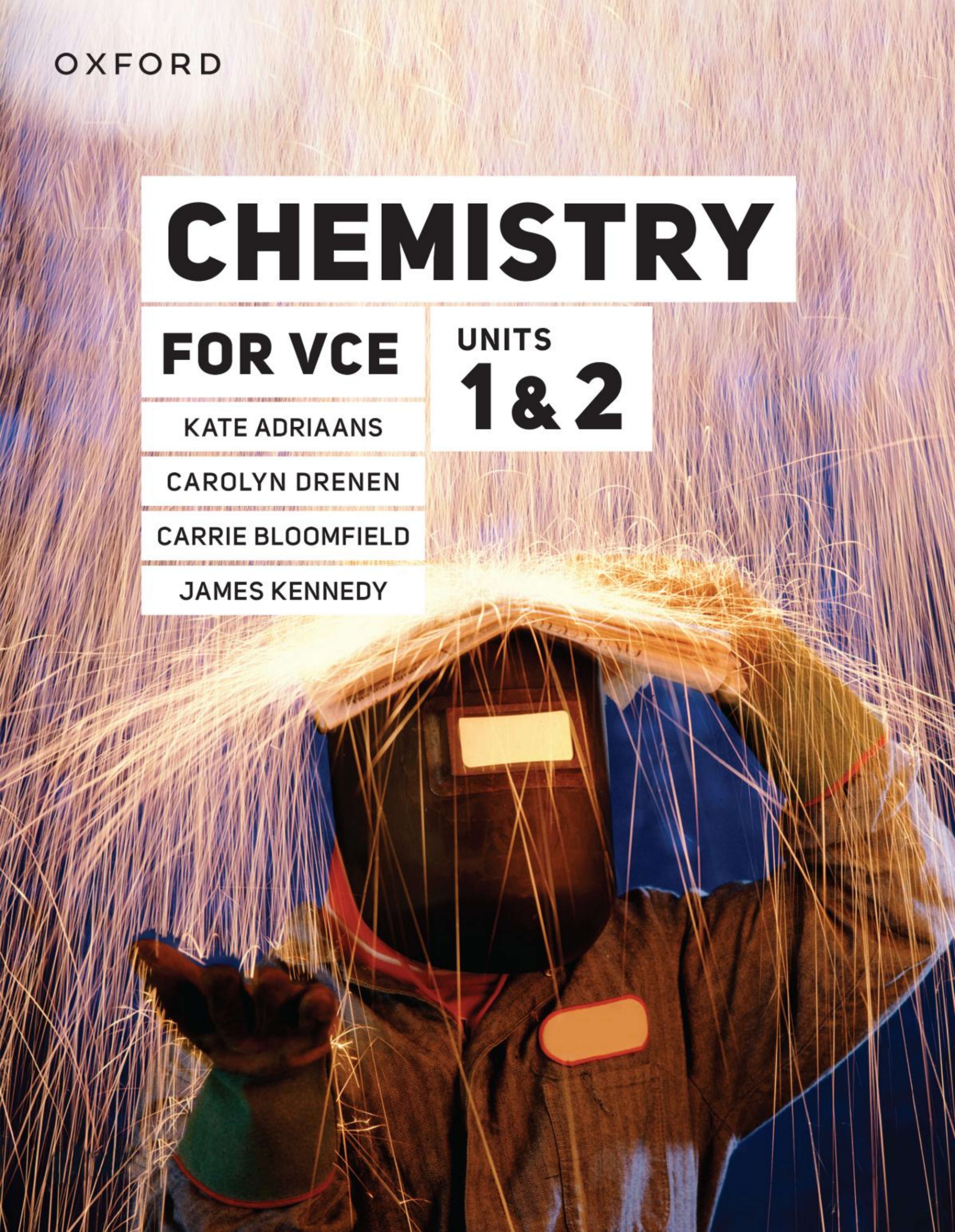

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UNITS 1 & 2

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CONTENTS

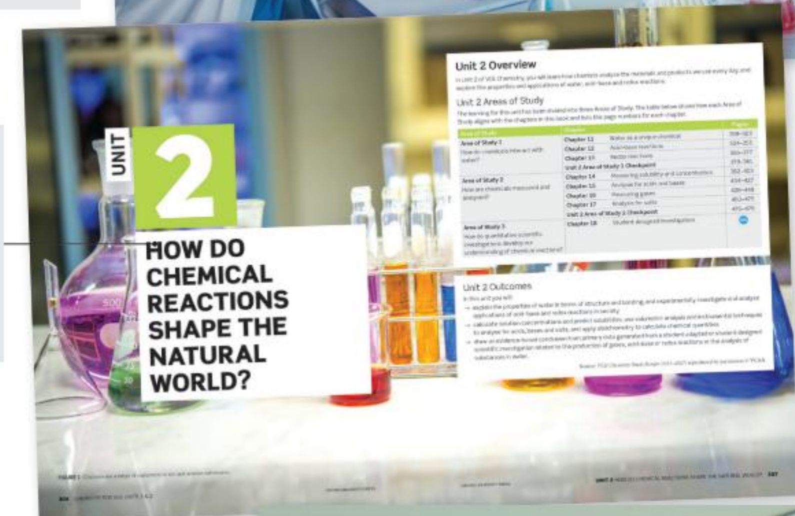
Using Chemistry for VCE Units 1 & 2

Key features of the Student Book:

» This Student Book combines complete coverage of the VCAA Chemistry Study Design 2023–2027 with clear and engaging design.
» Each print Student Book comes with complete access to all the digital resources available on Student obook pro.

Chemistry toolkit
The Student Book begins with a stand-alone reference chapter that includes:
- assessment advice and structured examples
- a step-by-step guide to preparing for your exam
- methods for presenting and analysing chemical data.

Unit openers
Each unit begins with a unit opener that includes:
- an **overview of Topics** in the Unit
- **Unit objectives** from the syllabus.



Groundwork
Questions for students to test their prior knowledge before they begin the chapter. Digital worksheets are available to help refresh their understanding.

Chapter openers
Each chapter begins with a chapter opener that includes key knowledge from the study design.

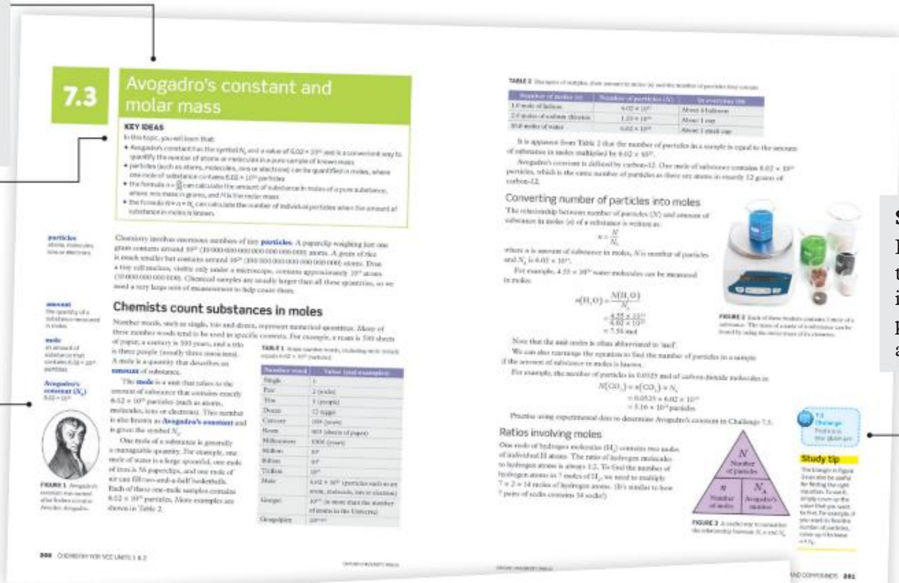
Practicals
A list of practicals within the chapter that allow students to apply their knowledge to real-life activities.

Topic-based approach
Content is presented in clearly structured topics. Each topic is labelled and numbered to help navigation.

Key ideas
Concept statements outlining the core content that students should take away from the topic.

Margin glossary
Definitions to all key terms.

Study tip
Practical advice to help students improve their performance in assessment tasks.

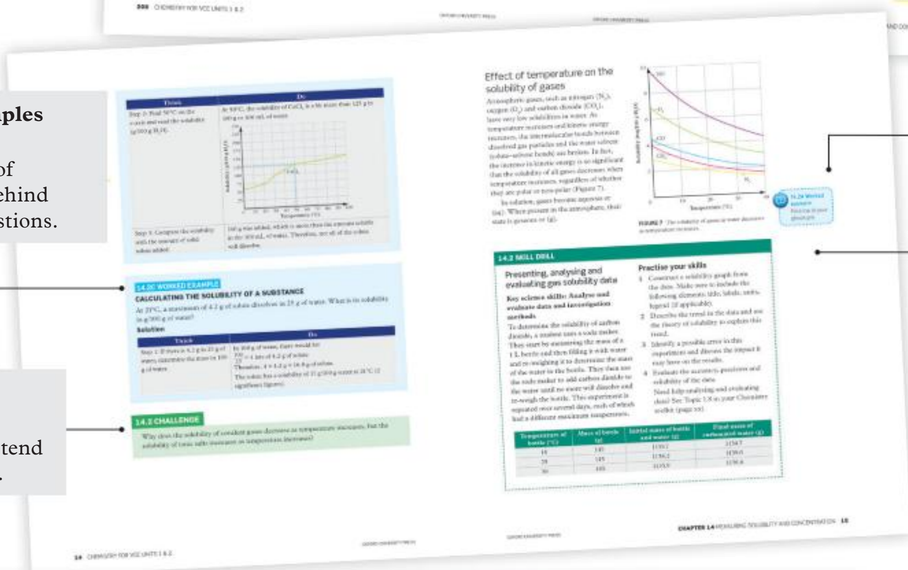


Worked examples
Step-by-step walkthroughs of the thinking behind answering questions.

Challenges
More difficult questions to extend understanding.

Digital only features
Dashed boxes for digital only features, accessible via the obook pro.

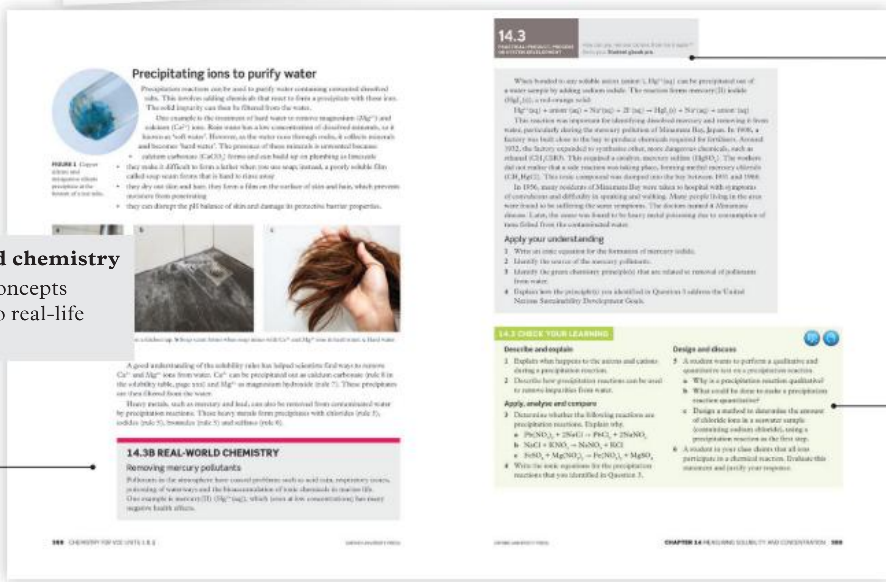
Skill drills
Students can practise their key science skills in context.



Practical links
Direct links to practicals.

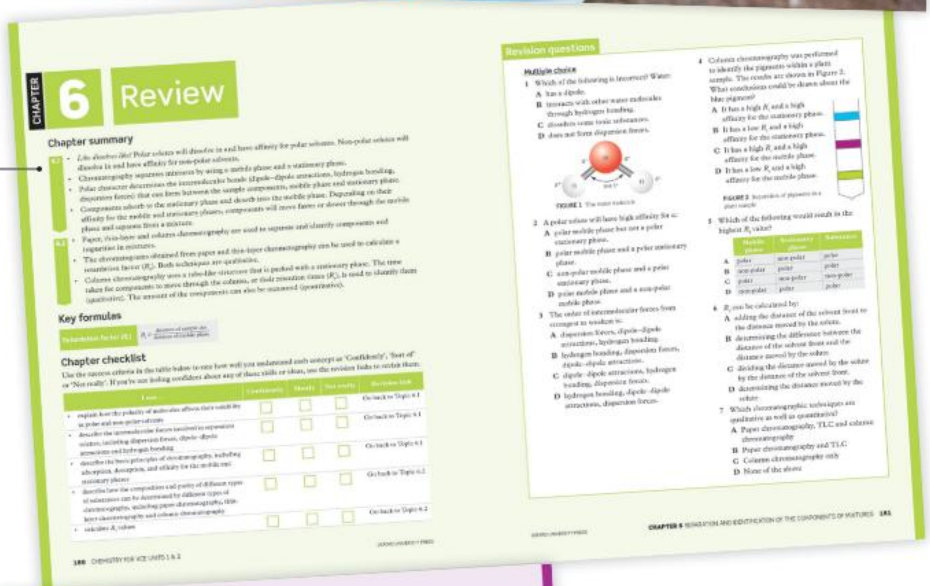
Real-world chemistry
Chemical concepts are linked to real-life examples.

Check your learning
Activity boxes with questions and tasks organised using cognitive verbs according to Bloom's taxonomy.



Practical manual

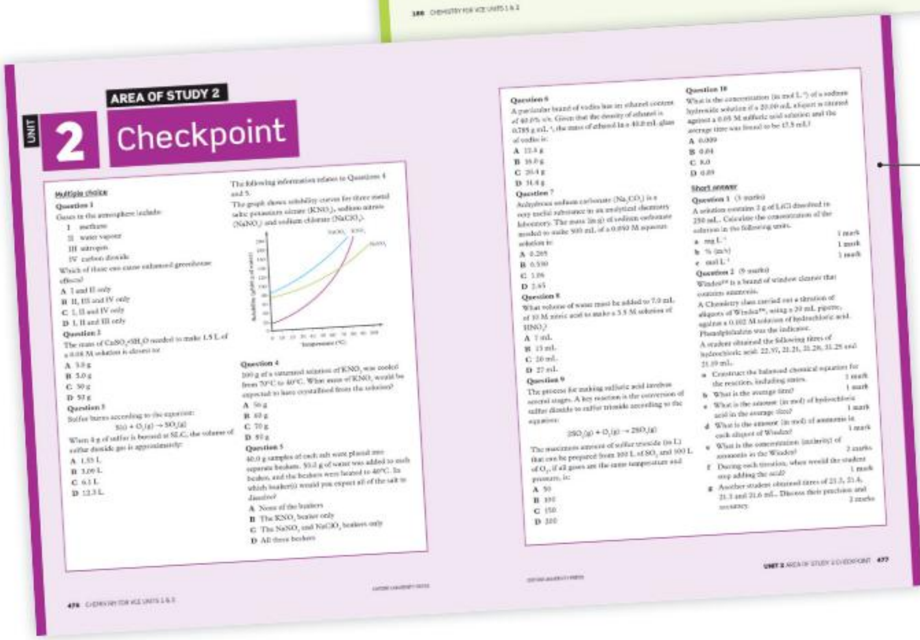
Each chapter contains a range of practicals that cover all eight scientific investigation methodologies. Practicals are available in the student book and via the obook pro.



Chapter reviews

Each chapter review includes:

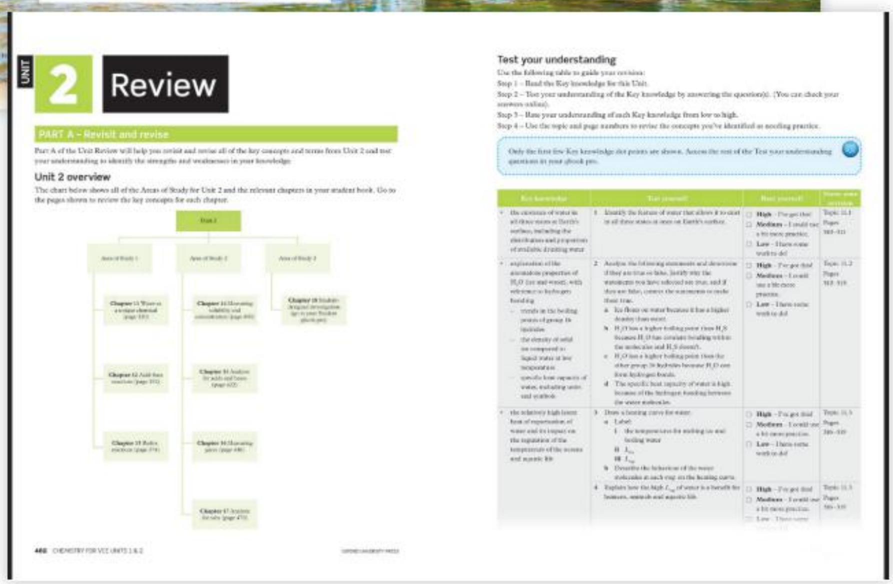
- a summary of key learning in each chapter
- key formulas relevant to the chapter
- a chapter checklist for students to evaluate their understanding of the key knowledge
- revision questions written to target assessment through multiple-choice and short-answer questions.



Checkpoints

Exam-style revision questions for each Area of Study, including multiple-choice and short-answer questions.

Online only chapters
Guidance for Unit 1 Outcome 3 Research investigation and Unit 2 Outcome 3 Student-designed investigation available via the obook pro.



Unit reviews
At the end of each unit, students can:
- revisit and revise the content with questions mapped to each key knowledge dot point
- gain tips to succeed in exams and apply their skills in a Think like an examiner activity
- answer practice exam questions.

Digital hotspots

Digital icons or hotspots found throughout the student book link to digital resources accessible via the obook pro.

Video – Watch a video demonstration for the Worked examples and Practicals

Assessment – Access a digital quiz for the topic or chapter

Resource – Access a worksheet or additional resource.

Key features of Student obook pro

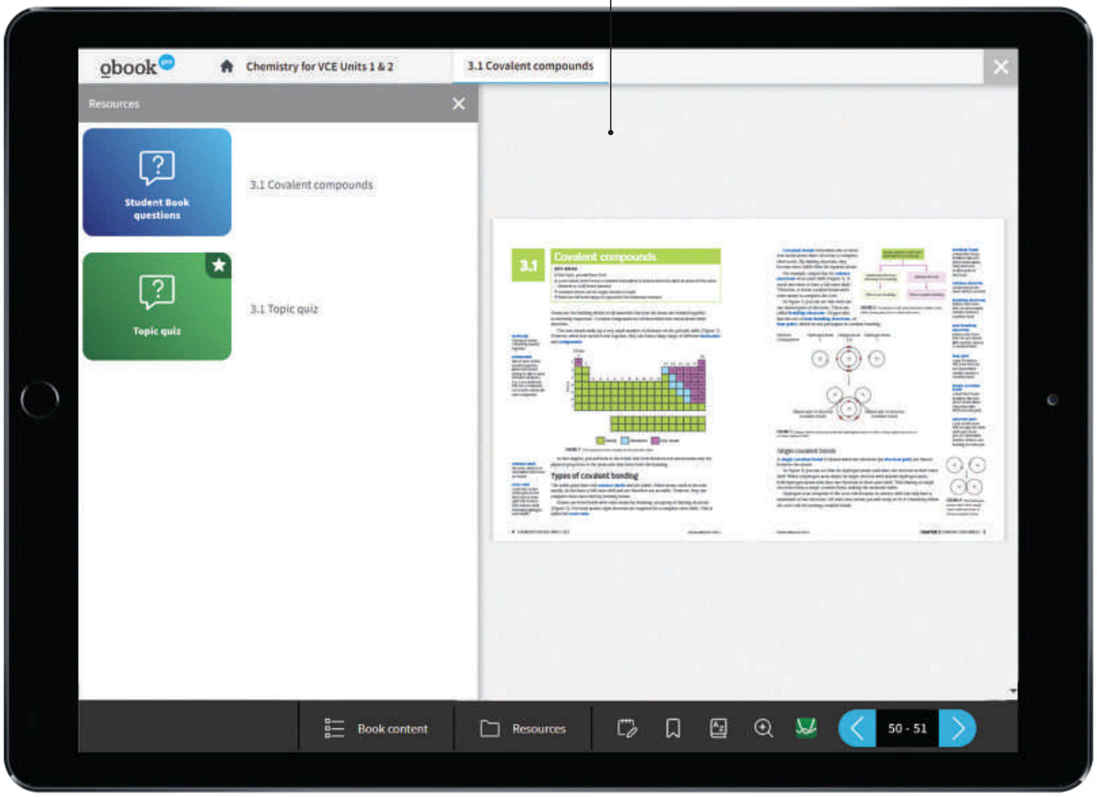
obook pro

> Student obook pro is a completely digital product delivered via Oxford's online learning platform, **Oxford Digital**.
> It offers a complete digital version of the Student Book with interactive note-taking, highlighting and bookmarking functionality, allowing students to revisit points of learning.
> A complete ePDF of the Student Book is also available for download for offline use and read-aloud functionality.

Focus on eLearning

Complete digital version of the Student Book

- This digital version of the Student Book is true to the print version, making it easy to navigate and transition between print and digital.



Interactive quizzes

- Each topic in the Student Book is accompanied by an interactive assessment that can be used to consolidate concepts and skills.
- These interactive quizzes are autocorrecting, with students receiving instant feedback on achievement and progress. Students can also access all their online assessment results to track their own progress and reflect on their learning.
- Each chapter is supported by a multiple choice quiz to give students further practice with exam-style questions.

Benefits for students

> integrated Australian Concise Oxford Dictionary look up feature
> targeted instructional videos for practicals and worked examples
> interactive assessments to consolidate understanding
> integrated Quizlet sets, including real-time online quizzes with live leaderboards
> access to their online assessment results to track their own progress.

obook pro

Key features of Teacher obook pro

> Teacher obook pro is a completely digital product delivered via **Oxford Digital**.

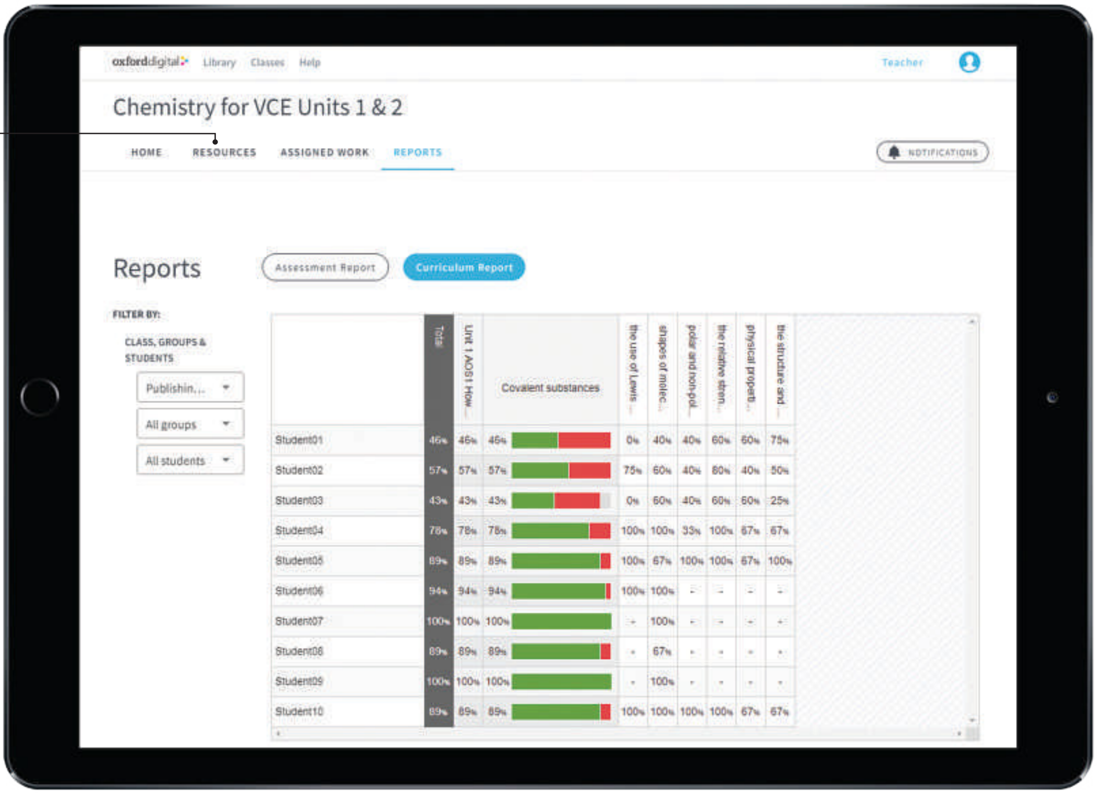
> Each chapter and topic of the Student Book is accompanied by full teaching support. Teacher notes are provided that clearly direct learning pathways throughout each chapter, including ideas for differentiation and practical activities.

> Teachers can use their Teacher obook pro to share notes and easily assign resources or assessments to students, including due dates and email notifications.

Focus on assessment and reporting

Complete teaching support

- Teaching support includes full lesson and assessment planning, ensuring there is more time to focus on students.



Additional resources

- Each chapter of the Student Book is accompanied by additional teaching and learning resources to help students progress.

Benefits for teachers

> In addition to online assessment, teachers have access to editable practice exams that are provided in topic 1.13 of the Chemistry toolkit. These exams are formatted like the VCAA Chemistry exam.

> Teachers are provided with laboratory support through experiment answer guidance, laboratory technician notes and risk assessments to ensure safe learning experiences.

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Kate Adriaans is the Director of Teaching and Learning Science at Frankston High School. She has nine years of experience teaching VCE Chemistry, Science and Environmental Science, and five years of experience as a VCAA assessor for Chemistry. Kate was awarded the CEA Travel Scholarship in 2019 and was an international winemaker before pursuing a career in teaching. She authored Oxford University Press's *Chemistry for Queensland Units 1 & 2 Student Workbook*.

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Carolyn Drenen has been teaching VCE Chemistry and Science in secondary schools for the past nine years, and is currently at Lalor North Secondary College. As an ECCN committee member, she has presented workshops at previous VCE Chemistry Conferences since 2015. She also connects with pre-service teachers in her role of university liaison. She authored Oxford University Press's *Chemistry for Queensland Units 3 & 4 Student Workbook*.

