 18, 71
 groups (periodic table), 10
 half-life of radioisotopes, 114
 halogenation, 151
 halogens, 55–6
 heat, 228
 heat of combustion (ΔH_c), 242–5
 heat of neutralisation, 245–6
 heat of reaction, 232
 Hess's law, 247–51
 heterogeneous mixtures, 120–1
 separating mixtures, 122–3
 homogeneous mixtures, 120–1
 separating mixtures, 122–3
 Hund's rule of maximum multiplicity, 18
 hydrated ionic crystals, 347
 hydrocarbons, 148
 alkanes, 148–51
 alkenes, 151–4
 benzene, 154–5

tests to distinguish alkanes and alkenes, 155–7
 hydrogen bonding, 276–7, 343–4
 hydrogen chloride (HCl), 269–70
 dipole–dipole attractions, 276
 hydrogen emission spectrum and atomic structure, 75–7
 hydrogen molecule (H_2), 268
 hydrophobic, 393
 ideal gas, 317
 ideal gas equation, 325–7
 immiscible liquids, 123
 intermolecular bonding, 144
 solubility and, 390–7
 intermolecular forces, 274, 344
 melting and boiling points, 277–8
 molecular structure, 264–8
 observable properties, 274–81
 polarity of molecules, 268–74
 solubility, 280
 vapour pressure, 278–9
 water, 344
 International Union of Pure and Applied Chemistry (IUPAC), 10, 97
 interpolation, 79
 intramolecular bonding, 144
 ion–dipole bonds, 391–2
 ionic bonding, 94–5
 formulas of, 98
 names of, 97–8
 periodic table and, 96–7
 ionic compounds, 136
 properties of, 137–8
 ionic radius, 47
 ionisation energy, 48–50
 ions, 94, 11
 electron configuration of, 20–1
 in solution, 370–82
 ions in solution
 determination of, 378–82
 solubility rules and writing precipitation equations, 370–8
 isotopes, 30–2
 chemical properties, 31
 physical properties, 31
 relative atomic mass, 32–4
 stable and unstable, 31
 isotopic composition, 68
 IUPAC. *see* International Union of Pure and Applied Chemistry (IUPAC)

kinetic energy, 228, 460
 kinetic theory, 317–18
 application, 320
 chemical reactions, 327–8
 gas laws, 320–7
 pressure, temperature and volume, 318–20
 lanthanoids, 16
 latent heat of fusion, 231
 latent heat of vaporisation, 231, 345
 lattice, 137
 Lavoisier, Antoine, 173
 law of conservation of energy, 230–7
 law of conservation of mass, 192
 in chemical equations, 207–9
 Lewis structures, 95
 limewater test, 378–82
 for carbon dioxide, 421
 limiting reactants, 212–13
 line spectrum, 71
 liquefied natural gas, 170
 lone pairs, 103
 lustre of metals, 140
 magnesium chloride, 95
 malleability of metals, 140–1
 malleable, 54
 Marsden, Ernest, 6
 mass
 chemical equations, 207–9
 law of conservation, 192
 moles, 198–200
 mass number (A), 11
 mass spectrometry, 68–71
 mass spectrum, 69
 mass–mass calculations, 211
 Maxwell–Boltzmann distribution, 460–1
 melting points, and intermolecular forces, 277–8
 Mendeleev, Dmitri, 42
 meniscus, 347
 metallic bonding, 101–2
 metallic character
 alkali metals, 55
 electron affinity, 54
 halogens, 55–6
 properties, 54
 metals, 139
 conductivity, 139–40
 lustre, 140