Reviewer: Matthew Di Petta

Matthew Di Petta is a VCE Chemistry and Psychology teacher, who is currently the Head of Chemistry at Melbourne Grammar School. Matthew completed a Science degree in both Chemistry and Psychology, and was fortunate enough to spend time teaching in Italy as part of his training. He has held roles with the VCAA, including as an exam assessor, and has presented exam revision lectures for both VCE subjects across Victoria.

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James Kennedy has been a VCE Chemistry teacher for nine years at several schools, including Haileybury, Loreto Mandeville Hall Toorak, Monash College and Wesley College. He has a wealth of experience in science communication and speaks at corporate events about how to tackle an irrational fear of 'chemicals'. His latest book, *Everything is Natural*, was published in 2021 by the Royal Society of Chemistry.

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Carrie Bloomfield has been teaching VCE Chemistry and Science in secondary schools for the past ten years and is currently a leading teacher at Mount Eliza Secondary College. She is a VCAA assessor for Chemistry and has presented on science inquiry skills at previous workshops for the *Oxford Science* series. Carrie co-authored Oxford University Press's *Chemistry for Queensland Units 1 & 2* and *Chemistry for Queensland Units 3 & 4*.

Reviewer: Carolyn Vaughan

Carolyn Vaughan has been teaching VCE Chemistry and Science at Frankston High School since 1990, where she has been the Assistant Director of Science for 11 years. She has a wealth of experience in developing Chemistry and Science curricula and really enjoys mentoring pre-service and graduate teachers.

CHAPTER 1

Chemistry toolkit

KEY SCIENCE SKILLS

→ Develop aims and questions, formulate hypotheses and make predictions
→ Plan and conduct investigations
→ Comply with safety and ethical guidelines
→ Generate, collate and record data
→ Analyse and evaluate data and investigation methods
→ Construct evidence-based arguments and draw conclusions
→ Analyse, evaluate and communicate scientific ideas

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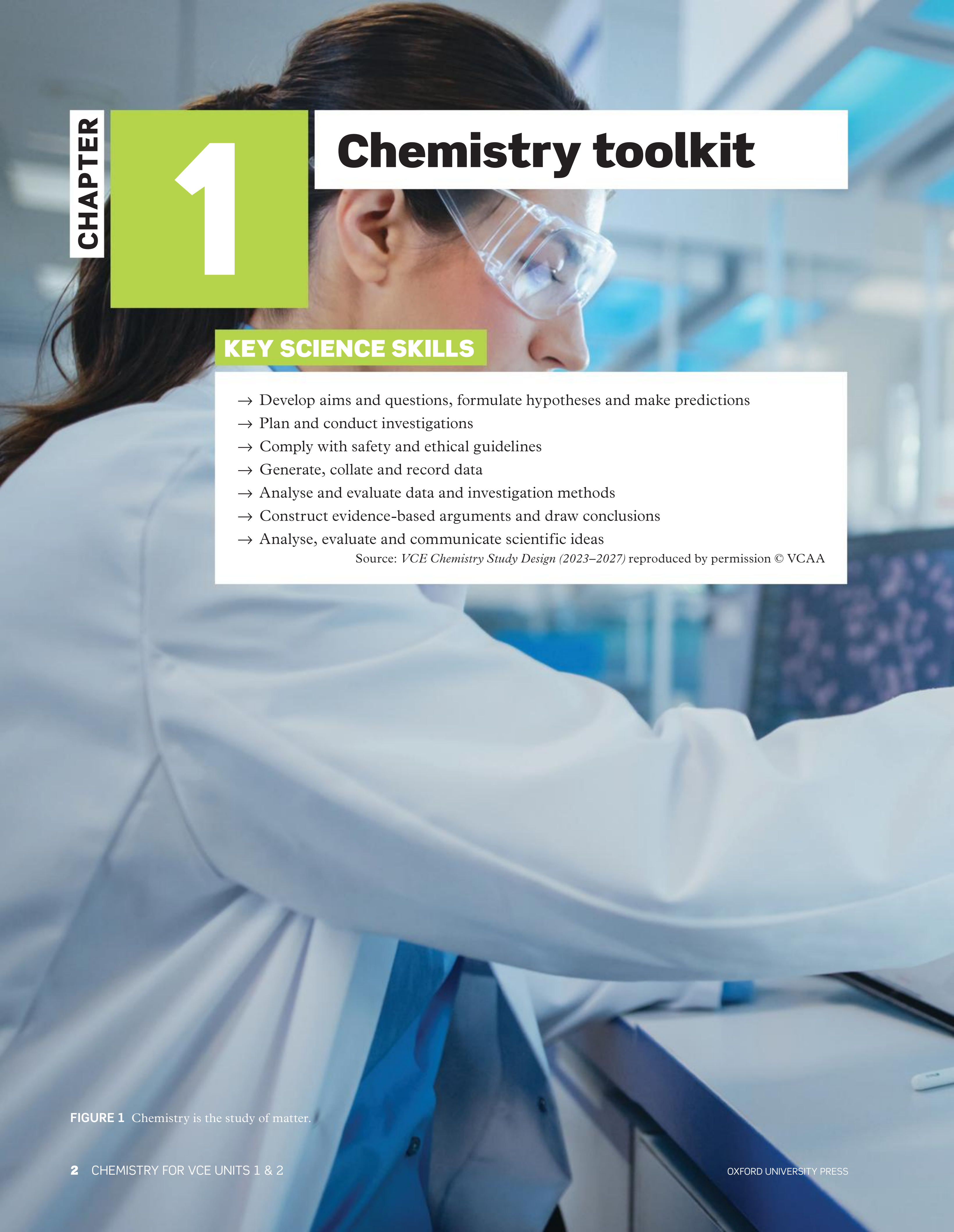

FIGURE 1 Chemistry is the study of matter.

GROUNDWORK

In Chapter 1, you will gain an overview of the VCE Chemistry course and build on your key science skills. You will also learn about the sustainability perspectives that are important for VCE Chemistry and gain skills to succeed in your assessments.

This chapter will build on skills you have already started developing in Years 7–10 Science. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

1A What is a hypothesis?

1A Groundwork resource
Hypotheses

1B Define qualitative data and quantitative data and give an example of each.

1B Groundwork resource
Qualitative and quantitative data

1C Identify at least three features you should include when presenting a graph or a table.

1C Groundwork resource
Presenting data

1D Contrast a controlled variable, an independent variable and a dependent variable in an experiment or investigation.

1D Groundwork resource
Variables

1.1 Overview of VCE Chemistry

KEY IDEAS

In this topic, you will learn that:

+ studying Chemistry can lead to a diverse range of career pathways
+ VCE Chemistry is divided into units and areas of study
+ the key science skills and their application are important for success in VCE Chemistry.

chemistry
the study of matter, including its structure, properties and behaviour

Chemistry is the study of the properties and behaviour of matter. It is often called the 'central science' because key concepts from Chemistry can be linked to other subjects. If you study VCE Biology or VCE Physics, you will see that some key ideas can be linked to what you are learning in VCE Chemistry.

FIGURE 1 Pharmacists use chemistry in their occupation.

Careers in chemistry

Studying Chemistry in VCE can lead to many opportunities when you finish school. Industries that chemists work in include (but are not limited to):

- medicine and health care (e.g. pharmacist, doctor, toxicologist, drug development)
- environment (e.g. environmental chemist, geochemist, water quality scientist)
- manufacturing (e.g. food and drink manufacturing, chemical engineering, materials development)
- science communication (e.g. science writer, teacher, science lecturer, science presenter).

Studying Chemistry in VCE also gives you a wide variety of transferrable skills, which are skills that you can use across many different disciplines. These skills include:

- problem-solving
- teamwork
- critical thinking
- communication
- research.

Structure of the VCE Chemistry course

Chemistry is one of the five science courses offered in VCE. When you study VCE Chemistry, you have an opportunity to engage in a range of inquiry tasks and develop key science skills. You will develop an understanding of the elements, the composition and behaviour of matter and the processes involved when matter reacts to produce useful materials in sustainable ways.

VCE Chemistry consists of four units (see Figure 2). Each unit is divided into Areas of Study. The Areas of Study that you will explore in Units 1 and 2 are outlined in Table 1.

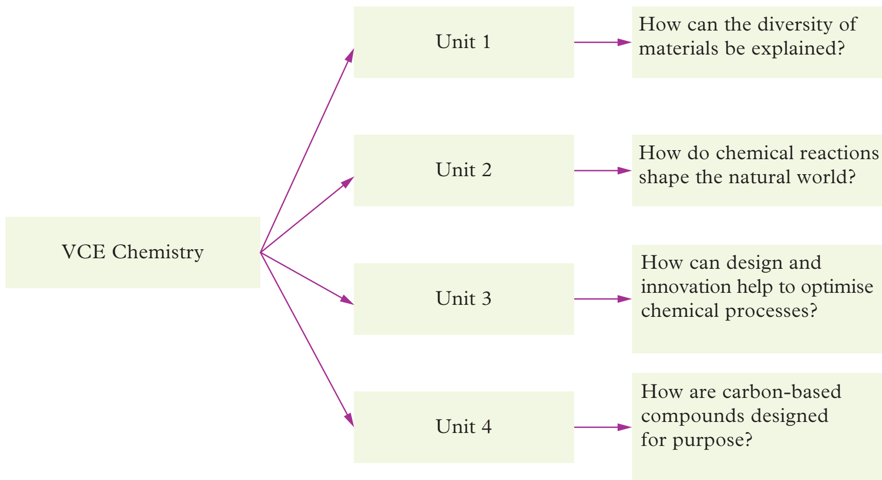

FIGURE 2 The structure of the VCE Chemistry course

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TABLE 1 Areas of Study for Units 1 and 2 Chemistry

Unit 1 How can the diversity of materials be explained?	
Area of Study	**Description**
1. How do the chemical structures of materials explain their properties and reactions?	In this Area of Study, you will learn about: • elements as the building blocks of useful materials • structures, properties and reactions of carbon compounds, metals and ionic compounds • separating the components of mixtures using chromatography.
2. How are materials quantified and classified?	In this Area of Study, you will learn about: • the measurement of quantities in chemistry • properties of organic compounds, including polymers.
3. How can chemical principles be applied to create a more sustainable future?	In this Area of Study, you will select and investigate a recent discovery, innovation, case study or issue related to the concepts you have covered in Areas of Study 1 and 2.
Unit 2 How do chemical reactions shape the natural world?	
Area of Study	**Description**
1. How do chemicals interact with water?	In this Area of Study, you will learn about: • the properties of water, including density, specific heat capacity and latent heat of vaporisation • acid–base and redox reactions, including their chemical equations and how they are used in the real world.
2. How are chemicals measured and analysed?	In this Area of Study, you will learn about: • the analysis of chemical reactions involving acids, bases, salts and gases • solubility of substances in water, the relationship between solubility and temperature, using solubility curves, and how to predict when a solute will dissolve or crystallise out of a solution • quantifying amounts in chemistry using volumetric analysis, the ideal gas equation, stoichiometry and calibration curves
3. How do quantitative scientific investigations develop our understanding of chemical reactions?	In this Area of Study, you will conduct an investigation that you have designed or adapted yourself. The investigation will relate to chemical equations and/or analysis that you have learnt about in Unit 2. To do this, you will write your own research question, and then conduct an experiment that generates primary data for you to answer your question. You will maintain a logbook for record, assessment and authentication purposes.

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How you will be assessed

In Units 1 and 2, you can be assessed in several different ways for the Areas of Study. These are outlined in Table 2. The assessment in VCE Chemistry is designed to test you against an Outcome for each Area of Study. The Outcome outlines what you should be able to do by the end of the Area of Study if you have completed it successfully.

TABLE 2 Areas of Study 1 and 2 Outcomes and Assessment tasks

<table>
<tr><th>Area of Study</th><th>Outcome</th><th>Assessment</th></tr>
<tr><td>Unit 1 Area of Study 1:
How do the chemical structures of materials explain their properties and reactions?</td><td>the student should be able to explain how elements form carbon compounds, metallic lattices and ionic compounds, experimentally investigate and model the properties of different materials, and use chromatography to separate the components of mixtures</td><td rowspan="4">For each Outcome in Areas of Study 1 and 2, you could be asked to complete one or more of the following tasks:
• a report of a laboratory or fieldwork activity, including the generation of primary data
• a comparison and evaluation of chemical concepts, methodologies and methods, and findings from at least two student practical activities
• reflective annotations of one or more practical activities from a logbook
• a summary report of selected practical investigations
• a critique of an experimental design, chemical process or apparatus
• an analysis and evaluation of generated primary and/or collated secondary data
• a modelling or simulation activity
• a media analysis/response
• problem-solving involving chemical concepts, skills and/or issues
• a report of an application of chemical concepts to a real-life context
• an analysis and evaluation of a chemical innovation, research study, case study, socio-scientific issue, secondary data or a media communication, with reference to sustainability (green chemistry principles sustainable development and/or the transition to a circular economy)
• an infographic
• a scientific poster.</td></tr>
<tr><td>Unit 1 Area of Study 2:
How are materials quantified and classified?</td><td>the student should be able to calculate mole quantities, use systematic nomenclature to name organic compounds, explain how polymers can be designed for a purpose, and evaluate the consequences for human health and the environment of the production of organic materials and polymers.</td></tr>
<tr><td>Unit 2 Area of Study 1:
How do chemicals interact with water?</td><td>the student should be able to explain the properties of water in terms of structure and bonding, and experimentally investigate and analyse applications of acid–base and redox reactions in society.</td></tr>
<tr><td>Unit 2 Area of Study 2:
How are chemicals measured and analysed?</td><td>the student should be able to calculate solution concentrations and predict solubilities, use volumetric analysis and instrumental techniques to analyse for acids, bases and salts, and apply stoichiometry to calculate chemical quantities.</td></tr>
</table>

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The Area of Study 3 Outcomes and Assessment tasks are slightly different.

TABLE 3 Area of Study 3 Outcomes and Assessment tasks

Area of Study	Outcome	Assessment
Unit 1 Area of Study 3: How can chemical principles be applied to create a more sustainable future?	the student should be able to investigate and explain how chemical knowledge is used to create a more sustainable future in relation to the production or use of a selected material.	a response to a question involving the production or use of a selected material, including reference to sustainability
Unit 2 Area of Study 3: How do quantitative scientific investigations develop our understanding of chemical reactions?	the student should be able to draw an evidence-based conclusion from primary data generated from a student-adapted or student-designed scientific investigation related to the production of gases, acid–base or redox reactions or the analysis of substances in water.	a report of a student-adapted or student-designed scientific investigation using a selected format, such as a scientific poster, an article for a scientific publication, a practical report, an oral presentation, a multimedia presentation or a visual representation

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You can use Chapter 10 Research investigation to guide you through Unit 1 Area of Study 3, and Chapter 18 Student-designed investigation to guide you through Unit 2 Area of Study 3.

Key science skills

The key science skills are applicable to all Areas of Study in Units 1–4 of the VCE Chemistry course. They are especially important for planning and conducting investigations for your assessment tasks.

In addition to key knowledge (which we will cover in Chapters 2–18), VCE Chemistry requires you to develop and apply a range of key science skills. These skills are specified in the VCAA VCE Chemistry Study Design and are listed in Table 4.

TABLE 4 Key science skills

Key science skill	VCE Chemistry Units 1–4
Develop aims and questions, formulate hypotheses and make predictions (see Topic 1.3)	• identify, research and construct aims and questions for investigation • identify independent, dependent and controlled variables in controlled experiments • formulate hypotheses to focus investigations • predict possible outcomes of investigations
Plan and conduct investigations (see Topic 1.4)	• determine appropriate investigation methodology: case study; classification and identification; controlled experiment; correlational study; fieldwork; literature review; modelling; product, process or system development; simulation • design and conduct investigations; select and use methods appropriate to the selected investigation methodology, including consideration of sampling technique and size, equipment and procedures, taking into account potential sources of error and causes of uncertainty; determine the type and amount of qualitative and/or quantitative data to be generated or collated • work independently and collaboratively as appropriate and within identified research constraints, adapting or extending processes as required and recording such modifications

(*continued*)

TABLE 4 continued

Key science skill	VCE Chemistry Units 1–4
Comply with safety and ethical guidelines (see Topics 1.5 and 1.6)	• demonstrate safe laboratory practices when planning and conducting investigations by using risk assessments that are informed by safety data sheets (SDS), and accounting for risks • apply relevant occupational health and safety guidelines while undertaking practical investigations • demonstrate ethical conduct when undertaking and reporting investigations
Generate, collate and record data (see Topic 1.7)	• systematically generate and record primary data, and collate secondary data, appropriate to the investigation, including use of databases and reputable online data sources • record and summarise both qualitative and quantitative data, including use of a logbook as an authentication of generated or collated data • organise and present data in useful and meaningful ways, including schematic diagrams, flow charts, tables, bar charts and line graphs
Analyse and evaluate data and investigation methods (see Topic 1.8)	• process quantitative data using appropriate mathematical relationships and units, including calculations of ratios, percentages, percentage change and mean • use appropriate numbers of significant figures in calculations • plot graphs involving two variables that show linear and non-linear relationships • identify and analyse experimental data qualitatively, handling where appropriate concepts of accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements; and errors (random and systematic) • identify outliers, and contradictory, provisional or incomplete data • repeat experiments to evaluate the precision of data • evaluate investigation methods and suggest ways to improve accuracy and precision, and to reduce the likelihood of errors
Construct evidence-based arguments and draw conclusions (see Topic 1.9)	• distinguish between opinion, anecdote and evidence, and scientific and non-scientific ideas • evaluate data to determine the degree to which the evidence supports the aim of the investigation, and make recommendations, as appropriate, for modifying or extending the investigation • evaluate data to determine the degree to which the evidence supports or refutes the initial prediction or hypothesis • use reasoning to construct scientific arguments, and to draw and justify conclusions consistent with the evidence and relevant to the question under investigation • identify, describe and explain the limitations of conclusions, including identification of further evidence required • discuss the implications of research findings and proposals

(continued)

TABLE 4 continued

Key science skill	VCE Chemistry Units 1–4
Analyse, evaluate and communicate scientific ideas (see Topic 1.10)	• use appropriate chemical terminology, representations and conventions, including standard abbreviations, graphing conventions, algebraic equations, units of measurement and significant figures • discuss relevant chemical information, ideas, concepts, theories and models and the connections between them • analyse and explain how models and theories are used to organise and understand observed phenomena and concepts related to chemistry, identifying limitations of selected models/theories • critically evaluate and interpret a range of scientific and media texts (including journal articles, mass media communications and opinions in the public domain), processes, claims and conclusions related to chemistry by considering the quality of available evidence • apply sustainability concepts (green chemistry principles, development goals and the transition from a linear towards a circular economy) to analyse and evaluate responses to chemistry-based scenarios, case studies, issues and challenges • identify and explain when judgements or decisions associated with chemistry-related issues may be based on sociocultural, economic, political, legal and/or ethical factors and not solely on scientific evidence • use clear, coherent and concise expression to communicate to specific audiences and for specific purposes in appropriate scientific genres, including scientific reports and posters • acknowledge sources of information and assistance, and use standard scientific referencing conventions

1.1 CHECK YOUR LEARNING

Describe and explain

1 Summarise how your knowledge and skills will be assessed in VCE Chemistry Units 1 & 2.

2 Reflect on the key science skills listed in Table 4 and identify 3–5 skills you need to practise most.

Apply, analyse and compare

3 Follow the link in your obook pro to access the 'Chemistry bullseye poster'. Research two different chemistry-related occupations that interest you and answer the following questions.

- **a** Write down some of the types of activities that a person does in these occupations.
- **b** What education or qualification was completed in order to work in these occupations?
- **c** What qualities or interests should a person who is interested in these occupations have?
- **d** What do you think is the most interesting part of these occupations? Explain your answer.
- **e** What do you think is the most difficult or repetitive part of these occupations? Explain your answer.

Design and discuss

4 Discuss the results of your research from Question 3 with another student.

1.2 Aboriginal and Torres Strait Islander knowledge, cultures and histories

KEY IDEAS

In this topic, you will learn that:

+ Aboriginal and Torres Strait Islander Peoples are Australia's first chemists and apply their unique scientific knowledge about the natural world to find (or refine) suitable food sources, and develop traditional medicines and other useful materials in sustainable ways
+ Aboriginal and Torres Strait Islander Peoples come from many diverse nations, each of which have their own unique cultures, histories and knowledges.

Aboriginal and Torres Strait Islander cultures are some of the oldest continuous cultures in the world and have existed on this land for at least 60 000 years. The Traditional Custodians of Australia are people who are of Aboriginal and/or Torres Strait Islander descent, and who identify and are accepted as such by their community.

Aboriginal and/or Torres Strait Islander Peoples people with family heritage from, and membership in, the ethnic groups that lived in Australia before British colonisation

As a collective group, **Aboriginal and/or Torres Strait Islander Peoples** may also be referred to as Indigenous Peoples, First Peoples and/or Traditional Custodians/Owners. Indigenous Peoples may also identify with their particular Community, Nation, Tribe or Clan that they, their family and/or their community recognise.

We say Aboriginal and Torres Strait Islander Peoples (not 'people') because there are many diverse language groups in Australia (representing more than 250 languages), each of whom have their own unique cultures, laws, practices and languages. So, when learning or writing about the cultures and practices of Aboriginal and/or Torres Strait Islander Communities, Nations, Tribes or Groups in VCE Chemistry, it is important to identify and acknowledge the right owners of that knowledge.

FIGURE 1 Aboriginal Language groups and Nations within the borders of Victoria

Aboriginal and Torres Strait Islander knowledges and cultures in VCE Chemistry

The diverse cultures and knowledge systems of Aboriginal and Torres Strait Islander Peoples have been acknowledged in the VCE Chemistry course.

In VCE Chemistry, you will learn how scientific thinking can be enhanced by considering how Aboriginal and Torres Strait Islander Peoples have developed and refined their own knowledge, including the use of plants as medicine and the use and modification of natural materials for useful purposes.

For example, people of the Kulin Nation in Victoria can use the sap of the river red gum (*Eucalyptus camaldulensis*) to seal burns, or the smoke of older manna gum (*Eucalyptus viminalis*) leaves to reduce fever. In Western Australia, the Noongar Peoples use infusions of crushed bracken fern (*Pteridium esculentum*) to relieve sores, or the juice from the same plant to relieve insect bites. In New South Wales, the Bidjigal and Gadigal Peoples use the long leaves of the gymea lily (*Doryanthes excelsa*) to produce string, and its long stem to craft a fishing spear.

a **b** **c**

FIGURE 2 Aboriginal and Torres Strait Islander Peoples hold specialised knowledge about the plants and animals on Country, such as how to use **a** the leaves of the manna gum (*Eucalyptus viminalis*) in Victoria, **b** the bracken fern in Western Australia or **c** the gymea lily in eastern New South Wales.

Study tip

In Unit 1 Area of Study 3, you might research how Aboriginal and Torres Strait Islander Peoples sustainably modify and process raw materials. It is important to be able to carefully evaluate your resources to make sure you're using appropriate ones. See your teacher for more information on how to do this.

This knowledge has been accumulated and refined over thousands of years. It is highly sophisticated and specialised for the areas in which the people live.

1.2 CHECK YOUR LEARNING

Apply, analyse and compare

1 Identify the Traditional Owners of the land your school or home is built on.

Design and discuss

2 Explain why modern medicine would reference the chemical knowledge of Aboriginal and Torres Strait Islander Peoples to treat ailments and infections.

3 Research one of the following topics.
- The use of plants for traditional medicines in Victoria
- The chemical processes that occur when Aboriginal and Torres Strait Islander Peoples detoxify poisonous items (such as cycad seeds) before eating them
- The traditional use of animal fat and/or plant pigments to make paint

4 Evaluate the resources you used in Question 3. Justify your response.

1.3 Developing aims, questions and hypotheses

KEY IDEAS

In this topic, you will learn that:

+ a research question states the specific problem or issue on which your investigation will be based
+ an aim is a statement of what is to be investigated
+ a hypothesis is a testable statement that should include a prediction about the outcome of an investigation, based on scientific reasoning.

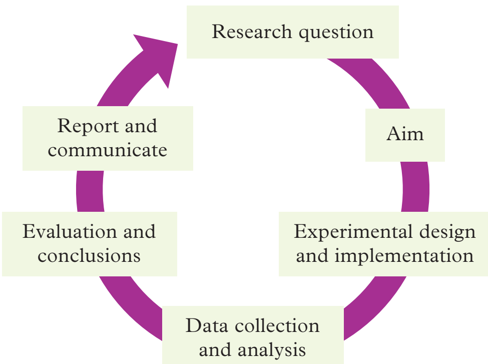

FIGURE 1 The general process of the scientific method. Adapted from: VCAA (2016) *Victorian Certificate of Education Chemistry Advice for Teachers 2016–2020*, page 6, Melbourne, Victorian Curriculum and Assessment Authority.

The **scientific method** is a cyclical process involving a number of steps (see Figure 1):

- developing a research question or aim
- formulating a hypothesis
- conducting experiments to collect, analyse and communicate data (including potential errors and uncertainty)
- developing conclusions
- communicating scientific ideas and understandings.

Developing a research question

In VCE Sciences, you will need to develop a research question for your research investigation (Unit 1) and student-designed investigation (Unit 2). When writing a research question, you should make sure the question is:

- **clear and specific**: your research question must contain enough detail for your audience to understand what you are investigating
- **concise**: your research question must be expressed in as few words as possible
- **complex**: your research question cannot be answered with a simple 'yes' or 'no', or with facts that are easily found; it must require enough scope for you to respond with an appropriate experiment with collection and analysis of primary and/or secondary data.

scientific method
a series of steps used to acquire knowledge in science, which involves observation, developing and testing hypotheses; collecting, analysing and communicating results

An example of a suitable research question for a scientific inquiry in VCE Chemistry is: Which plant biomass produces an efficient mass per unit volume of essential oil when extracted by steam distillation?

Writing an aim

aim
a statement that describes the purpose of an investigation

An **aim** is a short statement outlining what is to be investigated. The goal of your experiment or inquiry must be precisely stated. An example of an aim for a scientific inquiry in VCE Chemistry is:

To investigate the relationship between plant biomass and the volume of essential oil produced during extraction by steam distillation.

Independent and dependent variables should also be included where possible, without detail on the method or procedure used. Remember, an **independent variable** is what you change during a controlled experiment. The **dependent variable** is what you measure (it may or may not be affected by the independent variable). A **controlled variable** is one that is not changed or remains constant during an experiment.

independent variable
the variable that is changed or manipulated during an investigation

dependent variable
the variable that is observed or measured when the independent variable is changed during an investigation

controlled variable
a variable that is not changed or is constant during an investigation

Formulating a hypothesis and making a prediction

A **hypothesis** is a testable statement that predicts how the independent variable will affect the dependent variable in an investigation. It should include an explanation for the prediction based on scientific knowledge or reasoning.

A hypothesis cannot be wrong, only supported or negated by the investigation results.

hypothesis
a testable statement that includes a prediction about the outcome of an investigation based on scientific reasoning

If … then … because …

A good way to make sure you include everything in your hypothesis is to use a structure such as '*If … then … because*' or '*When … the … because*' statement.

Suggested phrases for formulating a hypothesis are listed in Figure 2.

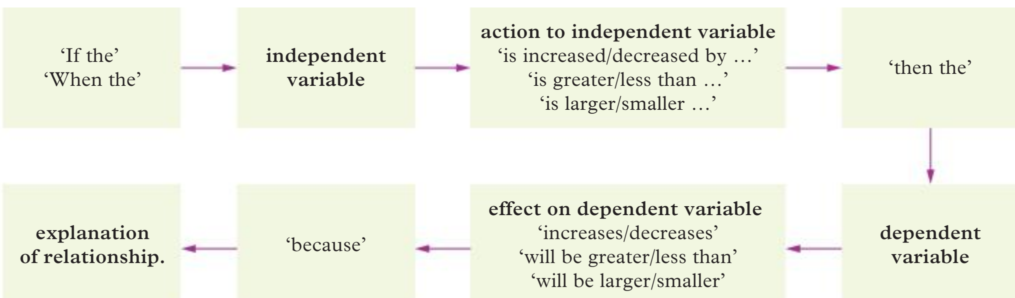

FIGURE 2 Constructing a hypothesis from a research question

An example of a hypothesis for a scientific inquiry in VCE Chemistry is:

When the hydrocarbon chain length of an alcohol is increased, then its solubility in water will decrease, because increasing the non-polar section of the molecule increases hydrophobicity and interferes with the solubility of the polar hydroxyl (–OH) functional group on the molecule.

See how to write an aim, research question and hypothesis in Worked example 1.3.

Study tip

Make sure you know how to write a hypothesis. This is the kind of skill that is often tested in assessment tasks or exams.

1.3 Worked example
Find me in your obook pro

1.3 Worked example
Video demonstration

1.3 CHECK YOUR LEARNING

Describe and explain

1 Suggest why the steps of the scientific method, as illustrated in Figure 1, are often shown as a circular process.

Apply, analyse and compare

2 Contrast an aim and a hypothesis.

Design and discuss

3 During studies of acids and bases, a student observed that when she poured lemon juice onto a plant, it shrivelled and died. The student wanted to use this observation as the basis for a practical investigation.

a Design a research question for this practical investigation.

b Write an aim for this practical investigation.

c Formulate a hypothesis for this investigation.

1.4 Planning and conducting investigations

KEY IDEAS

In this topic, you will learn that:

+ a variety of methodologies can be used for a scientific investigation
+ a logbook is required for recording raw data and must be submitted intermittently for assessment purposes.

In VCE Chemistry, you will complete at least 10 hours of practical work per unit.

Scientific investigations are important to the VCE Chemistry course and can be used to collect primary data or data from secondary sources. You can complete an investigation individually, in a small group or with the whole class; however, all the work required for assessment purposes (e.g. the logbook and the poster) must be your own work.

Scientific investigation methodologies

methodology
a system of methods used in a particular area of study or activity

A variety of **methodologies** can be used when planning and conducting scientific investigations. A methodology refers to the approach you take to answer your scientific question. Not all scientific questions should or need to be answered with a controlled experiment. There are other ways to find the answers to your scientific questions or collect data. Some of the methodologies that you will encounter in VCE Chemistry are outlined in Table 1.

TABLE 1 Key scientific investigation methodologies in VCE Sciences

Investigation methodology	Investigation outline
Case study	A case study involves the investigation of a specific event, an activity or a problem that contains a real or hypothetical situation. Case studies can be historical, involving the analysis of causes and consequences, and/or discussion of knowledge learnt from the situation; a real situation or a role play of an imagined situation, where plausible recommendations are to be made; or problem solving, where developing a new design, methodology or method is required.
Classification and identification	Classification refers to the arrangement of phenomena, objects or events into smaller, more manageable groups (e.g. classifying organic compounds into functional groups). Identification is a process of recognising things as belonging to particular sets or possibly being part of a new set.
Controlled experiment	A controlled experiment investigates the relationship between an independent variable and a dependent variable. All other variables are controlled so they don't affect the outcome of the investigation.
Fieldwork	Fieldwork involves going to a specific location to investigate a phenomenon or problem that is unique to that site. When completing fieldwork, you should record site-specific data in your logbook, such as observations and data collected about the conditions of the environment (e.g. air temperature, pH of the water).
Literature review	A literature review involves researching, gathering and interpreting secondary sources (also called 'literature' by academics) to answer a research question. This may be used to answer questions that have already been asked by other scientists, or to develop a background understanding before you start your own investigation.
Modelling	A model physically, conceptually or mathematically simulates a concept to assist understanding (e.g. modelling molecule structures, which are too small for us to observe in the classroom).

(*continued*)

TABLE 1 continued

Investigation methodology	Investigation outline
Product, process or system development	Product, process or system development involves designing an object, process or system to meet a need. This should link scientific knowledge to technological developments.
Simulation	Simulations involve using an existing model to investigate a scientific phenomenon by manipulating variables in the simulation. Simulations are useful when variables cannot be manipulated in real life.

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Logbook

You will need a logbook in hard copy or a digital form to keep primary and secondary data records from all types of scientific investigations undertaken in Units 1–4. You will need to submit your logbook to your teacher for some assessment tasks so they can check that the work you have submitted is your own.

Each entry must include the date and information you add in chronological order, and an acknowledgement of all secondary resources, as well as all expert advice and teacher assistance received.

Your logbook could also include the following information as you complete your investigation:

- planning notes for experiments
- a description of the activities you have carried out
- the results/data from guided activities or investigations (including outliers and/or risk identification and management)
- personal reflections made during or at the conclusion of demonstrations, activities or investigations
- any links to spreadsheet calculations or other digital records and presentation you might use
- any notes and electronic (or other images) taken on excursions, or any database extracts
- notes of any additional work completed outside of class time.

A well-organised logbook that contains all of this information will make it a lot easier to complete your scientific poster or report in Unit 2 Area of Study 3.

FIGURE 1 Modelling can help us to investigate scientific phenomenon if we are unable to observe the real thing in the lab.

FIGURE 2 Fieldwork allows us to investigate phenomena at a specific location. For example, this scientist is testing the pH of the water at this waterway.

1.4 CHECK YOUR LEARNING

Describe and explain

1 What are the advantages of keeping a logbook for investigations in VCE Chemistry?

2 What are the differences between primary and secondary data?

Apply, analyse and compare

3 Compare:

a a controlled experiment and fieldwork

b a case study and a literature review

c modelling and a simulation.

Design and discuss

4 Design a potential research question for one of the scientific investigation methodologies.

1.5 Safety in chemistry

KEY IDEAS

In this topic, you will learn that:

- safety must be your priority when performing scientific investigations
- a risk assessment is an organised way of identifying hazards and/or risk factors, and to implement controls for prevention when performing scientific investigations.

Chemists often work with dangerous goods and hazardous substances. You might encounter some of these dangerous substances when conducting practical work in VCE Chemistry.

Some hazards are clearly labelled for us. (See Figure 1 for examples of chemical hazard warning symbols.) But other hazards still need to be avoided to keep everyone safe. We can do this by following basic laboratory safety rules and conducting risk assessments.

Rules for laboratory safety

- **Always wait for teacher permission** before entering the laboratory, collecting materials and/or equipment and beginning your activities. Teacher and laboratory technician instructions must also be obeyed for OHS (occupational health and safety) reasons.
- **Always wear the correct personal protective equipment appropriately**. Laboratory coats should be buttoned up to protect your clothes. Laboratory glasses or goggles should be worn on the face for eye protection. Laboratory gloves should be worn whenever possible to protect skin from heat, cold and chemicals. Closed-toe shoes should be worn for foot protection and long hair should be tied back behind the neckline.
- **Always behave sensibly**. Treat the laboratory as a proper workplace environment. Be courteous and respectful to your peers, teachers and other staff who may enter.
- **Never eat or drink in the laboratory**. You don't know what substances have been in the laboratory previously, and ingestion of hazardous materials could be harmful.
- **Report all accidents immediately to the teacher in charge**, including spills, leaks, breakages, and faulty and/or damaged equipment, to minimise escalation of potential hazards.
- **Turn off all equipment and/or pack it away in a careful and safe manner** (for future use by others). Place waste materials and/or chemicals in appropriate containers for disposal and leave your workplace area clean, dry and tidy.

FIGURE 1 The Globally Harmonised System of Classification and Labelling of Chemicals (GHS) is used in Australia.

Risk assessments

A **risk assessment** is an organised way to identify:

- any risk factors and/or hazards that might occur during an investigation
- controls or actions to prevent accidents, injuries or harm to you (or anyone else who might be in the laboratory).

As part of VCE Chemistry, you may need to write and submit a risk assessment to accompany investigations that you have designed yourself.

FIGURE 2 When conducting experiments, you should wear a lab coat, gloves, safety glasses and closed-toe shoes, and tie back long hair.

Risk assessments are written for all workplace environments, including the laboratory, to meet model work health and safety acts, legislation and codes of practice.

risk assessment
a process of evaluating the potential risks that may be involved in an activity, e.g. performing an experiment

Risk assessments can be written in a variety of formats and some schools may use external programs to generate them. An example of a risk assessment for a laboratory-based activity (with annotations) is provided in your obook pro. A blank risk assessment template is also provided.

It is important to consider the information in each section of the risk assessment before assigning a final risk judgment. For example, to reduce the risk of using a hazardous substance, it could be:

- diluted
- dispensed from a dropper bottle
- used by a person who is wearing their personal protective equipment appropriately
- used when following standard laboratory rules.

1.5 Risk assessment
Find me in your obook pro

It is also important to implement and follow the control measures outlined for all potential hazards, physical and/or chemical, in the risk assessment, to ensure the safety of everyone involved in the activity.

Access an annotated risk assessment using the digital icon.

1.5 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is important to comply with safe work practices when planning and conducting laboratory-based activities.

2 Describe the process of how to make a final risk judgment of a laboratory-based activity process or procedure, using an example of a practical activity or teacher demonstration from Units 1 or 2.

Apply, analyse and compare

3 Compare the similarities and differences between the safety data sheets of one GHS substance classified as hazardous and another one classified as non-hazardous.

Design and discuss

4 An experiment is to be conducted to determine the solubility of potassium chloride in water at various temperatures, using a burette, small funnel, large test tube, retort stand with clamp, kettle, thermometer and large beaker. Design a risk assessment for this experiment.

1.6 Ethical understanding

KEY IDEAS

In this topic, you will learn that:

+ ethics are applied when performing all types of scientific investigations in VCE Chemistry.

ethics
moral principles that govern a person's behaviour or how an activity is conducted

Study tip

When collecting or reporting the results of scientific investigations, be careful to minimise data errors, be honest (do not mislead, or change or make up data) and avoid bias from unexpected data or results.

Ethical understanding is applied across Units 1–4 of VCE Chemistry, especially when performing scientific investigations, analysing data collected from primary and secondary sources, and identifying and investigating any issues relating to the application of scientific knowledge in society.

In VCE Chemistry, **ethics** can include:

- thinking about the impact of scientific investigations (including your own) on living and non-living things and the environment
- being honest when recording data and presenting information from your own investigations, and acknowledging when you are using someone else's ideas or data
- developing an opinion or argument for or against science-related ethical issues based on your understanding of ethical concepts (considering current and future needs) and scientific knowledge
- identifying factors that might influence decision making in science (such as personal values or economic, political or legal factors).

1.6 CHALLENGE

Ethical issues in chemistry

Modern ethical issues in chemistry include: manufacture and use of 'designer' polymers, nanoparticles and pharmaceuticals; sustainable consumption of materials and goods; price and/or safety of drugs made by pharmaceutical companies.

Research one of these issues and answer the following questions.

1 What known facts are presented?
2 What ethical concerns do they raise?
3 Which of the known facts are *relevant* to resolving the ethical concerns?
4 What additional facts might be relevant to the case?
5 How might the additional facts affect what is ethically at stake?
6 What options are available, and which seem the best from an ethical viewpoint?

FIGURE 1 Products developed by pharmaceutical companies can be expensive. Is this ethical?

1.6 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is important to understand and apply ethics in VCE Chemistry.

Apply, analyse and compare

2 Download a copy of the *Global Chemists' Code of Ethics*, from your obook pro. Compare this document with the key knowledge outcomes from the VCE Chemistry Study Design, giving reasons for how well they align.

Design and discuss

3 Design an A4-sized report or brochure of the arguments for or against (based on collected evidence) one of the ethical issues in Challenge 1.6. Identify and justify your own position on the issue.

1.7 Generating, collating and recording data

KEY IDEAS

In this topic, you will learn that:

+ data generated from scientific investigations is important evidence to support the trends, patterns and/or relationships between variables
+ raw data can be qualitative or quantitative, discrete or continuous
+ raw data from scientific investigations needs to be presented clearly so that it can be easily understood by your audience.

In VCE Chemistry and other sciences, investigations are important for developing explanations, supported by evidence (or data), to explain natural phenomena and events. Raw data generated from scientific investigations and other evidence recorded in your logbook must be presented in an appropriate format to illustrate any trends, patterns and/or relationships between the independent, dependent and controlled variables.

Types of data

From previous studies in statistics or mathematics, you know that there are many types of raw (or unprocessed) data that can be collected. Studies in VCE Chemistry involve the following raw data types.

- **Qualitative data** can be described in words, phrases or categories (e.g. classifying types of substances or describing investigation observations in dot points or sentences).
- **Quantitative data** can be described with numbers, quantities or other numerical values, often collected during scientific investigations. Quantitative data can be further classified as:
  - **discrete data** – set numerical values that are not related (e.g. comparing the energy content of different foods or fuels)
  - **continuous data** – numerical values within a specified range that are usually generated when measurement is involved (e.g. temperature, mass, length, voltage, volume, time, pH).

qualitative data
data that is not numerical, can be text, images or audio

quantitative data
data that can be counted or measured and is expressed as numbers

discrete data
data or information that can only be certain numerical values

continuous data
data or information that can be any numerical value

collate
collect and combine texts, information or data

Collating and recording data

All raw data from scientific investigations, either qualitative or quantitative, should be **collated** presented in a table, which has a title and labelled columns.

Organising qualitative data in tables

For qualitative data, the column labels should include a word or phrase description of the variables. Data presented in these tables often shows trends that may be compared and contrasted in the discussion section.

Organising quantitative data in tables

For quantitative data, the column labels should include a description of the independent and dependent variables and unit(s) if measurement is involved.

Data presented in these tables displays the numerical values of each variable without a clear indication of their relationship(s). So this means you need to use a graph (or other visual aid) to show any patterns or relationships.

TABLE 1 A template for organising quantitative data in a table

Independent variable (include units)	Dependent variable (include units)			
	Trial 1	Trial 2	Trial 3	Average

Graphing data

Scientific graphs have the following features in common:

- a title, which should be a descriptive statement that includes the convention ‘dependent variable versus independent variable’
- axes labels with units and scale numbers
- independent variable plotted on the horizontal (x) axis
- dependent variable is plotted on the vertical (y) axis
- properly scaled axes (of equal unit size) that fit the allocated space, and which ensure that the data is plotted on the graph (not beyond each axis). The graph should be as large as possible in the allocated space for clarity
- a colour coding system or different symbols for each different set of data points if more than one data set is plotted on the same pair of axes.

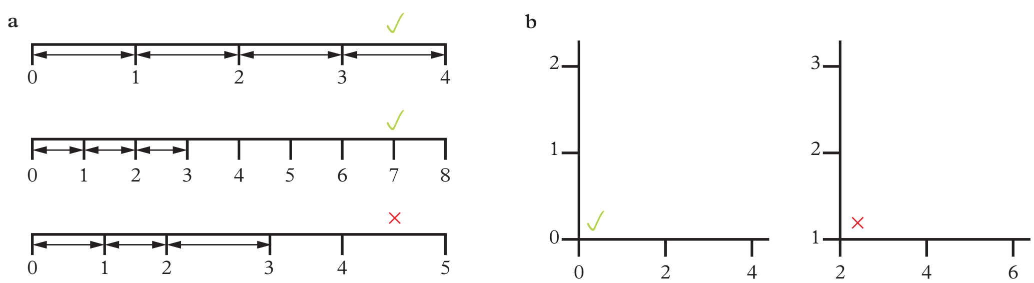


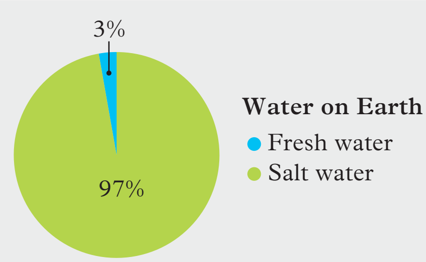
FIGURE 1 When setting up a graph, remember to **a** scale your axes evenly, and **b** begin your axes at 0 (or the origin).

Figure 1 shows some examples of what to do and what not to do when drawing graphs. The points on your axis scale should be the same distance apart. If you use numbers, they should also increase by the same amount each time.

Graphing discrete data

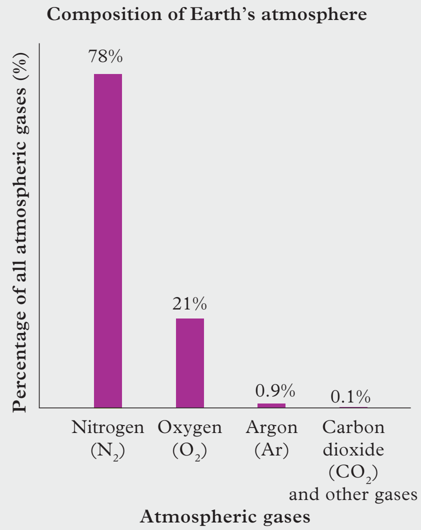
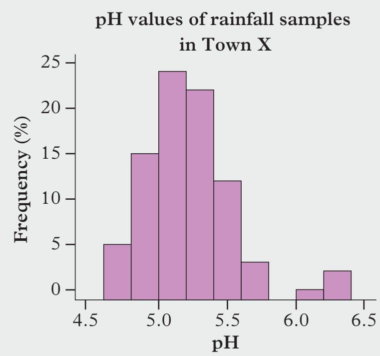
Qualitative or discrete data can be presented using different types of graphs (see Table 2).

TABLE 2 Graphing discrete data

Type of graph	Description	Example
Pie chart	Pie charts are drawn as circles and divided into segments that each represent a category. They show the proportion of each category as part of the whole data set.	

(continued)

TABLE 2 continued

Type of graph	Description	Example
Column graph	Column graphs are used to display categorical data. The bars can be plotted vertically or horizontally but should not be touching.	
Histogram	Histograms are used to show the **frequency** of numerical data groups. Each column width on a histogram can be a single data value or a data interval if the data is grouped (e.g. 0–5, 5–10, 10–15). The height of the column corresponds to the frequency, expressed as a number or percentage.	

frequency
the number of times an observation occurred in an experiment or a study

Graphing continuous data

Line graphs or scatterplots are useful ways of presenting continuous data sets where there are relationships between the dependent and independent variables.

A line graph is used to display one or more data sets that are in a defined relationship (Figure 2).

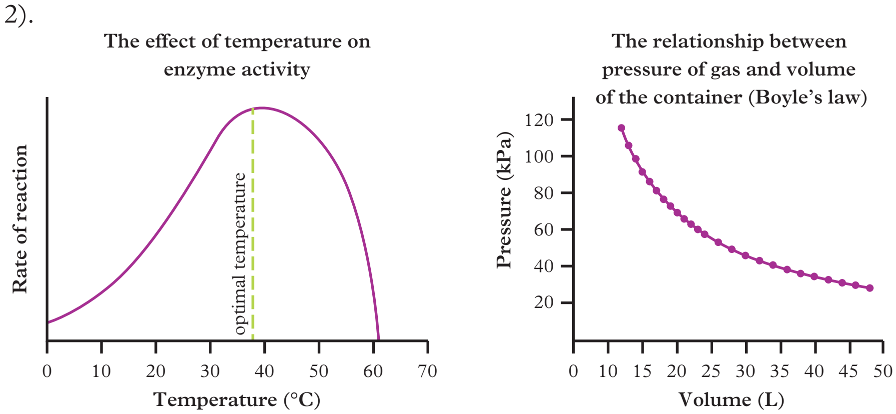


FIGURE 2 Line graphs are used to describe the relationship between two variables in a defined relationship.

A scatterplot is used to identify trends between two data sets that are not in a defined relationship.

Study tip

Do not confuse histograms with column graphs. Histograms are used for displaying numerical data, and their bars touch. Column graphs are used to show categories of data, and their bars do not touch.

line of best fit
a line drawn through a scatterplot of data points that expresses a relationship between those points

Line of best fit

A **line of best fit** can be drawn on a scatterplot to demonstrate any trends in the data (see Figure 3). Some points may be above, below or on the line of best fit. If the trend of the data is linear, the line of best fit is drawn with a straight edge. If the trend of the data is curved, a smooth curve should be drawn. If there is no visible trend between the data points (i.e. they are scattered all over the graph), then no line of best fit should be drawn.

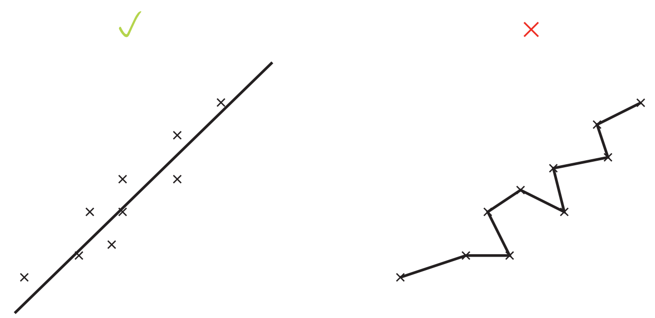


FIGURE 3 A line of best fit is drawn through the middle of your data points. You shouldn't connect the dots. A line of best fit can only be drawn if there is a strong positive or strong negative correlation.

Correlation

The relationship between two variables is called the **correlation.** If the data points on a scatterplot are close together, it could mean there is a stronger correlation between the variables.

- If one variable increases as the other variable increases, it is called a positive correlation. Weight and height are two variables that commonly have a strong positive correlation.
- If one variable increases as the other variable decreases, it is called a negative correlation.
- If there is no trend between the variables (data points are scattered all over the graph), the data is showing no correlation between the variables.

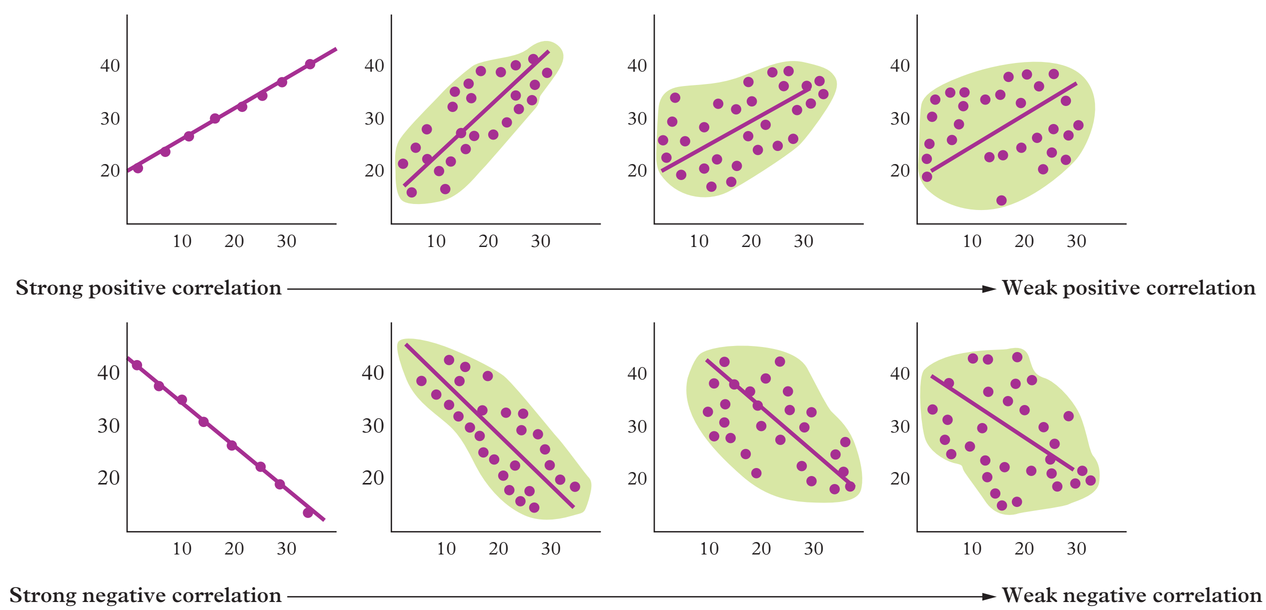


FIGURE 4 Scatterplots are used to describe the relationships between two variables that are not in a straight line.

1.7 CHECK YOUR LEARNING

Describe and explain

1 State the main features of scientific graphs.

Apply, analyse and compare

2 Contrast column graphs and histograms.

Design and discuss

3 The laboratory technician at your school has a waste solution of copper(II) sulfate, of unknown concentration, that needs to be contained and sent off-site for disposal. The local chemical waste processing plant has requested a statement of the exact concentration of this solution before they accept the container. Students have volunteered to help determine the concentration of this unknown solution by colorimetry.

The students prepare a series of copper(II) sulfate solutions of known concentration. They fill 3 mL cuvettes with the solutions. They also fill one cuvette with distilled water and another cuvette with the waste solution of unknown concentration. Each filled cuvette is placed in the colorimeter and the students measure the percentage absorbance of each coloured solution.