- malleability and ductility, 140–1
- properties, 141–2
- methane, 175–7
- mixtures, 119
- mobile phase, 290
- molar mass (M), 198
- molar volume of a gas, 315–16
- molarity, 350
- mole concept
 - Avogadro's number of particles, 192–6
 - chemical equations, 207–9
 - chemical formulas, 195
 - empirical and molecular formulas, 201–7
 - gas and volume, 315–17
 - law of conservation of mass, 192
 - mass, 198–200
 - percentage composition, 200–1
 - relative mass of substances, 196–8
 - stoichiometry, 209–20
- molecular formula, 105, 201
 - determining, 202–4
- molecular polarity, 391
- molecular size, 394–7
- molecules, 92
 - ammonia, 265
 - predicting shape, 264–6
 - water, 266
- mole–mole calculations, 209–11
- monatomic ions, 97
- monochromator, 79
- Moseley, Henry, 8
- multiple covalent bond, 104–5
- nanoparticles, 470
- nanotechnology, 166–7
- net ionic equation, 372–3
- neutral species equation, 372
- neutralisation reactions, 417–19
 - heat of, 245–6
- neutrons, 8, 10–12
- noble gases, 17
 - electron configuration, 101
- non-bonding pairs, 103
- non-electrolytes, 391
- non-metallic character
 - electron affinity, 54
 - halogens, 55–6
 - alkali metals, 55
 - properties, 54
- non-polar covalent bonds, 269
- non-polar polyatomic molecules, 270–2
- non-spontaneous chemical reactions, 171
- nucleus, 7, 8, 13
 - and atomic radius, 45
 - and isotopes, 31
- orbital notation, 18–19
- organic chemistry, 148
- oxides of elements in period 3, 57–8
- paper chromatography, 290–4
 - separating components, 297–302
- pascal (Pa), 319
- Pauli exclusion principle, 15, 18
- percentage composition, 200–1
- percentage yield, 213
- periodic table, 10
 - arrangement, 42–3
 - atomic radius, 45–6
 - covalent bonds, 103–4
 - electronegativity, 52–3
 - first ionisation energy, 47–50
 - groups and periods, 43–4
 - ionic radius, 47
 - ions, 96–7
 - metallic and non-metallic, 54–7
 - oxides of elements in period 3, 57–8
 - stable electron configurations, 44
 - successive ionisation energies, 50–2
 - valency, 46–7
- periods, 10, 43–4
- pH
 - Arrhenius model, 436–8
 - measuring, 434–5
 - scale, 432–6
 - recognising strong and weak acids and bases, 438–42
- phase change, 171–2
- photons of light, 72
- photosynthesis, 170
- physical properties (elements), 31
- plum pudding model, 6
- polar covalent bonds, 269
- polar polyatomic molecules, 270–2
- polarity of molecules, 268, 344
 - electronegativity, 268–70
 - polar and non-polar polyatomic molecules, 270–2
- polyatomic ions, 97
- pop test, 420
- precipitation reaction, 180, 370
- pressure, 399–401
 - kinetic theory, 319–20
 - rate of reaction, 462–5
 - temperature and volume, 320
- protons, 8, 10–12
- pure substances, 118
- quadruple covalent bond, 104
- qualitative analysis, 78
- quantitative analysis, 78
- quantum theory, 13
- quartz, 145
- radioactive decay, 31
- radioactivity, 114
- radioisotopes, 114
- rates of chemical reactions 452–62
 - factors that affect, 462–74
 - measuring, 453–6
- reacting quantities, calculation of, 359–61
- reaction equations, 416
- reactivity of elements, 43
- redox reactions, 181
- relative atomic mass (A_r), 11, 32–4
- relative formula mass, 197
- relative mass of substances (M_r), 196–8
- retardation factor (R_f), 291
- retention factor. *see* retardation factor (R_f)
- Rutherford, Ernest, 6
- salts, 417
- saturated compounds, 149
- saturated solutions, 361–2
- Schrödinger, Edwin, 76
- scuba diving, 337
- shielding effect, 48
- sieving, 122
- silica, 298
- silicon dioxide (SiO_2), 145
- single-displacement reactions, 178
- sodium bicarbonate, 422
- sodium chloride, 94
 - dissociation, 375
 - ionic compound, 197
 - Lewis structure, 95
 - properties, 138
 - reaction with silver nitrate, 180, 370, 371
 - solution in water, 391–2

- solubility, 350
 - effect of temperature, 398–401
 - equations, 375
 - intermolecular bonding, 390–7
 - intermolecular forces, 280
 - interpreting and analysing curves, 402–6
 - rules and writing precipitation equations, 370–8
- solutes, 349–50
- solutions, 120
 - of ionic compounds, 391–2
 - of molecular compounds, 392–4
- solvents, 347, 349–50
- Sorensen, Soren, 433
- specific heat capacity (*c*), 238–9
- spectator ions, 371
- spectroscopy, 71
- spontaneous chemical reactions, 171
- stable electron configurations, 44
- stable isotopes, 31
- standard atmospheric pressure, 242
- standard temperature and pressure (STP), 315
- stationary phase, 290
- steel, 142
- stoichiometric ratio, 209, 213
- stoichiometry, 208, 209
 - calculating yields, 214–20
 - experimental and theoretical yield, 213
 - limiting reactants, 212–13
 - mass–mass calculations, 211
 - mole–mole calculations, 209–11
- strength versus concentration, 439–42
- subatomic particles, 8–9
- sublimation, 55, 171
- subshells, 15–16
- successful collision, 458
 - increased number of, 468
 - increased percentage of, 468
- successive ionisation energies, 50–2
- supernate, 371
- supersaturated solution, 362
- surface area, 465–7
- surface tension, 344, 345–7
- suspension, 370
- temperature, 467–70
 - dependence of solubility of solids, 398
 - effect on solubility, 398–401
 - energy, 228–30, 460–1
 - enthalpy, 238–40
 - kinetic theory, 318–19
 - pressure and volume, 320
 - scales, 229
- tensile strength, 141
- tetrachloromethane, 271
- theoretical yield, 212, 213
- thermochemical equations, 231
- thin-layer chromatography (TLC), 290–3
 - investigation of, 294–7
 - separating components, 297–302
- Thomson, J.J., 6
- transition metals, 16, 97
- triple covalent bond, 104
- two-way paper chromatography, 294
- Tyndall effect, 121
- universal solvent. *see* water
- unsaturated compounds, 151
- unsaturated solution, 361
- unstable isotopes, 31
- valence shell, 46
- valence shell electron pair repulsion (VSEPR) theory, 264, 274
 - bond angles, 266
 - predicting shape of molecules, 264–6
- valency, 46–7, 93
- vapour pressure
 - and intermolecular forces, 278–9
 - of water, 344–5
- water
 - boiling point and vapour pressure, 344–5
 - density of ice, 345
 - intermolecular forces, 344
 - bonds, 343
 - polarity, 344
 - shape of molecule, 343
 - solubility of gases, 399
 - solvent ability, 347
 - structure and properties, 342–9
 - surface tension and capillary action, 345–7
- water of crystallisation, 347–9
- wavelength, 72
- WikiCell, 170
- yield
 - calculating, 214–20
 - experimental and theoretical, 213

Flexible online learning designed to support you

*Nelson MindTap puts **you** at the centre*

Access tools and content that make learning simpler yet smarter to help you go beyond drill and practice.



Watch video tutorials featuring expert teacher advice to unpack new concepts and develop your understanding.

Revise skills using quizzes, worksheets and skillsheets to practise your skills and build your confidence.

Navigate your own path, accessing the content, analytics and support you need whenever you need it in your learning journey.

Find everything you need to access your Nelson MindTap course at
cengage.com.au/nelsonmindtap