FIGURE 5 A colorimeter measures the amount of light absorbed by different solutions.

The raw data obtained from this practical investigation is given in Table 3.

TABLE 3 The students' raw data

Trial	$CuSO_4$ concentration ($mol\ L^{-1}$)	Measured absorbance (%)
1	0	0.00
2	0.2	0.06
3	0.5	0.16
4	0.7	0.22
5	0.9	0.28
6	Waste solution	0.20

a Describe the purpose of using distilled water in this practical investigation.

b Identify the dependent and the independent variables in this practical investigation.

c Design a visual representation of the data in Table 3 with appropriate labels. Use the evidence from your representation and Table 3 to identify the relationship between the dependent and independent variables in this practical investigation.

d Use the visual representation from part c to determine the $CuSO_4$ concentration in the waste solution from trial 6.

1.8 Evaluating data and investigations

KEY IDEAS

In this topic, you will learn that:

+ investigations that generate raw data must be valid, repeatable and reproducible
+ errors and outliers must be included and accounted for in data evaluations
+ significant figures must be considered in all calculations in VCE Chemistry.

The raw data generated from the design and implementation of scientific investigations must be analysed and discussed in terms of its quality and quantity; it must be **valid**, **repeatable** and **reproducible**. The VCE Chemistry Study Design provides specific guidance for terms and definitions for data, measurement, errors, significant figures and outliers.

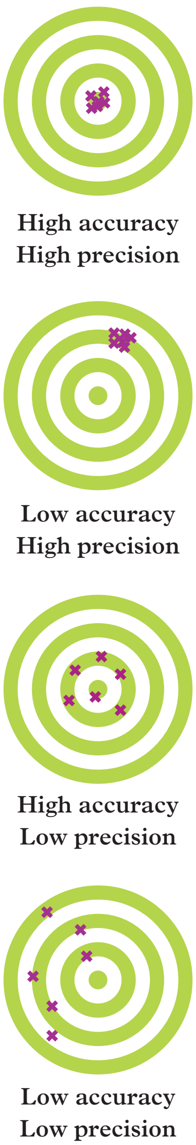


FIGURE 1 Accuracy and precision are different concepts.

Data and measurement

When analysing and evaluating the measurement of data, you should know a number of key terms and be able to use them. These terms are shown in bold.

When data is measured and recorded perfectly, a **true value** is obtained. If the value obtained from your experiment is close to the true value, then it is considered to be **accurate**. When the experiment is repeated and a set of replicate values is generated, you can assess the **precision** of the data by comparing how close (or similar) the values are to each other. Precision should not be confused with accuracy. It is important to understand the difference between these two concepts and be able to evaluate them. You can visualise this in Figure 1.

- **True value**: the value you would get if the data could be measured and recorded perfectly
- **Accuracy**: how close the value is to the true value; measurements can be described as 'more accurate' or 'less accurate' when compared with a true value
- **Precision**: how close the data values in a set are to each other; if the values in a set of data are all close (or the same), they can be described as precise

Once you determine that your data set is accurate and precise, you report this as a **measurement result**. This is usually the average of at least three trials but can also refer to a single result if only one trial was conducted.

It is also important to assess the **repeatability** of your experiment. This is whether the same data values can be produced again by the same experiment, under the same laboratory conditions. You can usually determine repeatability by conducting the same experiment at least three times.

Reproducibility is similar to repeatability. This is the ability to generate the same data values under slightly different conditions (such as with a different experiment, different measuring instrument or in a different laboratory). A reproducible experiment relies on a clear experimental method and well-defined variables. This is closely linked with accuracy.

Repeatability and reproducibility can be used to evaluate the quality or precision of measurement results.

- **Measurement result**: the final result reported in an investigation; it can refer to a single result if only one trial was conducted, or the average of at least three trials if the experiment was repeated

- **Repeatability**: whether the same data values can be produced again by the same experiment, under the same laboratory conditions
- **Reproducibility**: whether the same data values can be reproduced under slightly different conditions

Because different tools are used in experiments, **resolution** can be different. This is the smallest increment a tool can measure. For example, if the resolution of a burette is 0.5 mL, the experimenter needs to estimate the value if it lies between the marked lines. Measurement readings of 10.50 mL or 10.55 mL are possible, but you cannot claim a measurement reading of 10.53 mL.

Finally, you need to be able to assess the **validity** of an experiment. An experiment is considered valid if it investigates what it aims or claims to investigate. For example, if the controlled variables change or there is observer bias, the experiment would not be valid. Both experimental design and implementation should be considered when evaluating validity.

- **Resolution**: the smallest increment a tool can measure
- **Validity**: whether the experiment investigates what it aims or claims to investigate

FIGURE 2 In these two measurement tools, the resolution of the thermometer is 1°C, and the resolution of the beaker is 25 mL.

Measurement errors, uncertainty, significant figures and outliers

Although we may aim to measure true values when conducting an investigation, this is often not achievable because of variations and **errors**. The term 'measurement error' describes the difference between a measurement result and the true value.

error
the difference between an accepted or theoretical value and the experimental, observed or measured value

Conducting multiple trials of an investigation is a useful way to minimise the impact of measurement errors. By conducting more trials, you can record more data and therefore find average values for measurements across trials. The more trials you conduct, the more likely it is that the average value will be reliable.

The following list outlines different kinds of errors that might occur during your investigations.

- **Random errors** are unpredictable and affect the precision of a measurement. They are due to an error in the measurement process. For example, parallax error occurs when a measurement is different from the true value because it has been observed at the wrong angle. Random errors can be reduced by doing multiple trials and taking repeated measurements.

Study tip

Errors are not mistakes; they are slight changes in measurement that cannot be avoided when standard operating and/or laboratory procedures are used. Common types of errors include **reading errors**, **parallax errors** or observational errors and **instrument errors**.

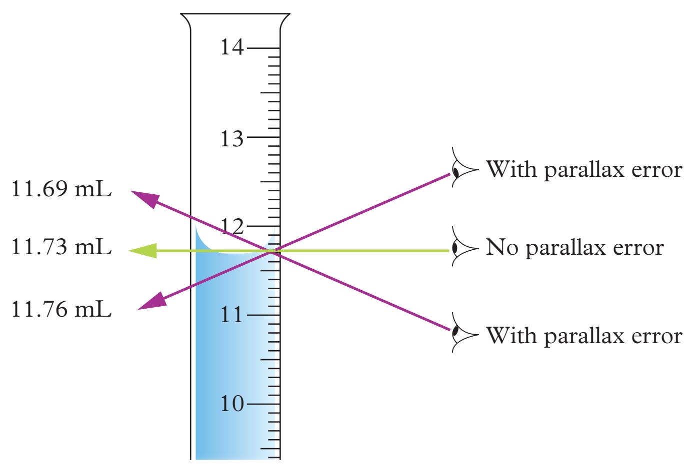


FIGURE 3 Parallax error occurs when a measurement is observed incorrectly from the wrong angle.

reading error
an error made by approximating a measurement if it falls between the markings of a measuring device

parallax error
an error made by not having your eye directly in line with the measurement

instrument error
an error due to the limited accuracy of a scientific instrument; usually, very small (and not significant) if used with standard operating and/or laboratory procedures

- **Systematic errors** are consistent and repeatable. They are caused by a problem with the method or equipment, such as an uncalibrated measuring instrument. Therefore, repeating the experiment will not reduce the impact of systematic errors.
- **Mistakes**, also called personal errors, are different from systematic or random errors. Mistakes should not be included in reporting and analysis. Instead, if you make a mistake, you need to repeat the experiment correctly. Careless measurements, or recording incorrect measurements, material spills and using incorrect equipment are examples of mistakes. Mistakes can be avoided by taking more care.
- **Uncertainty** occurs when you are unsure of the exact quantity being measured. For example, if you weigh a sample at 0.42 g, you may be uncertain if that number is 0.420 exactly, or 0.422. Each piece of equipment has a level of uncertainty.
- **Outliers** are data points or observations that are very different from the rest of the data set. Outliers need to be explained when analysing data in the discussion of a scientific report. For example, it could have been caused by an error during the investigation. Outliers should also be plotted on a graph (but they can be excluded when drawing a line of best fit). Repeating measurements can be a useful way to examine outliers.

It is important to be able to critically analyse the data that you have collected or has been provided to you. This will help you identify any contradictory or incomplete data, and recognise issues such as personal bias. This is called **data evaluation**.

data evaluation
the critical analysis of data that has been personally collected or provided to identify contradictory or incomplete data, or issues such as a personal bias

Percentage error

Percentage error refers to the difference between a measured value (e.g. one you have taken in your experiment) and a known or theoretical value. Percentage error can be a positive or negative number and is calculated as follows:

$$\%\ \text{error} = \frac{(\text{observed value}) - (\text{accepted or theoretical value})}{(\text{accepted or theoretical value})} \times 100$$

percentage error
the calculated percentage difference between an observed or measured value and an accepted or theoretical value

Significant figures

Significant figures are essential to calculations in VCE Chemistry. They are the digits of a number that are necessary to express a number accurately.

The VCAA provide the following guidelines for significant figures:

- all digits in numbers expressed in standard form are significant, e.g. 4.320×10^{-6} has 4 significant figures
- all non-zero numbers are significant, e.g. 42.3 has 3 significant figures
- zeros between two non-zero numbers are significant, e.g. 4.302 has four significant figures
- leading zeros are not significant, e.g. 0.0043 has 2 significant figures
- trailing zeros to the right of a decimal point are significant, e.g. 42.00 has 4 significant figures
- for numbers less than 1, 0.4 has 1 significant figure, 0.04 also has 1 significant figure whereas 0.40 has 2 significant figures and 0.400 has 3 significant figures
- whole numbers written without a decimal point will have the same number of significant figures as the number of digits, with the assumption that the decimal point occurs at the end of the number, for example 400 has 3 significant figures. Therefore, a stated volume of '400 mL' will be considered as having 3 significant figures.

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significant figures
the digits of a number that are used to express it to the required degree of accuracy, starting from the first non-zero digit

Study tip

If you are using values from the data book, remember to consider the number of significant figures in these values when you are rounding your final answer.

1.8 WORKED EXAMPLE

CALCULATING PERCENTAGE ERROR AND APPLYING SIGNIFICANT FIGURES

An experiment was performed to determine the density of lead and yields a value of 10.95 g cm^{-3}. The accepted value for the density of lead is 11.342 g cm^{-3}.

a Find the percentage error, showing all working.

b Round the answer to part a to four significant figures.

Solution

Think	Do
Step 1: Calculate the percentage error by substituting values into the formula: $\% \text{ error} = \frac{(\text{observed value}) - (\text{accepted or theoretical value})}{(\text{accepted or theoretical value})} \times 100$	$\% \text{ error} = \frac{10.95 - 11.342}{11.342} \times 100$ $= -0.034\,5618 \times 100$ $= -3.456\,18$
Step 2: Round –3.456 18 to four significant figures, using VCE guidelines.	–3.456%

1.8 CHECK YOUR LEARNING

Describe and explain

1 Which types of errors would affect the accuracy of data collected in a scientific investigation?

2 Explain why implementing multiple trials in scientific investigations improves the reliability of data obtained.

Apply, analyse and compare

3 Compare:

a accuracy and precision

b repeatability and reproducibility

c mistakes and errors.

4 The accepted boiling point of water is 100°C. The following boiling point values were measured by a student during an experiment under standard laboratory conditions. Comment on the accuracy and/or precision of the data obtained.

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

5 Round the following values to 3 significant figures.

a 4.078

b 36

c 02.90

Design and discuss

6 Design a revision summary tool to help you remember the application of significant figures in VCE calculations.

1.9 Constructing evidence-based arguments and conclusions

KEY IDEAS

In this topic, you will learn that:

+ analysis of raw data generated in a scientific investigation is important for constructing evidence-based arguments and conclusions
+ sentence structure and paragraph structure are important when constructing evidence-based arguments and conclusions.

Constructing evidence-based arguments and conclusions is an important key science skill required when studying VCE Chemistry. It is primarily used in the analysis, discussion and conclusion sections of the scientific investigation report.

How to construct evidence-based arguments and conclusions

Some broad questions that may be answered when constructing evidence-based arguments are:

- What is the origin of or who was responsible for the experimental results, findings, evidence or raw data collected?
- Are the experimental results, findings, evidence or raw data accurate, precise, reproducible and/or repeatable?
- What kinds of errors, inconsistencies and outliers may have affected the experimental results, findings, evidence or raw data collected?
- What background information (e.g. chemical concepts, scientific understandings, other researched information sources) was used to link or connect to the experimental results, learnings, evidence or raw data to the investigation question and to the aim?
- What series of steps or procedures could be used to improve the experimental design or methodology for future trials of this scientific investigation?

Structure of an argument

Argument structure usually follows a pattern, the length of which is determined by the number of arguments addressed. One argument usually corresponds to one paragraph.

Introduction (What is your argument/point of view?):

– Position statement (What is your hypothesis or summary of the scientific investigation?)

– List the arguments that you will make

Explain your supporting arguments:

– Supporting argument 1

» Point – identified in topic sentence

» Explanation – support with evidence, scientific finding, result or data

– Supporting argument 2:

» Point – identified in topic sentence

» Explanation – support with evidence, scientific finding, result or data

– Supporting argument 3:
  » Point – identified in topic sentence
  » Explanation – support with evidence, scientific finding, result or data

Reinforce your main point – 'To sum up...'; 'To conclude…'

An example is shown below and in Worked example 1.9.

1.9 Worked example
Find me in your obook pro

The hypothesis that increasing the volume of acid would negatively affect the pH, when titrating against a strong base, was supported.

Introduction

From the line graph of the results obtained in Table 1, the equivalence point of the titration (where the concentrations of the H^+ and OH^- ions are equal) occurred when 22 mL of acid was added and pH was 8.3. The end point of the titration was recorded at 26 mL of acid at pH 2.8. The data obtained in the experiment conformed to a strong acid/strong base titration[1], where a strong acid (e.g. HCl) was the standard solution in the burette being titrated into a strong base (e.g. sodium hydroxide, the analyte/titrand) in the conical flask.

Argument 1, with supporting fact from external source[1]

The main limitation for this experiment was that a small number of trials (one titration only) was conducted. Also, the amount of acid added before pH readings were recorded was in large increments (of 2 mL).

Argument 2, featuring limitations of the data collected

Suggestions for future improvement include conducting multiple titrations (or trials) using the same acid/base reagents, with smaller increments of acid added to the base, to obtain more reliable values for the equivalence point and end point, to within ±0.5 mL for accuracy and precision.

Argument 3, featuring suggestions for improvement

In conclusion, when a standard solution of a strong acid is titrated against a strong base, the pH decreases as the volume of acid increases, to produce an equivalence point and end point that is readily observed.

1 Clark J (November 2013) 'Titration (pH) Curves' Accessed 13 June 2022 from https://www.chemguide.co.uk/physical/acidbaseeqia/phcurves.html#top

Language features

When writing an argument, use the following language features to enhance your writing:

- Use connectives to show cause and effect; for example, 'As an outcome of ...', 'As a result of ...', 'because ...', 'consequently ...'
- Use scientific terminology and/or supporting figures.
- Use supporting facts and/or quotes from experts or other researched external sources of information; for example, 'A study conducted by XX found that …'
- Make evaluative statements; for example, 'The data shows a clear trend …', 'Evidence contradicts the argument that …'
- Use formal language and avoid personal language (such as I, you, he, she, we, they, me, him, her, us and them); for example, 'the hypothesis was supported', not 'I proved my hypothesis'.

1.9 Worked example
Video demonstration

1.9 CHECK YOUR LEARNING

Describe and explain

1 Explain why it is important to construct and understand evidence-based arguments and conclusions in science.

2 Identify three things you should (or should not) do when writing an evidence-based argument.

Design and discuss

3 Design an evidence-based analysis and conclusion for the experimental data presented in the 'Check your learning' Topic 1.7 Question 3.

1.10 Communicating

KEY IDEAS

In this topic, you will learn that:

+ effective science communication depends on knowing your audience and prioritising important information.

Science communication is not just found in your science subjects at school. It is found in everyday life: online (websites, blogs, videos, social media), in journalism, in advertising and in government policy.

Important points to remember in order to communicate science effectively:

- **Identify and understand your target audience –** presenting work for teachers or external examiners will be different from presenting to your peers.
- **Use language that is appropriate for your target audience –** if you are speaking to the general public, then limit the use of **jargon** wherever possible. If you need to incorporate **acronyms**, make sure that the full words are written in brackets next to the acronym on one occasion.
- **Only include essential content –** your communication should contain essential facts that enhance your target audiences' understanding of the science involved.
- **Use appropriate stylistic elements that are relevant to your target audience –** this may include (but is not limited to) the use of a specific presentation format (poster, report, infographic or digital media platform), diagrams, photographs, graphs, tables or other text elements (e.g. similes, metaphors or analogies). They should convey clear content and messages, and they should appeal to your audience and enhance the information so that the science is simple, easily understood and relatable without excess information or confusion
- **Use appropriate scientific literacy and conventions –** this may include (but is not limited to) the use of coherent, concise, objective and formal use of scientific language; the consistent use of tense and/or active or passive voice; definition of key scientific terms and accurate scientific representations (including models, symbols, units, balanced chemical equations and/or formulas) and a bibliography or reference list for citation of text and/or diagrams that are not of your own creation.

jargon
special words or expressions used by a profession or group that are difficult for others to understand

acronym
an abbreviation formed from the initial letters of each word from a group of words

Study tip

Spell check on your digital device is not always 100% reliable. Read, re-read or ask another person to review your work before final submission(s) so that the points of effective science communication have been covered.

Formats for communication

Science is not all about practical reports. There are many ways to communicate scientific concepts that you might like to try in VCE Chemistry. For example, you could present information as:

- an oral presentation
- a report
- an infographic
- a video or animation
- a multimodal presentation
- a scientific poster.

Scientific poster

In VCE Chemistry, you will produce a scientific poster as part of a major piece of assessment in Units 3 and 4. For this reason, some teachers will get you to practise this skill in Units 1 and 2.

The poster may be produced electronically or in hard copy and should not exceed 600 words, to meet criteria of conciseness, clarity and legibility. Note that tables, graphs and all references and/or acknowledgments are not included in the word count.

Figure 1 shows the required scientific poster format as specified by the VCAA Study Design. The centre of the poster should take up 20–25% of the poster space and contain a one-sentence summary of the major finding or outcome of the investigation that answers the investigation question.

	Title **Student name**	
Introduction Methodology and methods Results	Communication statement reporting the key finding of the investigation as a one-sentence summary	Discussion Conclusion
References and acknowledgements		

FIGURE 1 The format for the scientific poster

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1.10 CHECK YOUR LEARNING

Describe and explain

1 a Define some important elements of effective science communication.

b Describe which of these elements would be most effective for presenting your school-assessed coursework.

2 Why is it important to use appropriate scientific language and literacy throughout the poster?

Apply, analyse and compare

3 Compare the elements of science communication used in a random science video and a media article of choice. In what ways are they similar or different?

Design and discuss

4 Compare the results and the discussion sections of the poster.

5 Using appropriate science communication elements, design a presentation (of choice) for the experimental data, evidence-based analysis and conclusion created in the 'Check your learning' Topics 1.7 and 1.9, Question 3.

1.11

Sustainability

KEY IDEAS

In this topic, you will learn that:

+ studies of sustainability have been integrated into Units 1–4 of the VCE Chemistry Study Design.

sustainability
using natural resources more efficiently, to ensure supply of resources is continued into the future

Sustainability in chemistry refers to the efficiency with which we use natural resources to meet society's needs for chemical products and services. Sustainability is important so that we do not use up all of Earth's resources, leaving nothing for future generations.

It is a key focus of the VCE Chemistry Study Design. Where possible in VCE Chemistry, you should consider how you can link ideas to sustainability.

You will consider sustainability in terms of three main perspectives:

- sustainable development
- green chemistry principles
- linear and circular economies.

In this topic, you will explore the three sustainability perspectives and see how and where they fit into VCE Chemistry. These perspectives are explored throughout the student book. This includes Chapter 10 Research investigation, where you will select an investigation topic and develop a response to a research question relating to sustainability.

Sustainable development

Study tip

Further detail on the United Nations Sustainable Development Goals is found on the United Nations website (link via your obook pro).

The term 'sustainable development' is defined in a report by the World Commission on Environment and Development as development that meets the needs of people in the present without affecting the ability of people in the future to meet their needs.

This led the United Nations to identify their own Development Goals (see Table 1 and Figure 1) that address current global challenges for sustainability.

For VCE Chemistry, you need to be aware of nine of these goals and how they can be applied.

FIGURE 1 The 17 United Nations Sustainable Development Goals

TABLE 1 The nine United Nations Sustainable Development Goals relevant to VCE Chemistry and where they fit into Units 1 and 2

Goal	What would it look like to achieve this goal?	Student book chapter
2 Zero hunger	End hunger, achieve food security, improve nutrition and promote sustainable agriculture	8 (Topics 8.3, 8.4), 10
6 Clean water and sanitation	Provide sustainable access to water and sanitation to everyone	10, 11 (Topics 11.1, 11.3), 14 (Topic 14.3), 17 (Topic 17.1), 19 (Product, process or system development 11.1C, Case study 17.2B)
7 Affordable and clean energy	Provide access to affordable, reliable, sustainable and modern energy for everyone	8, 10, 13 (Topic 13.3), 16 (Topics 16.1, 16.4, 16.5)
9 Industry, innovation and infrastructure	Build resilient infrastructure, use inclusive and sustainable methods of building and production, foster innovation (i.e. new ideas)	9 (Topics 9.5, 9.6), 10, 13 (Topic 13.3), 14 (Topic 14.3), 19 (Product, process or system development 9.5)
11 Sustainable cities and communities	Make cities and human settlements inclusive, safe, resilient and sustainable	10
12 Responsible consumption and production	Ensure resources are consumed and produced sustainably	2 (Topic 2.3), 4 (Topic 4.3), 9 (Topics 9.5, 9.6), 10, 16 (Topics 16.1, 16.4, 16.5), 19 (Literature review 2.3, Fieldwork 4.3B, Product, process or system development 9.5)
13 Climate action	Take urgent action to combat the climate crisis, including drastically reducing greenhouse gas emissions	8 (Topics 8.3, 8.4), 10, 12 (Topic 12.6), 16 (Topics 16.1, 16.4, 16.5)
14 Life below water	Protect and use sustainably the oceans, seas and marine resources	10, 11 (Topics 11.1, 11.3), 12 (Topic 12.6), 17 (Topic 17.1), 19 (Case study 17.2B)
15 Life on land	Protect, restore and promote ecosystems on land, manage forests, combat desertification, and reverse land degradation and loss of biodiversity	10, 12 (Topic 12.6)

FIGURE 2 In places where energy sources are unaffordable or not accessible, people burn plastic as a source of fuel to cook food. The smoke from burning plastic produces toxic fumes. United Nations Sustainable Development Goal 7 aims to give everyone access to affordable and clean energy.

Green chemistry principles

green chemistry
an area of chemistry that focuses on designing safer and more sustainable new products

Green chemistry is an area of chemistry that focuses on designing safer and more sustainable new products. It is based on several principles that aim to reduce the impact of product or process development on the environment, resources and human well being.

The principles focus on reducing risks or hazards, the amount of materials consumed, the amount of energy used and unwanted wastes.

To achieve these aims, green chemistry relies on creativity and innovation in developing new products, processes and systems.

Seven green chemistry principles are important in VCE Chemistry (Table 2).

TABLE 2 The seven green chemistry principles important in VCE Chemistry and where they fit into Units 1 and 2

Green chemistry principle	Description	Student book chapter
Atom economy	Atom economy refers to how much of the initial reactant atoms are contained in the final product. Processes that are designed to have good atom economy maximise the incorporation of reactant materials into the final product.	2 (Topic 2.3), 4 (Topic 4.3), 10, 19 (Literature review 2.3, Fieldwork 4.3B)
Catalysis	Catalysis involves using a catalyst (such as an enzyme) to increase the rate of a chemical reaction. In green chemistry, the use of catalysis aims to generate the same product with less energy or waste in a reaction.	10
Design for degradation	Under this principle, products are designed to break down (degrade) into harmless products at the end of their life cycle. They do not persist or harm the environment.	8 (Topics 8.3, 8.4), 9 (Topics 9.5, 9.6), 10, 19 (Product, process or system development 9.5)
Design for energy efficiency	Under this principle, processes are designed to increase energy efficiency.	0, 13 (Topic 13.3), 16 (Topics 16.1, 16.4, 16.5)
Designing safer chemicals	Under this principle, chemical products are designed to be less harmful or toxic, yet still able to carry out their purpose.	8 (Topics 8.3, 8.4), 9 (Topics 9.5, 9.6), 10, 19 (Product, process or system development 9.5)
Prevention of wastes	This principle seeks to prevent waste rather than treating or cleaning it up after it has been created.	4 (Topic 4.3), 9 (Topics 9.5, 9.6), 10, 13 (Topic 13.3), 14 (Topic 14.3), 16 (Topics 16.1, 16.4, 16.5), 17 (Topic 17.1), 19 (Fieldwork 4.3B, Product, process or system development 9.5, Case study 17.2B)
Use of renewable feedstocks	A raw material or a **feedstock** (used for processing another product) should be made from sustainable materials (e.g. plant-based), rather than fossil fuels.	8 (Topics 8.3, 8.4), 9 (Topics 9.5, 9.6), 10, 16 (Topics 16.1, 16.4, 16.5)

feedstock
a raw material used to produce other goods

Linear and circular economies

Natural resources are becoming scarce. Humans are using and disposing of more products than we have ever done in the past. This is due to growing global populations, but also due to global spending habits.

As a society, we must rethink the way we consume and dispose of products to make sure our natural resources are available in the future.

In VCE Chemistry, you may be asked to compare how a **linear economy** and a **circular economy** make use of our resources.

linear economy
a way of managing resources that operates on a take–make–dispose model; new resources are used and disposed of after use

circular economy
a model of production that involves sharing, leasing, re-using, repairing, refurbishing and recycling existing materials and products as long as possible

Linear economy

A linear economy uses a 'take–make–dispose' model. Resources are extracted from nature or the Earth to make products that will be thrown away (sometimes after only one use). A simple representation is shown in Figure 3.

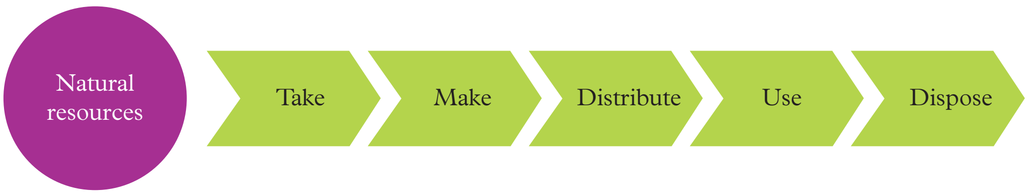

FIGURE 3 A representation of a linear economy

Circular economy

A **circular economy** is a more sustainable approach. It follows a continuous cycle that focuses on using and re-using resources repeatedly at different stages of production and consumption (Figure 4).

Linear and circular economy concepts are explored in the following chapters in the student book:

- Chapter 2 (Topic 2.3)
- Chapter 4 (Topic 4.3)
- Chapter 8 (Topics 8.3, 8.4)
- Chapter 9 (Topics 9.5, 9.6)
- Chapter 10
- Chapter 13 (Topic 13.3)
- Chapter 16 (Topics 16.1, 16.4, 16.5)
- Chapter 19 (Fieldwork 4.3B, Product, process or system development 9.5).

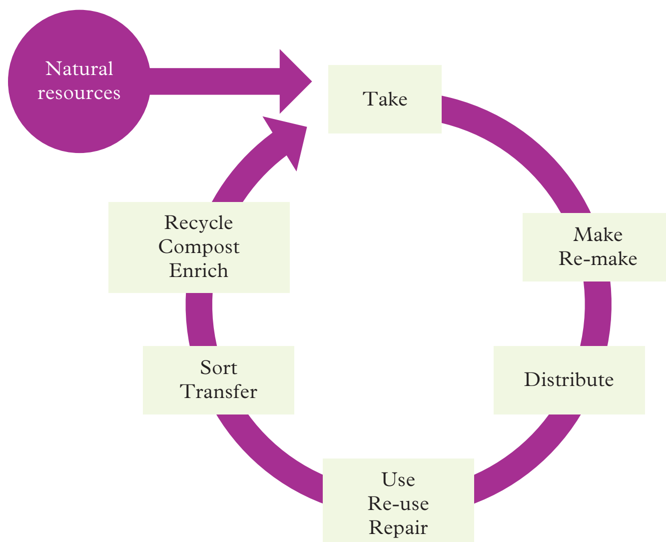

FIGURE 4 A representation of a circular economy

1.11 CHECK YOUR LEARNING

Describe and explain

1 Describe and explain which principles of green chemistry are the most relevant to studying VCE Chemistry.

2 Explain why the transition from a linear to a circular economy is being used as a strategy for more sustainable development.

Design and discuss

3 Some hazardous substances that have been minimised or eliminated by green chemistry processes include carbonyl chloride (for the manufacture of bisphenol A in plastics), sodium salts of alkyl benzene sulfonic acids (for detergents), sodium phosphate builders (for removing calcium and magnesium ions from hard water), chlorofluorocarbons (for refrigeration, air conditioning and aerosol propellants) and phosphorus- or chlorine-based pesticides (for crop pest control).

a Select one of the hazardous substances mentioned and research how green chemistry was used to replace it.

b Present your findings to your class.

FIGURE 5 You will learn about metal recycling as an example of a shift from a linear to a circular economy.

1.12 Balancing chemical equations

KEY IDEAS

In this topic, you will learn that:

+ balanced chemical equations obey the law of conservation of mass
+ construction and balancing of chemical equations is a prerequisite for completing quantitative calculations in VCE Chemistry.

Constructing and balancing chemical equations is an important skill in VCE Chemistry. School assessment questions often ask you to 'write the balanced chemical equation for the reaction between …' before you start the quantitative calculations.

The law of conservation of mass

law of conservation of mass
a law that states that in an isolated system, mass cannot be created or destroyed

isolated system
a thermodynamic system that cannot exchange matter or energy outside its boundaries

The **law of conservation of mass** states that in an **isolated system**, during a chemical reaction, mass is conserved. It cannot be created or destroyed.

This means that in chemical reactions, the chemical bonds between the reactants and products are broken and rearranged, so that the number of atoms involved remains unchanged. Balanced chemical equations obey the law of conservation of mass. But the number of atoms present does not change.

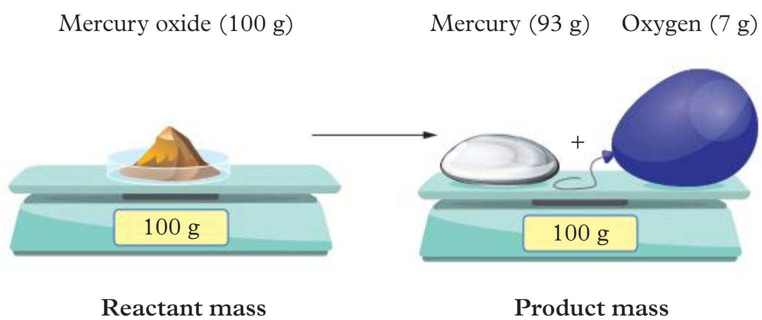

FIGURE 1 When mercury oxide (HgO) is heated, its atoms rearrange to form liquid mercury (Hg) and oxygen gas (O_2). The mass of the reactants and products is the same.

This concept is important to remember when balancing equations. There needs to be the same number of molecules on the reactant side as the product side.

Steps for balancing chemical equations

You can follow these steps when asked to balance a chemical equation.

Step 1 Write the equation with the known products and reactants

Write the chemical formulas for the reactants and products. These formulas cannot be changed to make the equation balance (i.e. do not change the subscripts).

$$CH_4 + O_2 \rightarrow H_2O + CO_2$$

This equation is **not** balanced.

Step 2 Tally the atoms

Tally how many atoms of each element are on either side of the arrow.

CH_4	+	O_2	$\rightarrow$	H_2O	+	CO_2
C = 1		C = 0		C = 0		C = 1
H = 4		H = 0		H = 2		H = 0
O = 0		O = 2		O = 1		O = 2

Reactants	Products
C = 1 + 0 = 1	C = 0 + 1 = 1
H = 4 + 0 = 4	H = 2 + 0 = 2
O = 0 + 2 = 2	O = 1 + 2 = 3

In this example, the number of carbon atoms is balanced (one on each side) but the numbers of hydrogen and oxygen atoms are not balanced.

Step 3 Add coefficients

Put numbers (coefficients) in **front** of the molecules to add further molecules. These numbers multiply all the atoms in the chemical formula (i.e. putting a 2 in front of H_2O means there are now twice as many H and twice as many O atoms).

Balance one element at a time.

Balance hydrogen, by adding the coefficient 2 to the product side.

$$CH_4 + O_2 \rightarrow 2H_2O + CO_2$$

CH_4	+	O_2	$\rightarrow$	H_2O	+	CO_2

Reactants	Products
C = 1 + 0 = 1	C = 0 + 1 = 1
H = 4 + 0 = 4	H = 2 + 0 = 2 × 2 = 4
O = 0 + 2 = 2	O = 1 × 2 + 2 = 4

Balance oxygen by adding the coefficient 2 to the reactant side.

$$CH_4 + 2O_2 \rightarrow 2\ H_2O + CO_2$$

CH_4	+	O_2	$\rightarrow$	H_2O	+	CO_2

Reactants	Products
C = 1 + 0 = 1	C = 0 + 1 = 1
H = 4 + 0 = 4	H = 2 + 0 = 2 × 2 = 4
O = 0 + 2 × 2 = 4	O = 1 × 2 + 2 = 4

There are now equal amounts of C, H and O on both sides of the equation.

Step 4 Check your work

Check your work by re-counting all types of atoms on both sides of the arrow (this equation is now **balanced**).

Figure 2 is a visual representation for this chemical reaction.

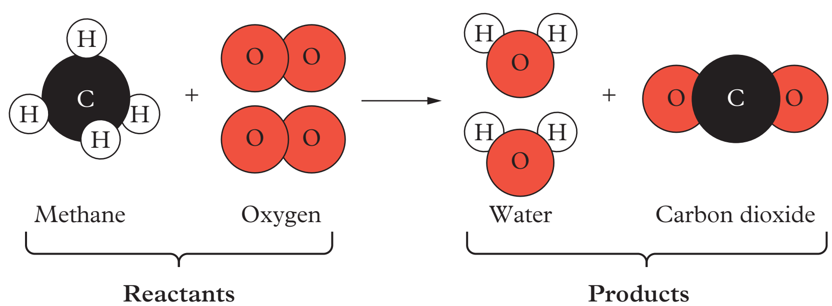

FIGURE 2 The reactants (methane and oxygen) have the same number of C, H and O atoms as the products (water and carbon dioxide) of this reaction.

Step 5 Add the states

If you know the states, place the following symbols after the formulas; (g) for a gas, (l) for a liquid, (s) for a solid, (aq) for an aqueous solution (in water).

For the previous equation for the complete combustion of methane, we would write:

$$CH_4(g) + 2O_2(g) \rightarrow 2H_2O(g) + CO_2(g)$$

In this case, the methane is being burned at a high temperature, so the water will be in a gaseous state.

1.12 CHECK YOUR LEARNING

Describe and explain

1 Explain why balanced chemical equations obey the law of conservation of mass.

Apply, analyse and compare

2 Balance the following chemical equations:

a $Na(s) + Cl_2(g) \rightarrow NaCl(s)$

b $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

c $Ag_2CO_3(s) \rightarrow Ag_2O(s) + CO_2(g)$

d $C_2H_4(g) + O_2(g) \rightarrow CO(g) + H_2O(g)$

e $P_4(s) + O_2(g) \rightarrow P_2O_5(s)$

f $Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$

g $HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

h $Ca(CN)_2(s) + HCl(aq) \rightarrow HCN(aq) + CaCl_2(aq)$

i $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$

j $Mg(s) + O_2(g) \rightarrow MgO(s)$

Design and discuss

3 Using Figure 2 as an example, design a visual representation of one of the balanced chemical equations in Question 2.

FIGURE 3 A balanced equation has the same number of each type of atom on either side of the arrow.

1.13 Preparing for assessment

KEY IDEAS

In this topic, you will learn that:

+ organisational skills are important for revision
+ understanding the meaning of command terms (e.g. describe, discuss, identify) will help you to answer questions well in assessments.

Preparing for any assessment in VCE Chemistry requires organisational skills that can be practised throughout the year. It is a good idea to practise your revision skills, in particular, because these will become important in Units 3 and 4.

Organisational skills

The organisational skills that are useful for VCE studies include:

- creating a timetable for your studies, homework and other commitments
- setting SMART goals
- incorporating reflection and/or evaluation of strategies for continuous improvement.

Creating a timetable

It is important to have a balance when you are in VCE. Make sure you have time to study, continue your extracurricular activities (such as hobbies or a part-time job) and rest properly.

1.13 Study plan template
Find me in your obook pro

Follow the link in your obook pro to access a template you could use to create a study timetable, or you could use information technology resources to create your own version. Make sure you block out the time spent at school (including travel), playing sport, for family commitments, having meals/snacks and at part-time work. Be realistic – look at the amount of time you have left; you may have to make some sacrifices to succeed in your VCE studies without sacrificing your mental health.

For the remaining time, allocate homework and/or study time for each subject, making sure that you are actively applying the skills learnt in class rather than only trying to remember the content.

Setting SMART goals

Setting and achieving goals in VCE studies improves your motivation to succeed. Goals are most effective when they are SMART:

- Specific – clear and explicit
- Measurable – can be measured or is a measure of success
- Attainable – can be challenging (for learning and growth) but reachable
- Relevant – are meaningful and worthwhile (i.e. what is to be accomplished)
- Time-based – are committed to deadlines or set times.

When setting your SMART goals, consider your strengths and weaknesses, as well as your thoughts and feelings about subject(s). An example of a SMART goal would be 'By the end of Term 1, I would like to achieve an average of B+ or higher for my internal assessment tasks in Chemistry' rather than the less specific 'improve in Chemistry'.

Reflection

Every few weeks, pause to reflect on your progress and the effectiveness of your organisational approach. Aim for continuous improvement – if something is not quite working, then change to a more reliable (and repeatable) strategy. Celebrate small successes and continue to challenge yourself by setting new goals.

The advantages of having good organisational skills include improving your marks in school-assessed coursework and reducing your stress, worry, **procrastination** and uncertainty levels. It will allow you to work smarter (not harder), avoid last-minute panics and gain a sense of accomplishment, in the knowledge that you did your best.

procrastination
the action of delaying or postponing something

Examination tips

If you are sitting an exam in Units 1 and 2, there are a few strategies that could help you prepare.

Before the exam:

- Start revising early. Cramming is stressful!
- Eat a healthy diet and cut back on processed or junk food and sugar-laden beverages.
- If you can, have a separate area where you revise or study from where you sleep because a good night's sleep is essential in the lead-up to the exam to help aid concentration.
- On the day of the examination, eat breakfast so that you are not distracted by hunger or the noise of your rumbling stomach.
- Take a bottle of water with the labels removed to maintain hydration.
- Arrive in plenty time before the examination; 15 minutes before the start of reading time is usually sufficient.
- Allow time for a bathroom visit before entering the examination venue.

During the exam:

- Use your reading time effectively – read the instructions and the questions carefully and plan out your responses.
- Once writing time has started, think about quickly jotting down some of the ideas that came to mind during reading time, so that you don't forget them later!
- Make use of your extra writing spaces or ask for a spare piece of paper.
- Don't dwell on questions you find difficult. Move on and return to them once you've finished the other questions well.
- On the multiple-choice questions, eliminate those answer choices that you think are incorrect.
- Beware of careless mistakes such as missing units, incorrect decimal places in conversions and using incorrect formulas and significant figures in answers.
- When drawing chemical compounds, chemical bond types or semi-structural formulas, use large, clear and accurate diagrams.

After you finish learning all of the content in Units 1 and 2, your teacher may assign you a practice exam that tests your knowledge and skills on everything you've learnt!

Study tip

Studying for your VCE exams should begin during the first week of the first topic, as smaller and consistent study sessions create better student outcomes (higher marks) than longer and irregular study sessions.

Understand command terms

Make sure you understand what the command terms mean, so that you can provide succinct and appropriate responses. Table 1 lists the VCAA VCE command terms that you may encounter in assessment tasks and exams.

TABLE 1 Command terms

Term	Explanation
account of	Describe a series of events or transactions.
account for	State reasons for; report on.
analyse	Identify components/elements and the significance of the relationship between them; draw out and relate implications; determine logic and reasonableness of information.
apply	Use; employ in a particular situation or context.
assess	Make a judgment about, or measure, determine or estimate, the value, quality, outcomes, results, size, significance, nature or extent of something.
calculate	Determine from given facts, figures or information; obtain a numerical answer showing the relevant stages in the working; determine or find (e.g. a number, answer) by using mathematical processes.
clarify	Make a statement or situation more comprehensible.
compare	Recognise similarities and differences and the significance of these similarities and differences.
construct	Make, build, create or put together by arranging ideas or items (e.g. an argument, artefact or solution); display information in a diagrammatic or logical form.
contrast	Show how things are different or opposite.
deduce	Draw a conclusion from given information, data, a narrative, an argument, an opinion, a design and/or a plan.
define	Give the precise meaning and identify essential qualities of a word, phrase, concept or physical quantity.
demonstrate	Show ideas, how something can be done or that something is true by using examples or practical applications, or by applying algorithms or formulas.
describe	Provide characteristics, features and qualities of a given concept, opinion, situation, event, process, effect, argument, narrative, text, experiment, artwork, performance piece or other artefact in an accurate way.
discuss	Present a clear, considered and balanced argument or prose that identifies issues and shows the strengths and weaknesses of, or points for and against, one or more arguments, concepts, factors, hypotheses, narratives and/or opinions.
distinguish	Make clear the differences between two or more arguments, concepts, opinions, narratives, artefacts, data points, trends and/or items.
evaluate	Ascertain the value or amount of; make a judgment using the information supplied, criteria and/or own knowledge and understanding to consider a logical argument and/or supporting evidence for and against different points, arguments, concepts, processes, opinions or other information.
examine	Consider an argument, concept, debate, data point, trend or artefact in a way that identifies assumptions, possibilities and interrelationships.
explain	Give a detailed account of why and/or how with reference to causes, effects, continuity, change, reasons or mechanisms; make the relationships between things evident.
extract	Select relevant and/or appropriate detail from an argument, issue or artefact.
extrapolate	Infer and/or extend information that may not be clearly stated from a narrative, opinion, graph or image by assuming existing trends will continue.
identify	Recognise and name and/or select an event, feature, ingredient, element, speaker and/or part from a list or extended narrative or argument, or within a diagram, structure, artwork or experiment.
infer	Derive conclusions from available information or evidence, or through reasoning, rather than through explicit statements.
interpret	Draw meaning from an argument, point of view, description or diagram, text, image or artwork and determine significance within context.
investigate	Observe, study or carry out an examination in order to establish facts and reach new conclusions.
justify	Show, prove or defend, with reasoning and evidence, an argument, decision and/or point of view using given data and/or other information.
list	Provide a series of related words, names, numbers or items that are arranged consecutively.

(continued)

TABLE 1 continued

Term	Explanation
name	Provide a word or term (something that is known and distinguished from other people or things) used to identify an object, person, thing, place etc.
outline	Provide an overview or the main features of an argument, point of view, text, narrative, diagram or image.
persuade	Induce (someone) to do something through reasoning or argument; convince.
predict	Give an expected result of an upcoming action or event; suggest what may happen based on available information.
propose	Suggest or put forward a point of view, idea, argument, diagram, plan and/or suggestion based on given data or stimulus material for consideration or action.
recall	Present remembered ideas, facts and/or experiences.
recommend	Put forward and/or approve (someone or something) as being suitable for a particular purpose or role.
recount	Retell a series of events or steps in a process, usually in order.
state	Give a specific name or value or other brief answer without explanation or calculation.
suggest	Put forward for consideration a solution, hypothesis, idea or other possible answer.
summarise	Retell concisely the relevant and major details of one or more arguments, text, narratives, methodologies, processes, outcomes and/or sequences of events.
synthesise	Combine various elements to make a whole or an overall point.

Source: *Glossary of command terms* reproduced by permission © VCAA

1.13 CHECK YOUR LEARNING

Describe and explain

1 Describe at least two strategies that can be useful for answering examination questions.

2 Explain why having good organisational skills is important for VCE studies.

Apply, analyse and compare

3 Analyse the revision strategies that you currently use and develop a SMART goal for improvement in this area.

4 Contrast the following command terms.

a Deduce and calculate

b Outline and identify

c Evaluate and justify

Design and discuss

5 Use ICT resources to create and print your personal study timetable, using a template similar to that shown in Figure 1. Keep it visible and stick to it!

STUDY PLAN

Mon /	Tues /	Wed /	Thur /	Fri /	Sat /	Sun /
Day-	Day-	Day-	Day-	Day-	Day-	Day-
06:00						
07:00						
08:00						
09:00						
10:00						
11:00						
12:00						
13:00						
14:00						
15:00						
16:00						
17:00						
18:00						
19:00						
20:00						
21:00						
22:00						
23:00						
24:00						
To-do list						
•	•	•	•	•	•	•
•	•	•	•	•	•	•
•	•	•	•	•	•	•
•	•	•	•	•	•	•

FIGURE 1 An example of a study plan template

CHAPTER 1 Review

Chapter summary

1.1
- The VCE Chemistry course has a clear structure.
- There are many career pathways available from studying Chemistry.

1.2
- Aboriginal and Torres Strait Islander Peoples were Australia's first chemists and applied their unique scientific knowledge about the natural world to find (or refine) suitable food sources and develop traditional medicines and other useful materials in sustainable ways.

1.3
- The key science skills and their application are important for success in VCE Chemistry.
- A research question states the specific problem or issue on which your investigation will be based.
- An aim is a statement of what is to be investigated.
- A hypothesis is a testable statement that may include a prediction of an outcome of an investigation.

1.4
- A variety of methodologies can be used for a scientific investigation.
- A logbook is required for recording raw data and must be submitted intermittently for assessment purposes.
- The scientific poster has a specific format in VCE Science.

1.5
- Safety is your priority when performing scientific investigations.
- A risk assessment is an organised way of identifying hazards and/or risk factors and implementing controls for prevention when performing scientific investigations.

1.6
- Skills and knowledge of ethical understanding are applied when performing all types of scientific investigations in VCE Chemistry.

1.7
- Data generated from scientific investigations is important evidence to support the trends, patterns and/or relationships between variables.
- Raw data can be qualitative or quantitative, discrete or continuous.
- Raw data from scientific investigations needs to be presented clearly so that it can be easily understood by your audience.

1.8
- Investigations that generate raw data must be valid, repeatable and reproducible.
- Errors and outliers must be included and accounted for in data evaluations.
- Significant figures must be considered in all calculations during VCE Chemistry.

1.9
- Analysis of raw data generated in a scientific investigation is important for constructing evidence-based arguments and conclusions.
- Sentence and paragraph structure is important for constructing evidence-based arguments and conclusions.

1.10
- Effective science communication depends on knowing your audience, avoiding jargon and prioritising information.

1.11
- The three perspectives of sustainability in VCE Chemistry are sustainable development, green chemistry principles, and linear and circular economies.

1.12
- Balanced chemical equations obey the law of conservation of mass.
- Constructing and balancing chemical equations is a prerequisite for completing quantitative calculations in VCE Chemistry.

1.13
- Organisational skills are important for assessment preparation.
- Understanding command terms can help you to answer questions properly on assessment tasks.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• understand the structure of the VCE Chemistry course, including the Areas of Study, their Outcomes, and Assessments	☐	☐	☐	Go back to Topic 1.1
• understand that Aboriginal and Torres Strait Islander Peoples' knowledges and perspectives have informed and enhanced scientific thinking	☐	☐	☐	Go back to Topic 1.2
• develop aims and questions, formulate hypotheses and make predictions	☐	☐	☐	Go back to Topic 1.3
• plan and conduct investigations, including using a logbook	☐	☐	☐	Go back to Topic 1.4
• describe the eight key scientific investigation methodologies relevant to VCE Chemistry	☐	☐	☐	Go back to Topic 1.4
• comply with safety and ethical guidelines, including understanding basic lab safety rules and the importance of risk assessments	☐	☐	☐	Go back to Topic 1.5
• describe the role of ethical understanding in VCE Chemistry	☐	☐	☐	Go back to Topic 1.6
• generate, collate and record data, including differentiating between different types of data (qualitative, quantitative, discrete and continuous data) and using graphs to represent data	☐	☐	☐	Go back to Topic 1.7
• analyse and evaluate data and investigation methods, including considering the validity, repeatability and reproducibility of data, identifying errors and outliers, and using significant figures	☐	☐	☐	Go back to Topic 1.8
• construct evidence-based arguments and draw conclusions	☐	☐	☐	Go back to Topic 1.9
• analyse, evaluate and communicate scientific ideas	☐	☐	☐	Go back to Topic 1.10
• describe the three different perspectives on sustainability, including sustainable development, green chemistry principles, and linear and circular economies	☐	☐	☐	Go back to Topic 1.11
• balance chemical equations	☐	☐	☐	Go back to Topic 1.12
• understand the importance of organisational skills in preparing for assessments, including creating and using a timetable, setting SMART goals and reflecting on progress	☐	☐	☐	Go back to Topic 1.13

Revision questions

Multiple choice

1 The law of conservation of mass means that:
   A chemical formulas are conserved during a reaction.
   B there are different types of atoms in the reactants from those of the products.
   C the coefficients used to balance chemical equations.
   D the mass of the reactants must equal the mass of the products.

2 The balanced chemical equation for the burning of propane in oxygen is:
   A $C_3H_8 + 10O_2 \rightarrow 3CO_2 + 4H_2O$
   B $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   C $C_3H_8 + 13O_2 \rightarrow 3CO_2 + 5H_2O$
   D $2C_3H_8 + 13O_2 \rightarrow 6CO_2 + 10H_2O$

3 Which of the following best describes the required VCAA format for the scientific poster?
   A Aim, hypothesis, methodology, results, discussion, conclusion, references
   B Introduction, methodology, results, communication statement, discussion, conclusion, references
   C Title, student name, introduction, methodology, results, communication statement, discussion, conclusion, references and acknowledgements
   D Abstract, introduction, methodology, results, discussion, conclusion, acknowledgements

4 What is the variable that is purposely changed in an experiment or practical activity?
   A Controlled variable
   B Dependent variable
   C Independent variable
   D Random variable

5 How many significant figures are in the measurement: Temperature = 0.0230°C?
   A 2
   B 3
   C 4
   D 5

6 Poor precision in scientific measurements may arise from:
   A theoretical values being too inflexible.
   B human error.
   C limitations of the measuring instrument.
   D both human error and the limitations of the measuring instrument.

7 If an accepted value is 15.63 and a student measured 12.82, 12.96, 13.02 and 12.99 in subsequent trials of an experiment, their data would be described as:
   A accurate but not precise.
   B precise but not accurate.
   C both accurate and precise.
   D neither accurate nor precise.

8 Deduce which of the following is not compatible with the principles of green chemistry.
   A Maximising energy use in each process
   B Minimising the use of toxic chemicals by replacing them with safer alternatives
   C Maximising the atom efficiency of each reaction pathway
   D Minimising the formation of wastes and by-products

Short answer

Describe and explain

9 Describe at least two examples of chemical processes used by Aboriginal or Torres Strait Islander Peoples for food sources, medicinal or construction purposes that are not mentioned in Topic 1.2.

10 Explain some of the language features used for constructing evidence-based arguments for scientific investigations.

Apply, analyse and compare

11 A student set up the equipment in Figure 1 to determine how the measured volume of a gas changed with the temperature. The results of the practical investigation are given in Table 1.

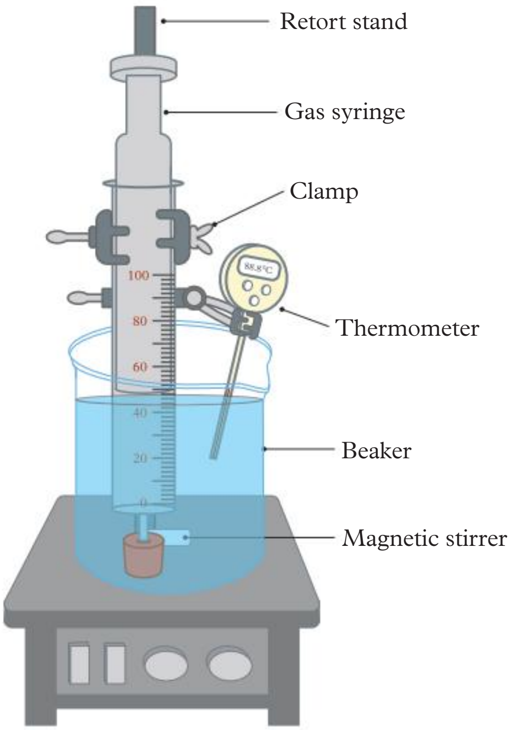

FIGURE 1 Equipment set up for the practical investigation

TABLE 1 Results of the practical investigation

Temperature (°C)	Volume (mL)
88.8	49
81.3	47
76.5	47
71.2	47
67.9	47
62.7	46
57.2	46
51.9	45
47.1	45
43.9	44
41.2	43

a Identify the dependent and the independent variables in this practical investigation.

b Write an aim for this practical investigation.

c Formulate a hypothesis for this practical investigation.

d Create a visual representation of the data in Table 1 (with appropriate labels) to identify the relationship between the dependent and independent variables in this experiment.

Design and discuss

12 Design an evidence-based analysis and conclusion for the data presented in Table 1 and the visual representation from Question 11d.

13 Design a risk assessment for the practical investigation in Question 11.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 1

HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED?

FIGURE 1 Polymers are large molecules made up of repeating units called monomers.

Unit 1 Overview

In Unit 1 of VCE Chemistry, you will explore the chemical structures and properties of materials, how they are measured and how innovations in manufacturing can lead to the development of more sustainable products.

Unit 1 Areas of Study

The learning for this unit has been divided into three Areas of Study. The table below shows how each Area of Study aligns with the chapters in this book and lists the page numbers for each chapter.

Area of Study	Chapter		Pages
Area of Study 1 How do the chemical structures of materials explain their properties and reactions?	**Chapter 2**	Elements and the periodic table	50–73
	Chapter 3	Covalent substances	74–117
	Chapter 4	Reactions of metals	118–139
	Chapter 5	Reactions of ionic compounds	140–161
	Chapter 6	Separation and identification of the components of mixtures	162–185
	Unit 1 Area of Study 1 Checkpoint		186–189
Area of Study 2 How are materials quantified and classified?	**Chapter 7**	Quantifying atoms and compounds	190–211
	Chapter 8	Families of organic compounds	212–249
	Chapter 9	Polymers and society	250–289
	Unit 1 Area of Study 2 Checkpoint		290–293
Area of Study 3 How can chemical principles be applied to create a more sustainable future?	**Chapter 10**	Research investigation	pro

Unit 1 Outcomes

In this unit, you will:

→ explain how elements form carbon compounds, metallic lattices and ionic compounds, experimentally investigate and model the properties of different materials, and use chromatography to separate the components of mixtures

→ calculate mole quantities, use systematic nomenclature to name organic compounds, explain how polymers can be designed for a purpose, and evaluate the consequences for human health and the environment of the production of organic materials and polymers

→ investigate and explain how chemical knowledge is used to create a more sustainable future in relation to the production or use of a selected material.

CHAPTER 2

Elements and the periodic table

KEY KNOWLEDGE

- → the definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons
- → the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements
- → critical elements (for example, helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery

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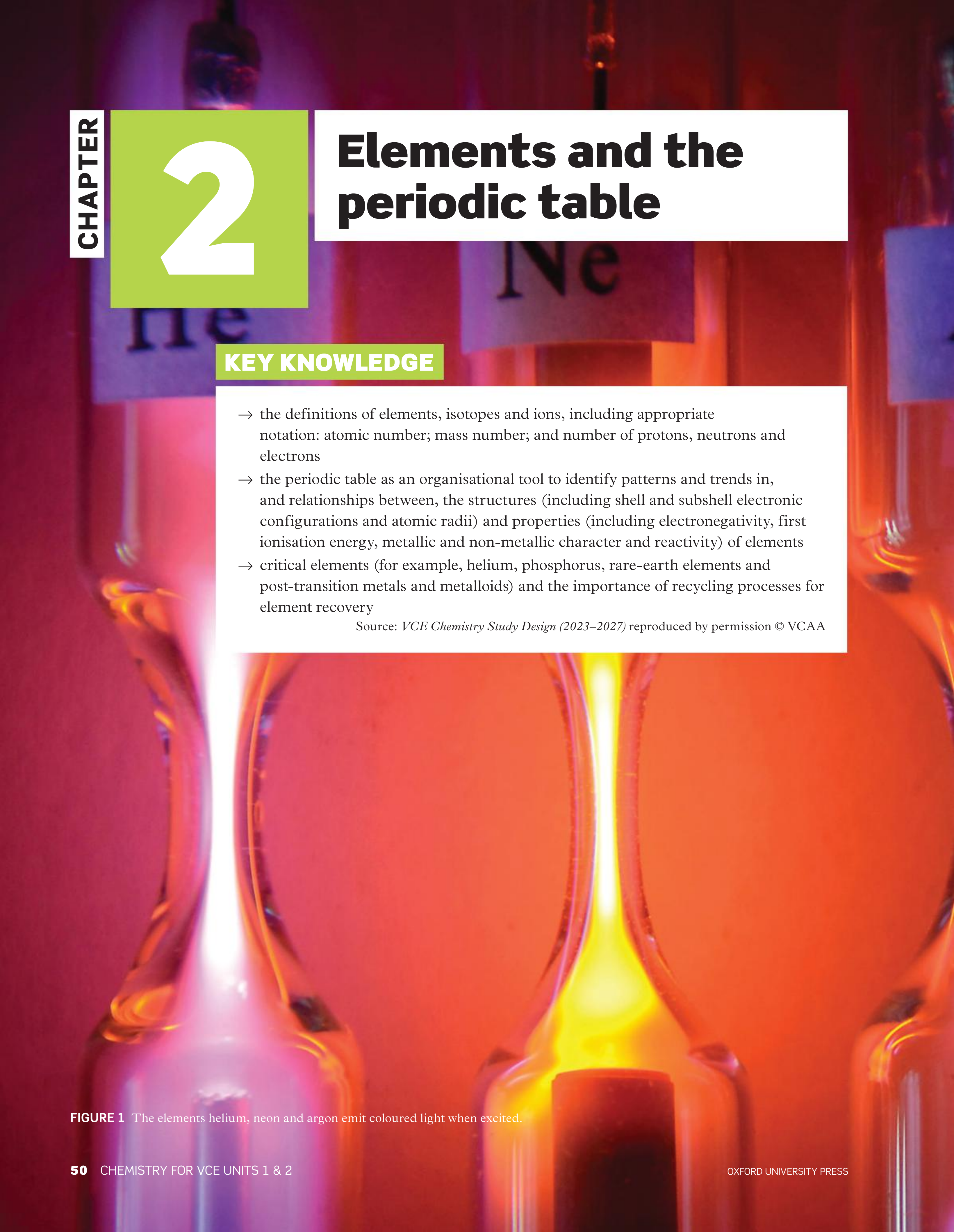
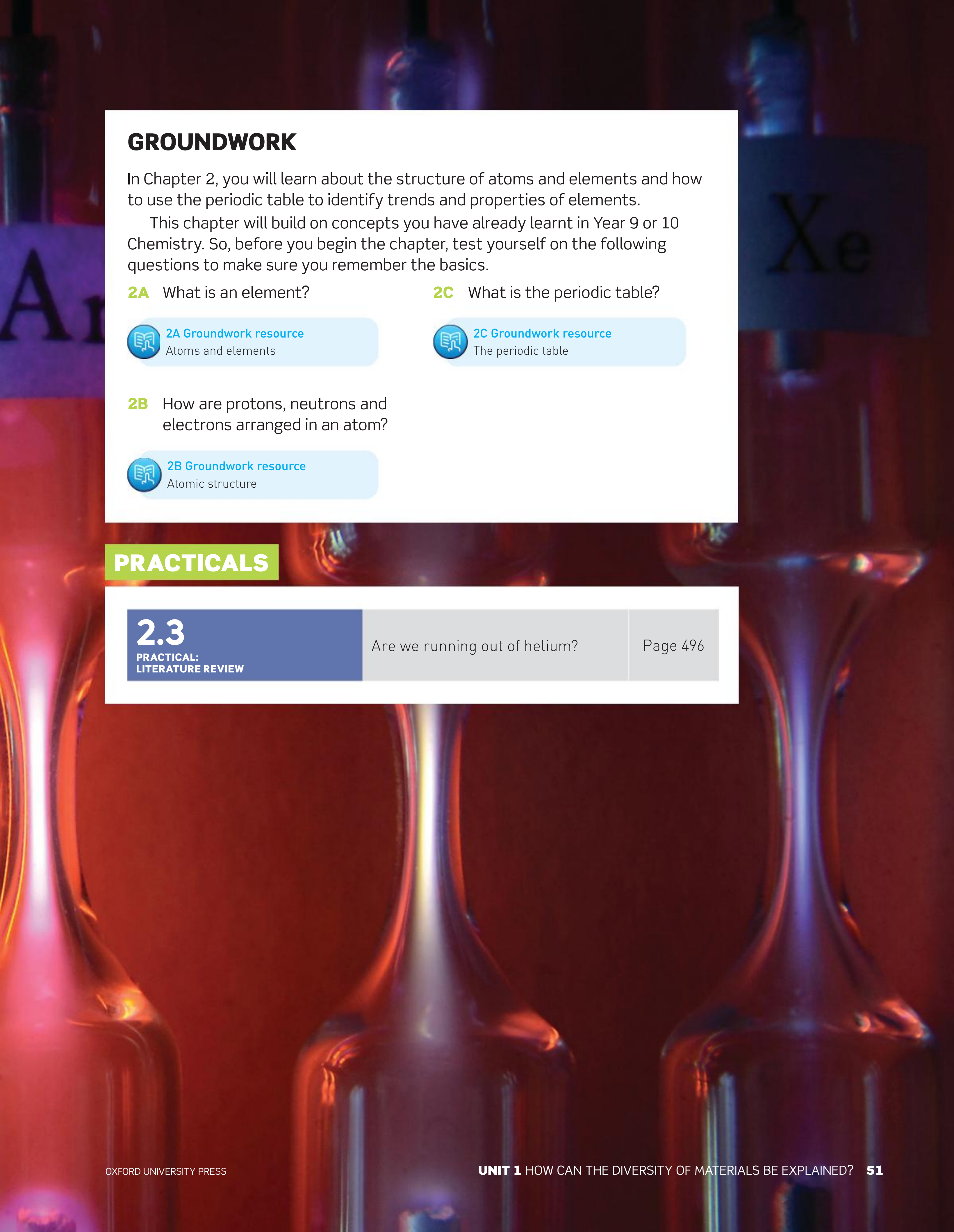


FIGURE 1 The elements helium, neon and argon emit coloured light when excited.



GROUNDWORK

In Chapter 2, you will learn about the structure of atoms and elements and how to use the periodic table to identify trends and properties of elements.

This chapter will build on concepts you have already learnt in Year 9 or 10 Chemistry. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

2A What is an element?

2A Groundwork resource
Atoms and elements

2B How are protons, neutrons and electrons arranged in an atom?

2B Groundwork resource
Atomic structure

2C What is the periodic table?

2C Groundwork resource
The periodic table

PRACTICALS

2.3 PRACTICAL: LITERATURE REVIEW	Are we running out of helium?	Page 496

2.1 Elements, isotopes and ions

KEY IDEAS

In this topic, you will learn that:

+ each of the 118 elements in the periodic table has a unique number of protons in the nucleus
+ atoms consist of protons and neutrons in the nucleus as well as electrons that orbit around the nucleus in shells
+ isotopes are atoms with the same atomic number but different mass numbers because they have different numbers of neutrons
+ an ion is a charged atom that has either lost or gained electrons.

atom
the smallest unit of matter, which consists of protons, neutrons and electrons

element
a pure substance that consists of only one type of atom

subatomic particle
a particle that makes up an atom, e.g. a proton, a neutron and an electron

proton
a positively charged subatomic particle

neutron
a neutral subatomic particle

electron
a negatively charged subatomic particle

nucleus
the dense region of an atom, which consists of protons and neutrons

electrostatic attraction
the attraction force between oppositely charged particles

atomic number
the number of protons in an atom

Almost everything we see and interact with is made from **atoms**, the smallest units and building blocks of matter.

Pure substances that consist of only one type of atom are called **elements**. Scientists have so far identified 118 elements, which have been organised into a periodic table. Of the 118 elements discovered so far, 90 are found naturally in our solar system and the remainder have been synthesised artificially. In this topic, you will look at the structure and components of elements and atoms.

Atomic structure

Atoms are made up of three main **subatomic particles**:

- **protons** (positively charged)
- **neutrons** (neutral)
- **electrons** (negatively charged).

Each atom has a dense **nucleus**, which contains positively charged protons and neutral neutrons. Because neutrons are neutral, the nucleus has an overall positive charge from the presence of protons. Orbiting outside the nucleus are the electrons (Figure 1).

Electrons are held close to the nucleus as a result of the forces of attraction between the negative electrons and the positive nucleus. The force of attraction between oppositely charged particles is called **electrostatic attraction**.

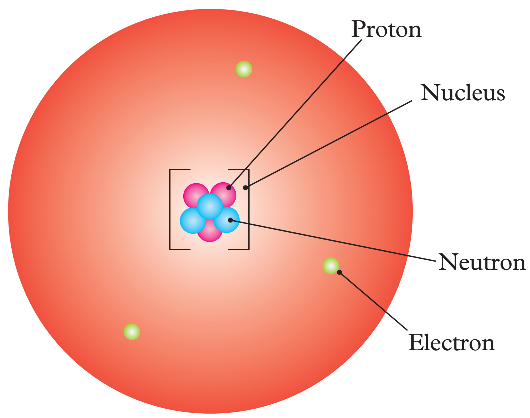


FIGURE 1 The basic model of an atom

The atomic number of elements

Each element has a unique number of protons in its nucleus, called its **atomic number**. For example, any atom with 79 protons in its nucleus is always gold and any atom with 30 protons in its nucleus is always zinc. Elements are arranged in increasing order of atomic number in the periodic table.

Each element also has its own name and chemical symbol consisting of one or two letters. The first letter is always capitalised and the second letter (if there is one) is always lower case. Chemical symbols do not always match the English name of the elements. This is because some symbols are derived from other languages. For example, tungsten is W, from its German name, *wolfram*; and iron is Fe, from its Latin name, *ferrum*.

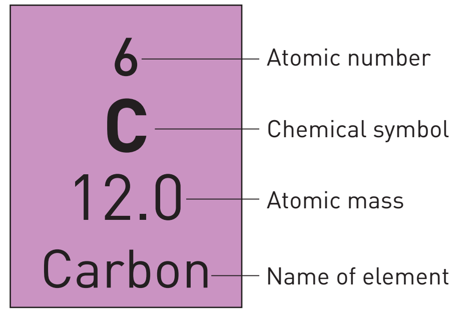


FIGURE 2 The element carbon has an atomic number of 6, atomic mass of 12.0 and chemical symbol C.

Protons, electrons and neutrons

Protons and neutrons exist in the nucleus and have a mass of 1 atomic mass unit (amu). The sum of protons and neutrons in an element give its **mass number**. The electrons are much smaller and have a mass approximately 1837 times less than that of a proton or neutron.

mass number
the sum of protons and neutrons in an atom

TABLE 1 A comparison of subatomic particles

Subatomic particle	Location	Charge	Relative mass
Proton	Nucleus	Positive (1+)	1
Neutron	Nucleus	Neutral (0)	1
Electron	Electron shells outside the nucleus	Negative (1–)	0.000 544*

*The relative mass of an electron is so low that we approximate it to zero when calculating mass numbers.

Isotopes

Isotopes are atoms with the same number of protons but different numbers of neutrons. This means they have the same atomic number but different mass number. Isotopes have the same chemical properties, such as reactivity, but different physical properties, such as density.

isotopes
atoms of the same element that have different numbers of neutrons

TABLE 2 Isotopes of lithium have the same number of protons but different numbers of neutrons

Isotope	In the nucleus		Outside the nucleus in shells
	Protons	Neutrons	Electrons
Lithium-6	3	3	3
Lithium-7	3	4	3
Lithium-8	3	5	3

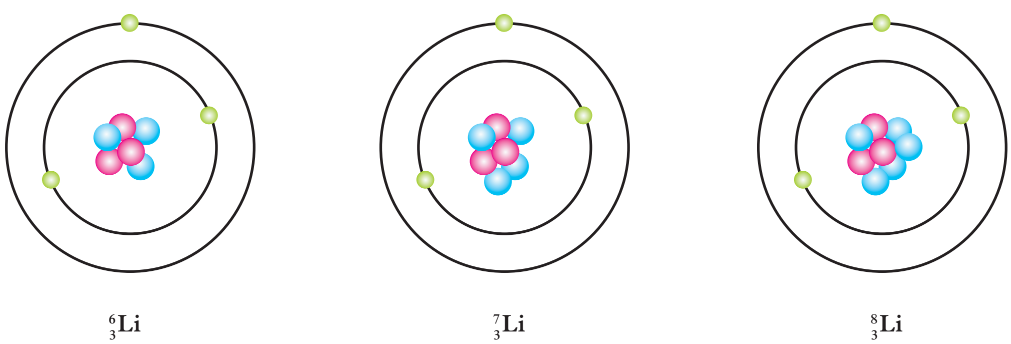


FIGURE 3 Lithium-6, lithium-7 and lithium-8 are isotopes.

The first 20 elements have similar numbers of protons and neutrons in their nuclei. Heavier elements (atomic number ≥ 20) usually have more neutrons than protons. Stable isotopes such as carbon-12 and carbon-13 have nuclei that do not break down over time. Radioactive isotopes have too many or too few neutrons and undergo radioactive decay. Radioactive isotopes are also called **radioisotopes** because they break down over time and emit radiation.

radioisotope
an isotope with too many or too few neutrons, which breaks down by undergoing radioactive decay

TABLE 3 A comparison of the properties of isotopes of any element

Isotopes have the same…	Isotopes have different…
number of protons	number of neutrons
atomic number	mass number
chemical properties	physical properties
element name and symbol	abundances in nature

Ions

ion
a charged atom formed by gaining or losing electrons

cation
a positively charged ion formed after an atom loses electrons

anion
a negatively charged ion formed after an atom gains electrons

Ions are atoms that have either a positive or a negative charge. Ions form when atoms gain or lose electrons. When an atom loses electrons, it forms a positively charged **cation**. When an atom gains electrons, it forms a negatively charged **anion**. The number of electrons lost or gained is called the charge of the ion.

TABLE 4 A comparison of anions, atoms and cations

Anion (–)	Atom	Cation (+)
Has **gained** electrons	Number of electrons **equals** number of protons	Has **lost** electrons
Has a **negative** charge	Has **no charge** (is electrically neutral)	Has a **positive** charge

Nuclide notation review

nuclide notation
an abbreviated way to show basic information about an atom or ion, e.g. $^{7}_{3}Li^{+}$

Atoms are represented by using **nuclide notation** (Figure 4) where:

- ***A*** represents the mass number of the nuclide, which is the number of protons + neutrons
- ***Z*** represents the atomic number of the nuclide, which is the number of protons only
- **X** is the elemental symbol, which depends entirely on the atomic number of the element.

Ions are written in this same style but include a charge written at the top right of the element symbol (Figure 5). This charge equals the number of electrons *lost* or *gained*. A positive charge indicates that the ion has lost electrons. A negative charge indicates that the ion has gained electrons. No charge is written for atoms because all atoms are neutral.

Mass number A

Atomic number Z

$$^{A}_{Z}\mathrm{X}$$

Element symbol

FIGURE 4 Nuclide notation

$$^{24}_{12}\mathrm{Mg}^{2+}$$

FIGURE 5 The nuclide notation of a magnesium cation: mass number = 24, atomic number = 12 and the atom has lost 2 electrons, giving the ion a 2+ charge.

TABLE 5 Examples of some atoms and ions written in nuclide notation

Atom or ion in nuclide notation	Number of protons	Number of neutrons	Number of electrons	Charge
$^{1}_{1}H$	1	0	1	0 (atom)
$^{25}_{12}Mg$	12	13	12	0 (atom)
$^{37}_{17}Cl^{-}$	17	20	18	1– (anion)
$^{56}_{26}Fe^{2+}$	26	30	24	2+ (cation)
$^{197}_{79}Au$	79	118	79	0 (atom)

Study tips

The number of neutrons can be calculated as mass number minus atomic number:

$$N = A - Z$$

The number of electrons can be calculated as atomic number minus charge:

$$e^{-} = Z - \text{charge}$$

Atoms are always neutral. Ions are always charged.

2.1 Real-world chemistry
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2.1 WORKED EXAMPLE

CALCULATING THE NUMBER OF NEUTRONS AND ELECTRONS IN AN ATOM OR ION

Calculate the number of neutrons and electrons in $^{24}_{12}Mg^{2+}$.

Solution

Think	Do
Step 1: Calculate the number of neutrons by subtracting atomic number from the mass number.	The number of neutrons in $^{24}_{12}Mg^{2+}$ = 24 (mass number) – 12 (atomic number) = 12 neutrons
Step 2: Calculate the number of electrons by subtracting the charge from the atomic number.	The number of electrons in $^{24}_{12}Mg^{2+}$ = 12 (atomic number) – 2 (charge) = 10 electrons

2.1 CHECK YOUR LEARNING

Describe and explain

1 Identify the charge of each subatomic particle.
 a Proton
 b Neutron
 c Electron

2 Describe the location of protons, neutrons and electrons in an atom.

3 Explain why the nucleus of an atom has a net positive charge.

Apply, analyse and compare

4 Write the following in nuclide notation.
 a Nitrogen-14
 b Sodium-23
 c Krypton-82
 d Magnesium-25

5 Calculate approximately how many electrons equal the mass of one proton.

6 Calculate the number of protons, neutrons and electrons in the following.
 a $^{16}_{8}O^{2-}$
 b $^{27}_{13}Al$
 c $^{79}_{35}Br^{-}$

Design and discuss

7 Draw a Venn diagram to show the similarities and differences between a lithium atom and a lithium-8 isotope.

2.2 The periodic table

KEY IDEAS

In this topic, you will learn that:

- the 118 elements are organised in the periodic table in order of increasing atomic number
- elements in the same group (column) of the periodic table have similar chemical properties; elements in the same period (row) of the periodic table have the same number of electron shells
- electron configurations can be written using shell notation (e.g. 2,8,8,1) or subshell notation (e.g. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$)
- there are several trends in the periodic table from left to right and from top to bottom, which can be explained by differences in the core charge and number of electron shells of each element.

Origins of the periodic table

The periodic table is usually credited to Dmitri Mendeleev, a 19th-century Russian chemist. Mendeleev organised approximately 60 elements into a periodic table based on their reactivities. Like many great scientists, Mendeleev did not work alone. Mendeleev's periodic table built on the work of other scientists who also placed elements into groups based on their reactivity (Figure 1).

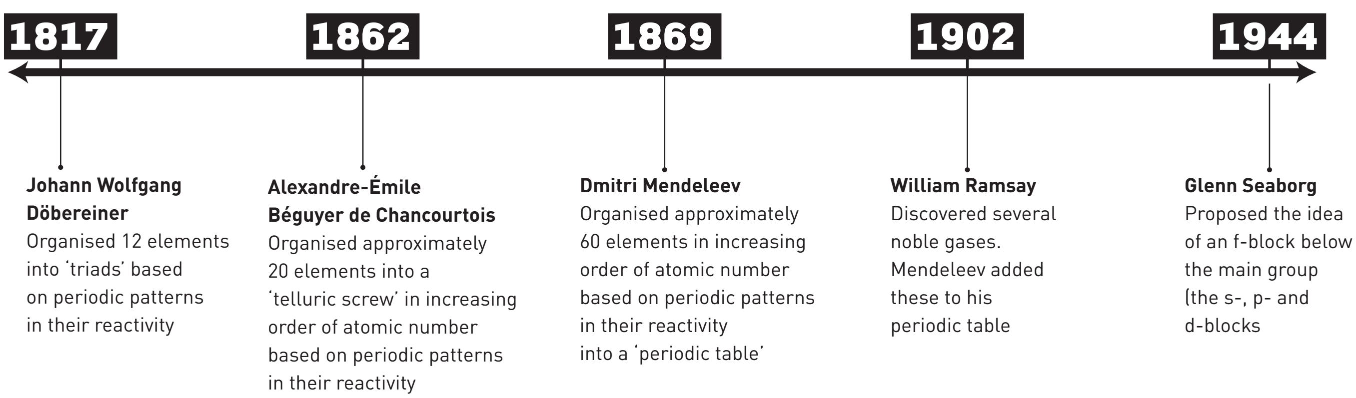

FIGURE 1 Contributions from some scientists to the development of the modern periodic table

The structure of the periodic table

valence electron an electron in the outermost shell of an atom

The modern periodic table contains 118 elements in increasing order of atomic number (Figure 2). Each group (column) contains elements with similar chemical properties and the same number of **valence electrons**. Each period (row) contains elements with the same number of electron shells.

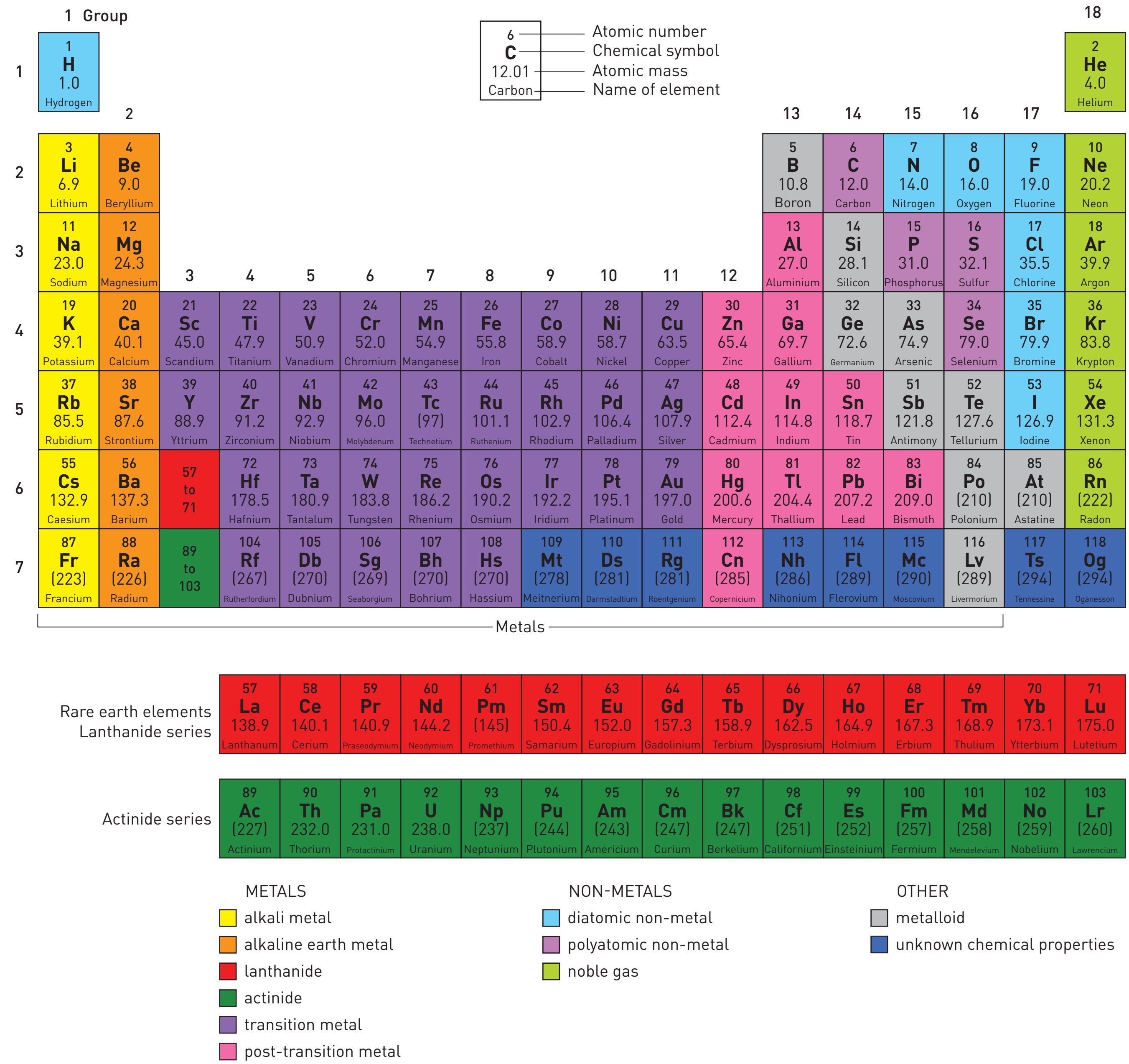

FIGURE 2 The periodic table of elements

TABLE 1 Groups of the periodic table

Group number	Group name	Number of valence electrons of elements in this group
1	Alkali metals	1
2	Alkaline earth metals	2
11	Coinage metals	1
3–11	Transition metals	Variable
12–15	Post-transition metals	Variable
15	Pnictogens	5
16	Chalcogens	6
17	Halogens	7
18	Noble gases	8

Groups

Groups of the periodic table are numbered from 1 to 18. Elements in the same group have the same number of electrons in their valence shells and some of the groups have names as well as numbers. For example, group 1 is called the alkali metals because they react violently with acids. Group 17 is called the halogens (‘salt-makers’ in Greek) because they react with metals to make salts.

Periods

The only similarity between elements in a period is the number of electron shells they have. For example, elements in period 5 (from rubidium to xenon) all have five shells of electrons. Other than the number of shells of electrons, elements in a single period have very little in common. The properties of elements in a period vary dramatically.

Blocks s, p, d and f of the periodic table

There are four blocks in the periodic table (Figure 3).

- The s-block consists of groups 1 and 2 – these are usually metals that react with acids and steam.
- The d-block consists of groups 3–12 – this block is also known as the transition block because it contains transition metals such as silver.
- The p-block consists of groups 13–18 – elements in this block are the non-metals. Elements in group 18 are the chemically inert noble gases.
- The f-block consists of the two rows at the bottom of the table – this block includes the lanthanides and actinides.

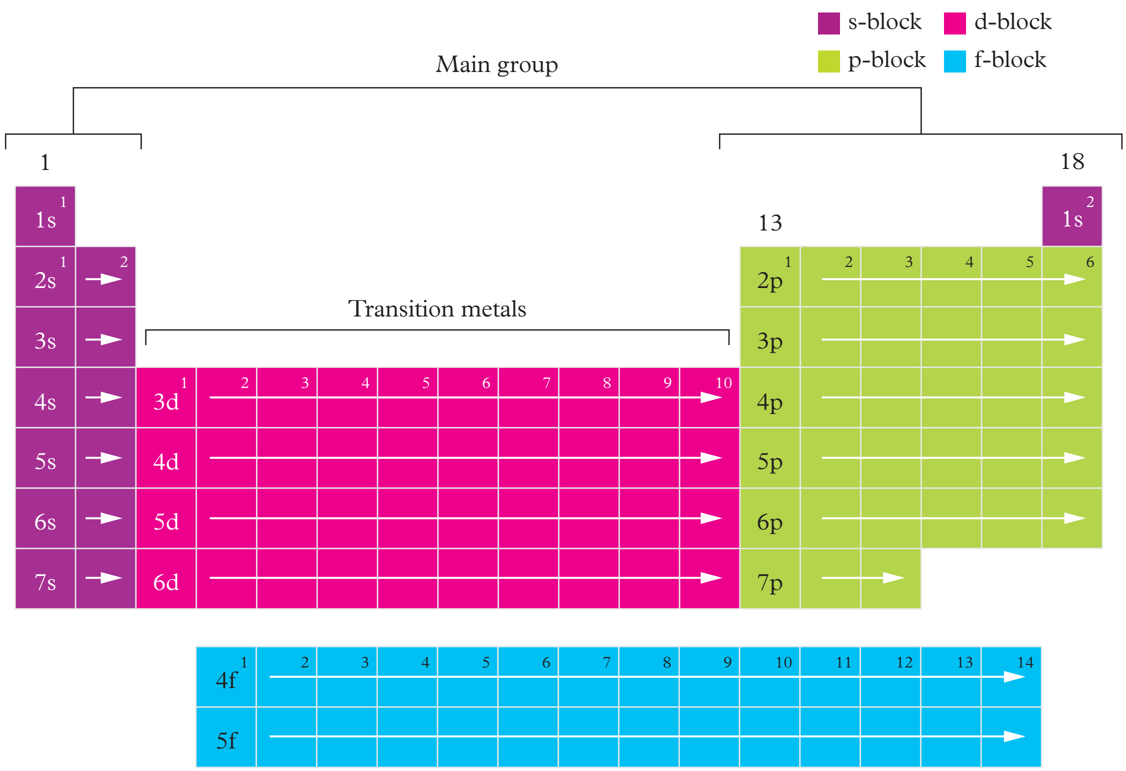

FIGURE 3 Blocks s, p, d and f in the periodic table

Electron shell configurations

In 1913, after studying the absorption and emission spectra of hydrogen, scientist Niels Bohr developed a model to explain the arrangement of electrons in an atom.

Bohr's model suggested that electrons:

- orbit around an atom's nucleus
- exist in fixed energy levels, which sit at specific distances from the nucleus
- can move between energy levels but to do so they must gain or lose energy.

Shells

After Bohr proposed his model for hydrogen, scientists began applying this model to other elements. They named the energy levels that Bohr proposed electrons were grouped in 'electron shells'. Electron shells are numbered outwards from the nucleus (where 1 = closest to nucleus) and can each hold a maximum number of electrons. Table 2 summarises the number of electrons each shell can hold.

TABLE 2 A summary of the maximum number of electrons each electron shell can hold

Shell number (n)	Maximum number of electrons held
1	2
2	8
3	18
4	32
n	$2n^2$

Writing electron shell configuration

Electron shell configuration can be written by using the following set of rules:

1 Each shell can contain a maximum number of electrons.

2 Lower energy shells (those closest to the nucleus) are filled before the higher energy shells.

3 Electron shells are filled in this specific order:
- The first two electrons go into the first shell.
- The next eight electrons go into the second shell.
- The next eight electrons go into the third shell.
- Two electrons are then placed into the fourth shell; after this, the remainder of the third shell is filled.

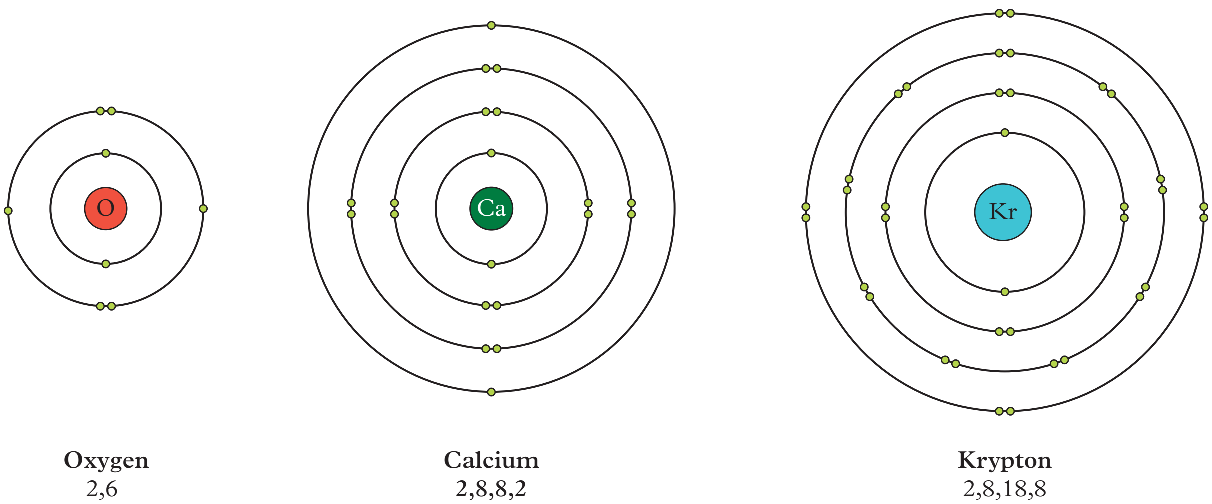

FIGURE 4 Bohr diagrams of oxygen, calcium and krypton showing the placement of electrons in shells. Their shell configurations are written underneath.

Electron subshell configurations

In 1926, Austrian physicist Erwin Schrödinger noticed that Bohr's model of electrons orbiting in shells was incomplete. Bohr's model did not explain why electrons do not crash into the nucleus. Schrödinger developed complex quantum mechanical equations to describe the true movement of electrons around atomic nuclei. We still use his quantum mechanical model of the atom today.

Subshells

Schrödinger's quantum mechanical model of the atom has electrons in specific energy levels which correspond to the 'shells' in Bohr's model. Within each shell, there are one, two, three or four **subshells**, each of which has its own name consisting of a digit and a letter. The first 10 subshells are 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d.

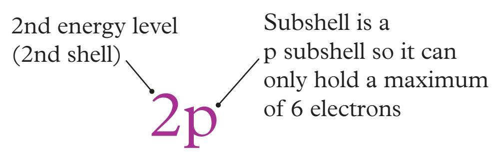

FIGURE 5 Subshell notation

subshell
a subdivision of an electron shell that can hold a maximum of 2, 6, 10 or 14 electrons depending on its type (s, p, d or f)

In each of these subshell names, the digit represents the energy level (or 'shell') in Bohr's model. The letter afterwards represents the type of subshell. Each type of subshell is a different shape and can hold a different maximum number of electrons. This is summarised in Table 3.

TABLE 3 A summary of the different subshells and the number of electrons they can hold

Subshell	Maximum number of electrons it can hold
s	2
p	6
d	10
f	14

orbital
a subdivision of an electron shell where two electrons can orbit the nucleus; each electron shell contains 1, 3, 5 or 7 orbitals

Orbitals

Each subshell contains one, three, five or seven **orbitals**, which are each able to hold a maximum of two electrons. This is summarised in Table 4.

TABLE 4 The number of orbitals and maximum number of electrons that can be held in each type of subshell

Subshell type	Number of orbitals present	Maximum number of electrons it can hold
s	1	2
p	3	6
d	5	10
f	7	14

Study tips

You can easily obtain the order of the subshells by quickly drawing Figure 6 during a test or exam

Aufbau principle
a principle that states electrons always occupy available orbitals with the lowest energy

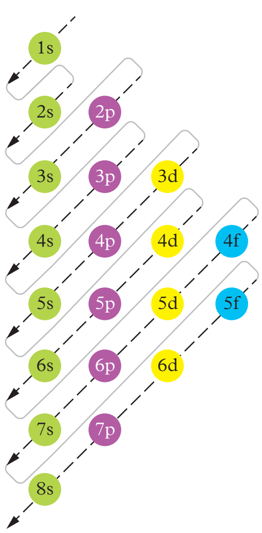

FIGURE 6 The Aufbau principle states that each electron occupies the lowest energy orbital available. The long zig-zag arrow shows the order in which electron subshells are filled.

Aufbau principle

The **Aufbau principle** describes how electrons always occupy available orbitals with the lowest energy. For example, the 1s subshell, which contains one orbital and can hold a maximum of two electrons, has the lowest energy and is always filled first. The 2s subshell has the second lowest energy and is always filled second. After 2s is filled, the 2p orbital will be filled before the 3s orbital because 2p has the lower energy level (Figure 6).

The order in which to fill subshells can be easily remembered by drawing Figure 6. Read the subshells diagonally from top right to bottom left to obtain the correct order in which subshells are filled by electrons.

To write the **electronic configuration** of an atom, write each subshell with the number of electrons it contains in superscript. Electron configurations of the first 30 elements are shown in Table 5. Notice how the electrons occupy subshells in order from lowest to highest energy.

electronic configuration
the distribution of electrons in an atom

TABLE 5 Schrödinger (subshell) electron configurations of the first 30 elements

Element	Symbol	*Z*	Electron configuration	
			Bohr (shells)	Schrödinger (subshells)
Hydrogen	H	1	1	$1s^1$
Helium	He	2	2	$1s^2$
Lithium	Li	3	2,1	$1s^2\ 2s^1$
Beryllium	Be	4	2,2	$1s^2\ 2s^2$
Boron	B	5	2,3	$1s^2\ 2s^2\ 2p^1$
Carbon	C	6	2,4	$1s^2\ 2s^2\ 2p^2$
Nitrogen	N	7	2,5	$1s^2\ 2s^2\ 2p^3$
Oxygen	O	8	2,6	$1s^2\ 2s^2\ 2p^4$
Fluorine	F	9	2,7	$1s^2\ 2s^2\ 2p^5$
Neon	Ne	10	2,8	$1s^2\ 2s^2\ 2p^6$
Sodium	Na	11	2,8,1	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium	Mg	12	2,8,2	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminium	Al	13	2,8,3	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon	Si	14	2,8,4	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
Phosphorus	P	15	2,8,5	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Sulfur	S	16	2,8,6	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Chlorine	Cl	17	2,8,7	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Argon	Ar	18	2,8,8	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Potassium	K	19	2,8,8,1	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Calcium	Ca	20	2,8,8,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
Scandium	Sc	21	2,8,9,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$
Titanium	Ti	22	2,8,10,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$
Vanadium	V	23	2,8,11,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$
Chromium	Cr	24	2,8,13,1	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^1\ 3d^5}$
Manganese	Mn	25	2,8,13,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$
Iron	Fe	26	2,8,14,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$
Cobalt	Co	27	2,8,15,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$
Nickel	Ni	28	2,8,16,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$
Copper	Cu	29	2,8,18,1	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^1\ 3d^{10}}$
Zinc	Zn	30	2,8,18,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

Study tips

Some textbooks and websites write the subshells so the third shell is grouped together; for example, $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $\mathbf{3d^3}$, $\mathbf{4s^2}$ instead of $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $\mathbf{4s^2}$, $\mathbf{3d^3}$.

Exceptions: copper and chromium

Some elements between hydrogen and xenon have electron configurations that do not entirely follow the Aufbau principle. For example, we would expect chromium and copper to have electron configurations ending in $3d^4$ and $3d^9$ respectively. However, because $3d^4$ and $3d^9$ are unstable electron configurations in neutral atoms, one electron from the 4s subshell in chromium and copper moves to occupy the 3d subshell. This results in $3d^5$ and $3d^{10}$ electron configurations for chromium and copper. You need to memorise the electron configurations of these two exceptions to the Aufbau principle.

Electron configurations of ions

Anions gain electrons in Aufbau order

When atoms gain electrons and form anions, the extra electrons continue to fill subshells according to the Aufbau principle. For example, when a neutral atom of fluorine (F), $1s^2\ 2s^2\ 2p^5$, gains one electron to form a fluoride ion (F^-), $1s^2\ 2s^2\ 2p^6$, the extra electron occupies the next available orbital within the 2p subshell.

Cations lose their outermost s electrons first

Atoms that lose electrons form cations. The outermost s electrons are lost first, followed by the rest of the electrons in reverse Aufbau order. For example:

Co: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$ (neutral atom in ground state)

Co^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$ (the two 4s electrons have been removed)

Co^{3+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4$ (the two 4s electrons and one 3d electron have been removed)

Another example is iron:

Fe: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$ (neutral atom in ground state)

Fe^+: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^6$ (one 4s electron has been removed)

Fe^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$ (the two 4s electrons have been removed)

Fe^{3+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$ (the two 4s electrons and one 3d electron have been removed)

Note that although $3d^4$ and $3d^9$ are unstable electron configurations in neutral atoms, they can be stable electron configurations in charged ions.

Electron configurations of atoms and ions in excited states

excited state
the state of an atom when electrons are not in their lowest energy orbital

Heat, light and electrical energy can excite atoms and ions. Being in an **excited state** means that one or more of the electrons is not in the lowest energy orbital: it is instead promoted to a higher energy orbital. We denote the excited state with an asterisk (*).

For example, a sodium ion (Na^+) may absorb energy when heated and have any one of its electrons promoted to a higher energy level, forming a sodium ion in an excited state, Na^{+*}.

$$Na^+:\ 1s^2\ 2s^2\ 2p^6$$

$$Na^{+*}:\ 1s^2\ 2s^2\ 2p^5\ 3s^1$$

There are also many more possibilities for the electron configuration of Na^{+*}, including $1s^2\ 2s^2\ 2p^5\ 3p^1$ or $1s^2\ 2s1\ 2p^6\ 3s^1$.

ground state
the state of an atom when electrons are in their lowest energy state as predicted by the Aufbau principle

Atoms and ions in an excited state are unstable and quickly revert to the **ground state** (the lowest energy configuration of electrons).

Trends in the periodic table

The organisation of the periodic table informs us about more than an atom's make up and electron configuration. The table also considers relative properties of elements, which can also be used to identify trends. In this topic, we will look at how core charge and shielding relate to horizontal and vertical trends. We will also look at key trends that can be read from the table including:

- atomic radius
- metallic character
- electronegativity
- reactivity
- ionisation energy.

Core charge

Core charge is the attractive force between the valence electrons and the nucleus of the atom. An inner electron shell is any electron shell that is not a valence electron shell. Core charge can be calculated using the following equation:

Core charge = number of protons – number of inner electrons

For example, a sodium atom has 11 protons and 10 inner electrons; therefore, the core charge of sodium equals 11 – 10 = 1+. The core charge of a neutral atom is always equal to the number of valence electrons it has. Horizontal trends in the periodic table are explained by differences in core charge. From left to right across a period, the core charge increases, which means electrons are held more closely to the nucleus.

core charge
the number of protons minus the number of inner electrons

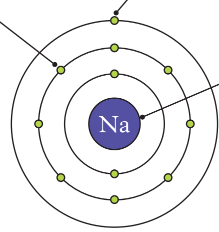

FIGURE 7 A sodium atom has a core charge of 1+.

Shielding effect

The **shielding effect** is the reduction in core charge experienced by valence electrons because of the presence of inner electrons. The shielding effect increases as the number of inner shell electrons or **inner shells** increase.

Vertical trends in the periodic table are explained by differences in the shielding effect. As you move down any group, the number of inner shells increases and so does the shielding effect. The more inner shells an atom has, the further valence electrons are held from the nucleus.

shielding effect
the shielding of valence electrons by inner electrons that alters the nuclear charge felt by valence electrons

inner shell
any shell of electrons that is not a valence shell

Atomic radius

Atomic radius is defined as half the distance between two nuclei of the same element that are bonded together (Figure 8). It is often used to describe the size of an atom. As you move down a group, atomic radius increases because of the shielding effect. The atomic radius trend is shown in Figure 9 on page 64.

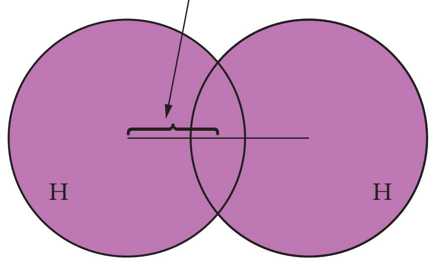

FIGURE 8 How the atomic radius of helium is determined

atomic radius
half the distance between two nuclei of the same element that are bonded together

metallic character
how readily an atom loses a valence electron

Metallic and non-metallic character

Metallic character describes how readily an atom loses a valence electron. Group 1 metals have the greatest metallic character because they lose electrons most readily. Non-metals have the lowest metallic character because they tend not to lose electrons (instead, they accept electrons). As such, **non-metallic character** describes how readily an element will accept electrons.

non-metallic character
how readily an atom gains a valence electron

Metallic character increases as you move down a group and decreases from left to right across a period. Overall, there is an increase in metallic character from top right diagonally to bottom left of the table. Conversely, there is an increase in non-metallic character from bottom left of the table diagonally up to top right. The metallic character and non-metallic character trends are shown in Figure 9.

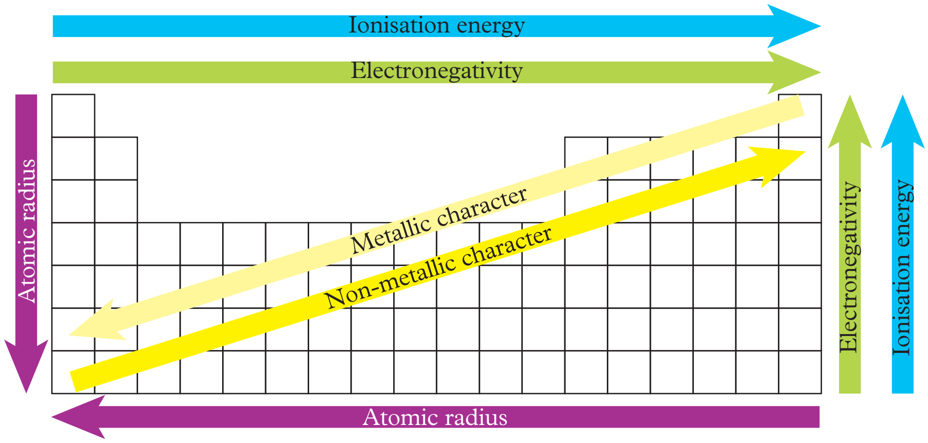

FIGURE 9 A summary of periodic table trends

Electronegativity

electronegativity the tendency for an atom to attract electrons when in a chemical bond

Electronegativity is the tendency for an atom to attract electrons when in a chemical bond. Metals have low electronegativity because of their low core charge and large atomic radius. Because of their low electronegativity, metals tend to donate electrons when they form ionic bonds.

In contrast, non-metals have high electronegativity and tend to accept electrons when they form ionic bonds. Electronegativity tends to decrease down a group and increase from left to right across a period (see Figures 9 and 10).

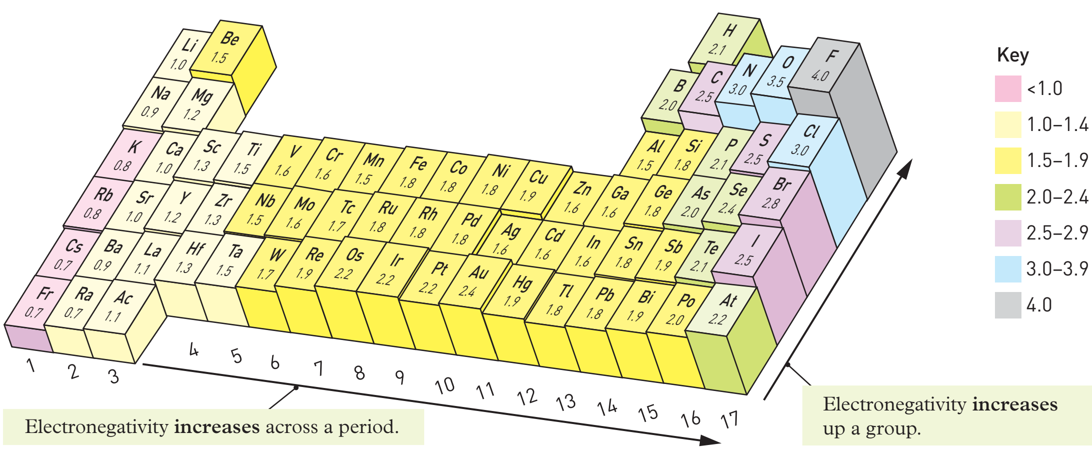

FIGURE 10 A summary of the electronegativity trends across the periodic table and the approximate electronegativity values of elements

2.2 Skill drill
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Reactivity

reactivity how readily a chemical substance undergoes a chemical reaction

Reactivity is the tendency for an element to react with itself or other elements. The most reactive elements in the periodic table are the elements with the highest and lowest electronegativity. Fluorine and francium are the most reactive elements in the periodic table. Elements in group 18 have a full valence shell, which makes them inert and, in most instances, unreactive.

2.2 Challenge
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First ionisation energy trend

First ionisation energy is defined as the amount of energy required to remove one valence electron. Metals in group 1 have the lowest first ionisation energy because after losing one electron, they achieve a full outer shell. Non-metals have the highest first ionisation energy because in losing one valence electron, they become even further from having a full outer shell. You can see the first ionisation energy trend in Table 6.

first ionisation energy
the energy required to remove the first valence electron

TABLE 6 A summary of explanations for each of the four trends in the periodic table in terms of core charge and shielding effect

Trend	Explanation of period trend (left to right)	Explanation of group trend (top to bottom)
Atomic radius	Increasing core charge causes valence electrons to be held more tightly to the nucleus, resulting in a smaller atomic radius.	The presence of more shells of electrons makes the atomic radius larger.
Metallic character	Increasing core charge causes electrons to be held more tightly to the nucleus so valence electrons are lost less readily (lower metallic character).	The presence of more shells of electrons causes a shielding effect, so the core charge is felt less. Therefore, electrons are lost more readily (greater metallic character).
Electronegativity	Increasing core charge causes electrons to be held more tightly to the nucleus, resulting in greater attraction between the nucleus and the valence shell of electrons (greater electronegativity).	The presence of more shells of electrons causes a shielding effect so the core charge is felt less, and electronegativity decreases.
First ionisation energy	Increasing core charge causes electrons to be held more tightly to the nucleus, which means more energy is required to remove one valence electron, so first ionisation energy increases.	The presence of more shells of electrons causes a shielding effect, so the core charge is felt less and first ionisation energy decreases.

2.2 CHECK YOUR LEARNING

Describe and explain

1 Explain why there are 10 groups in the d-block of the periodic table.

2 Define ‘metallic’ and ‘non-metallic’ character.

3 Explain why atomic radius decreases going from left to right across a period.

4 State the element in:

a group 1, period 3

d group 15, period 2

c group 17, period 4.

Apply, analyse and compare

5 Explain why sodium and lithium have similar chemical properties.

6 Arrange the following elements in order of increasing ionisation energy: fluorine, magnesium, potassium, silicon.

7 State and compare the subshell electron configurations of K^+, Ar and Cl^-.

8 State and compare the subshell electron configurations of Zn^{2+} and Cu^+.

Design and discuss

9 Discuss any three trends in the periodic table that can be observed going down group 17.

2.3 Critical elements

KEY IDEAS

In this topic, you will learn that:

+ critical elements are elements that are at risk of becoming in very short supply if they are not managed sustainably and recycled where possible
+ helium, phosphorus, neodymium, indium and germanium are critical elements
+ the circular economy of critical elements requires the recycling and recovering elements from waste.

Some elements are at serious risk of becoming depleted within the next 100 years. Many of these elements are crucial for producing the technology that allows us to maintain our high standard of living. We must manage sources of these elements sustainably and replace them with sustainable alternatives where possible. Where replacements for these endangered elements do not exist, we must recover and recycle those elements from waste to ensure a sustainable supply of those elements into the long-term future.

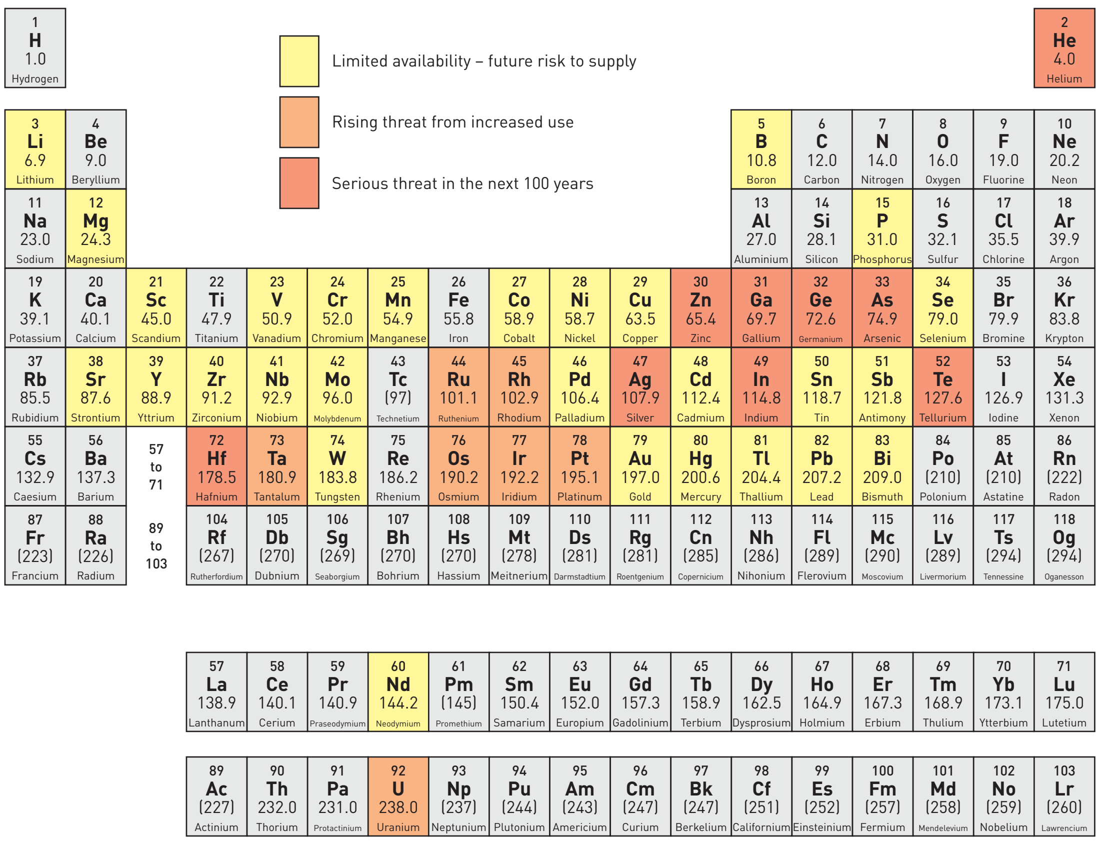

FIGURE 1 A periodic table showing endangered elements and their risk status

Critical elements

critical element
a highly used element at risk of depletion in the near future

Critical elements are those used in industries that contribute to our standard of living and our economy. They are considered 'critical' because their limited supply can significantly affect large industries and economies. The high demand for these elements has placed several critical elements in positions of endangerment and potential extinction.

2.3
PRACTICAL: LITERATURE REVIEW

Are we running out of helium? Go to page **496**.

Examples of critical elements are:

- helium – a group 18 non-metal with atomic number 2
- phosphorus – a group 15 non-metal with atomic number 15
- rare-earth elements – the 15 lanthanide metals as well as yttrium and scandium
- specific post-transition metals – metallic elements that sit between transition metals and metalloids in the periodic table
- specific metalloids – elements with properties that vary between those of a metal and non-metal.

In this topic, we will look at helium and indium (a post-transition metal) as examples of critical elements and examine steps that can be taken to reduce their risk of extinction.

Helium

Helium, despite being the second most abundant element in the Universe, is currently endangered and at risk of extinction. Supplies of helium are finite; we cannot produce it artificially or extract it from the atmosphere. Additionally, replenishment of helium supplies will take millions of years. These factors combined with our high use of helium have placed the element in a critically endangered position.

The use of helium goes well beyond filling balloons and aircrafts. Helium's primary use is to cool substances to extremely low temperatures and create controlled environments. This is needed to run superconductors, MRI (magnetic resonance imaging) scanners in hospitals, NMR (nuclear magnetic resonance) spectrometers in chemistry laboratories and several other machines. Helium is also used to detect pipeline leaks because of its small size. The percentage of helium supplied used across several industries is shown in Figure 2.

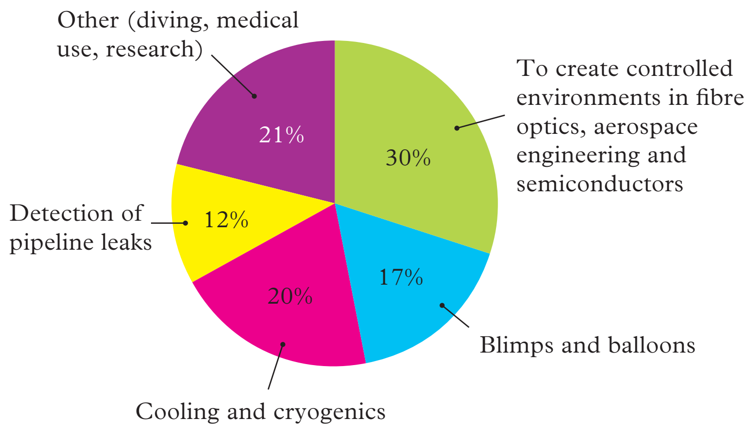

FIGURE 2 The percentage of helium used across different industries

Alternatives to helium

In most industries, finding a substitute for helium is near impossible because of its combination of being inert, having a low boiling point, small atomic size and cooling ability. However, in some instances hydrogen can replace helium. For example, balloons such as weather balloons can be filled with hydrogen gas instead of helium gas. Hydrogen gas is approximately half as dense as helium and therefore provides approximately twice as much lift when it is used in a balloon.

Unfortunately, mixtures of hydrogen gas and air can explode when ignited, which poses a significant safety risk. For this reason, hydrogen balloons cannot replace helium party balloons. However, hydrogen can safely replace a portion of the helium that is used in weather balloons and airships.

Recovering and recycling helium

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

FIGURE 3 Indium is obtained as a by-product from the mining of other metals.

Indium

Indium is a post-transition metal that is used to make LCDs (liquid-crystal displays), LEDs (light-emitting diodes), infrared lasers (for fibre optic cables) and some types of solder and glass. Most touchscreen appliances use indium tin oxide as a transparent conducting material. Indium is rarely mined directly; instead, it is found in very low concentrations as a by-product of mining other metals. Estimates of how much indium remains vary considerably. Some studies suggest that at current usage rates, we will face a shortage of indium in just a few decades.

Alternatives to indium

Alternatives to indium include carbon nanotubing and graphene. However, their properties do not align with all of indium's major uses. A new alternative for indium in touchscreens is being developed by researchers at the University of Sydney. It is a combination of tungsten oxide and silver, coated onto glass. If successful, it could take the pressure off the demand for indium and provide a substitute with higher recycling potential.

Recovering and recycling indium

Recycling indium from used electronic devices is extremely challenging because it is very difficult to separate indium from the other elements and polymers in the product. Also, the amount of indium recoverable from each electronic device is usually too small to make recycling worthwhile at current prices. Institutions can improve indium recycling by promoting take-back initiatives where specific branded items can be more easily recycled in their own stream. People can reduce the demand for indium by upgrading electronics less often, donating devices and recycling old devices correctly.

Other critical elements

Three other critical elements are phosphorus, neodymium and germanium. Table 1 summarises the uses, issues, alternatives and recycling methods for these critical elements.

TABLE 1 A summary of the critical elements phosphorus, neodymium and germanium

Critical element	Primary uses	Issues with supply	Alternatives	Recycling methods
Phosphorus	• As a key ingredient of fertiliser (phosphate is needed for plant growth)	• A finite resource that takes millions of years to mineralise • Approximately 80% of phosphorus is lost or wasted in the supply chain (most being lost to water bodies)	• Although there is no replacement for the phosphorus needed by plants, plant growth can also be encouraged by using organic manure, struvite (found in urine) or fungal species (such as *Trichderma harzianum*)	• Chemical precipitation • Enhanced biological phosphorus removal • Phosphorus recovery at wastewater treatment plants
Neodymium (rare-earth element)	• To make powerful magnets (Nd is a permanent magnet) for electronic devices, electric and hybrid vehicles, wind turbines and aerospace equipment	• Relatively sparse deposits • Increasing demand • Recovery is costly and still developing • Many rare-earth elements are lost during electronic waste recycling	• Tesla's induction motor uses alternatives to the rare-earth elements used in other electric car motors	• Current recycling costs are high and recycling process is underdeveloped • Researchers are developing recycling technologies for rare-earth elements

TABLE 1 continued

Critical element	Primary uses	Issues with supply	Alternatives	Recycling methods
Germanium (metalloid)	• As a semiconductor in electronic devices • In infrared optics • To make alloys	• Scarce natural supply with increasing demand for use • Less than 1% is recycled	• Chalcogenides for infrared optics	• Tannin precipitation • Solvent extraction

Critical elements and the circular economy

In Chapter 1, you learnt that a **circular economy** is one strategy for managing scarce or endangered resources. Instead of disposing of critical elements after their use, the circular economy focuses on recovering these resources.

circular economy
a model of production and consumption that involves sharing, leasing, re-using, repairing, refurbishing and recycling existing materials and products as long as possible

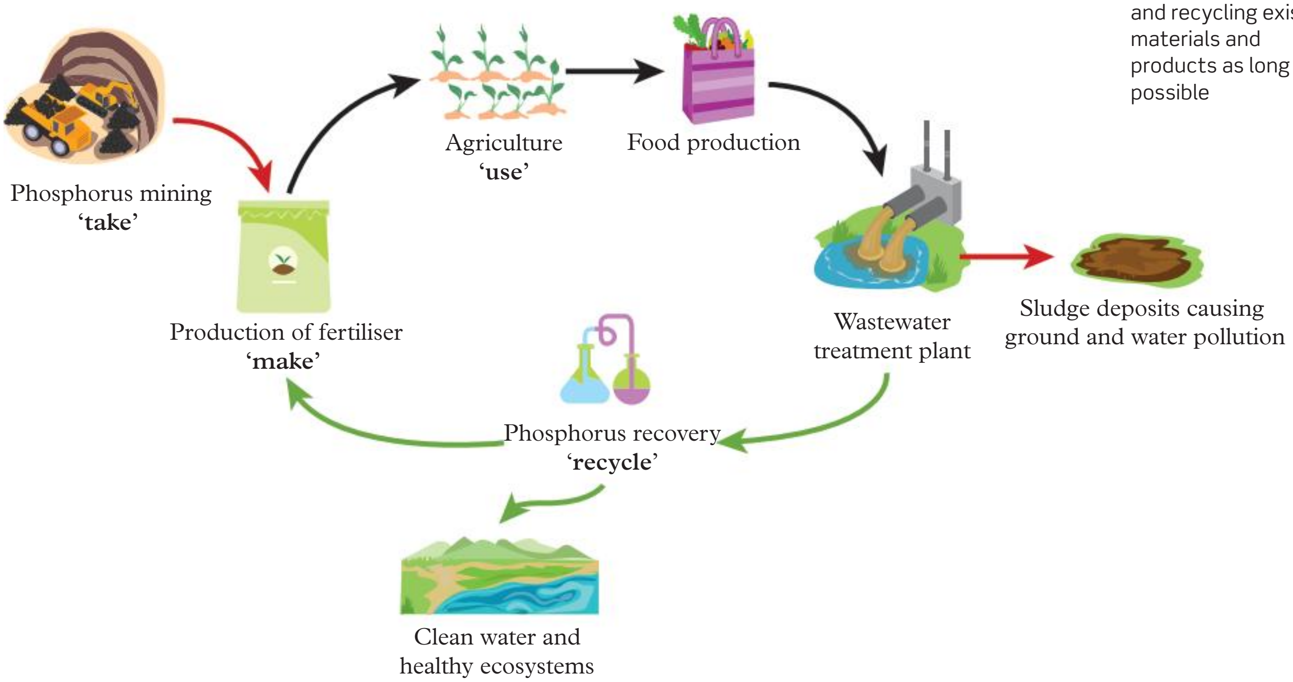

FIGURE 4 An example of a circular phosphorus economy

If no steps are taken to recover critical elements, industries will continue to exploit these resources until they are depleted. It is imperative that industries requiring critical elements shift towards a circular economy to conserve our finite resources. Figure 4 shows how the phosphorus linear economy (where approximately 80% of the critical element is currently lost) can be improved as a circular economy through recovery of phosphorus from waste.

2.3 Skill drill
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2.3 CHECK YOUR LEARNING

Describe and explain

1 List two critical elements and their uses.

2 Describe an issue faced by the indium recycling industry.

Apply, analyse and compare

3 Explain how the circular economy is a potential solution for preserving endangered elements.

Design and discuss

4 Propose an economic incentive that would encourage people to recycle their old electronic devices rather than throw them away.

5 Propose an economic incentive that would encourage manufacturers of electronic devices to recover endangered elements from the devices they collect through their recycling programs.

CHAPTER 2 Review

Chapter summary

2.1

- Each element is made of atoms that have a unique number of protons in their nuclei.
- Atoms also contain neutrally charged neutrons in the nucleus and negatively charged electrons that orbit the nucleus in shells.
- Isotopes are atoms with the same atomic number but different mass numbers as a result of different numbers of neutrons.
- Ions have different numbers of electrons than protons, which gives them an overall positive (cations) or negative (anions) charge.
- Nuclide notation describes the mass number (A), atomic number (Z), elemental symbol and charge (if any) of an atom or ion.

2.2

- The modern periodic table organises 118 elements in increasing order of atomic number (Z).
- Elements in the same period (row) have the same number of electron shells.
- Elements in the same group (column) have the same number of valence electrons and similar chemical properties.
- Electrons orbit the nucleus in shells (holding up to 2, 8, 18 or 32 electrons each), which consist of a number of subshells (holding up to 2, 6, 10 or 14 electrons each), which contain orbitals (which can hold up to 2 electrons each).
- Electrons occupy the lowest energy orbital first before occupying higher energy orbitals according to the Aufbau principle.
- The four trends in the periodic table (atomic radius, metallic character, electronegativity and first ionisation energy) are explained by differences in core charge from left to right, and differences in the number of electron shells and the shielding effect from top to bottom.

2.3

- Critical elements such as helium and indium are at risk of becoming in short supply in the next 100 years.
- To ensure a long-term supply of critical elements, we must embrace a circular economy that includes recovery and recycling of these elements once used, as well as finding more sustainable alternatives where possible.

Key formulas

Number of neutrons in an atom or ion	$N = A - Z$
Number of electrons in an atom or ion	$e^- = Z -$ charge

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• define element, isotope and ion	☐	☐	☐	Go back to Topic 2.1
• use nuclide notation to show atomic number, mass number and number of protons, neutrons and electrons	☐	☐	☐	Go back to Topic 2.1
• apply subshell electronic configurations to elements and ions	☐	☐	☐	Go back to Topic 2.2
• use the periodic table as a tool to identify patterns and properties in elements, including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity	☐	☐	☐	Go back to Topic 2.2
• identify critical elements and explain the importance of recycling processes for element recovery	☐	☐	☐	Go back to Topic 2.3

Revision questions

Multiple choice

1 The ground state electron configuration for Ni^{2+} is:

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 4p^6$

B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8$

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

2 The maximum number of electrons that can occupy a 4p subshell is:

A 2

B 6

C 10

D 14

3 As you move down the elements of group 1 of the periodic table:

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

4 Which of the following represents the nuclide symbol for a cation that has the same number of electrons as the atom argon?

A ${}^{19}_{9}F^-$

B ${}^{23}_{11}Na^+$

C ${}^{35}_{17}Cl^-$

D ${}^{39}_{19}K^+$

5 Which of the following lists the atoms in increasing order of first ionisation energy?

A C, N, O

B O, N, C

C N, O, C

D C, O, N

6 What is the electron configuration of a sodium ion in an excited state?

A $1s^2\ 2s^2\ 2p^5$

B $1s^2\ 2s^2\ 2p^6$

C $1s^2\ 2s^2\ 2p^5\ 3s^1$

D $1s^2\ 2s^2\ 2p^6\ 3s^1$

7 How many orbitals are in a 4s subshell?

A 1

B 2

C 3

D 5

8 Which of these atoms has the most neutrons?

A $^{57}_{27}Co$

B $^{58}_{26}Fe$

C $^{56}_{25}Mn$

D $^{56}_{28}Ni$

9 The ground state electronic configuration for a copper atom is:

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^2$

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

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^1$

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

10 The ground state electronic configuration for a copper(II) ion is:

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3d^9\ 3p^6$

B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^1$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^1$

Short answer

Describe and explain

11 Define:

a electronegativity

b metallic character

c first ionisation energy

12 Determine the core charge of:

a carbon

b sodium

c magnesium

d silicon

e phosphorus.

13 Explain why chromium and copper have unusual electron configurations.

14 Electron configurations can be abbreviated by replacing the core electron configuration with the symbol for the matching noble gas in square brackets. For example, potassium can be written as $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$ or [Ar] $4s^1$. Write the abbreviated electron configurations of the following elements.

a Silicon

b Copper

c Selenium

d Iron

15 Describe the similarities between elements in the same group of the periodic table.

16 Describe the similarities between elements in the same period of the periodic table.

17 Describe one recycling initiative that could recover helium from laboratories.

Apply, analyse and compare

18 Explain why a lithium atom has a smaller atomic radius than a sodium atom.

19 Explain why fluorine has greater electronegativity than nitrogen.

20 Compare the chemical and physical properties of chlorine-35 and chlorine-37.

21 Write the electron subshell configurations for the following atoms and ions.

a Fe

b Fe^{2+}

c Fe^{3+}

d $Fe^{\star}$ (in an excited state)

22 Identify the symbol of the element in:

a period 2, group 1

b period 3, group 2

c period 3, group 17.

23 Identify the number of protons, neutrons and electrons in the following atoms and ions.

a $^{54}_{24}Cr^{3+}$

b $^{51}_{23}V^{3+}$

c $^{30}_{14}Si^{4+}$

d $^{81}_{35}Br^{-}$

24 a Place the following atoms in order from smallest to largest radius: As, Br, K, V.

b Explain your answer to part **a**.

Design and discuss

25 Discuss how the recovery of germanium through recycling could contribute to a circular economy.

26 Draw a Venn diagram that shows the similarities and differences between magnesium-24 and magnesium-26 isotopes.

27 Discuss which trends in the periodic table are caused by the shielding effect and which trends are caused by core charge.

You can find the following resources for this section in your obook pro:

Quizlet

Compete in teams or against yourself to test your knowledge.

Chapter quiz

Test your understanding of Key Knowledge in this chapter.

CHAPTER 3

Covalent substances

KEY KNOWLEDGE

- → the use of Lewis (electron dot) structures, structural formulas and molecular formulas to model the following molecules: hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene
- → shapes of molecules (linear, bent, pyramidal, and tetrahedral, excluding bond angles) as determined by the repulsion of electron pairs according to valence shell electron pair repulsion (VSEPR) theory
- → polar and non-polar character with reference to the shape of the molecule
- → the relative strengths of intramolecular bonding (covalent bonding) and intermolecular forces (dispersion forces, dipole–dipole attraction and hydrogen bonding)
- → physical properties of molecular substances (including melting points and boiling points and non-conduction of electricity) with reference to their structure
- → the structure and bonding of diamond and graphite that explain their properties (including heat conductivity and electrical conductivity and hardness) and their suitability for diverse applications

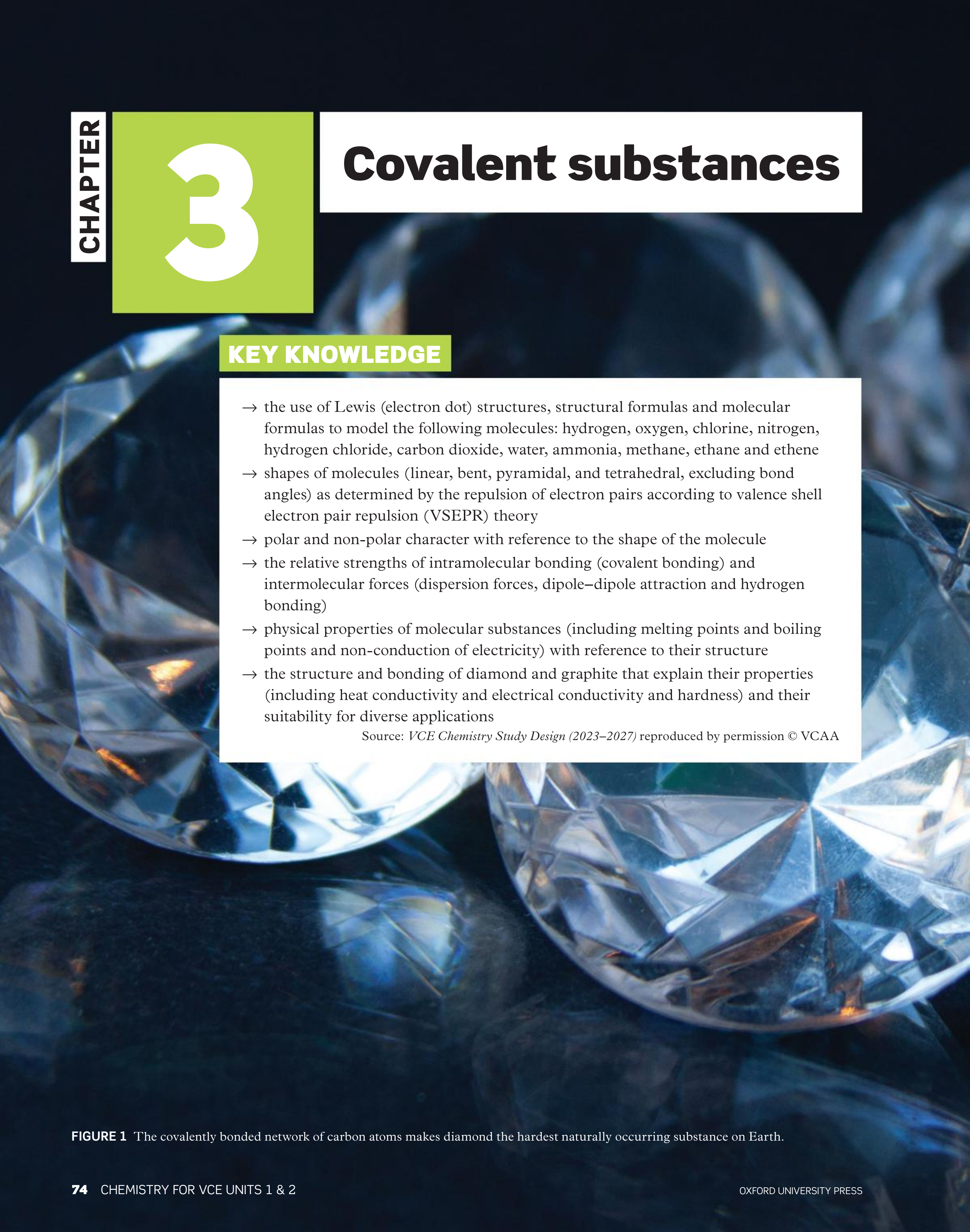


FIGURE 1 The covalently bonded network of carbon atoms makes diamond the hardest naturally occurring substance on Earth.

GROUNDWORK

In Chapter 3, you will learn about how covalent compounds form and how their structure is related to their physical properties.

This chapter will build on concepts you have already learnt in Year 10 Science and Chapter 2. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

3A What type of atoms form covalent bonds and where are these elements located in the periodic table?

3A Groundwork resource
The periodic table

3B How can atoms gain a full valence shell of electrons?

3B Groundwork resource
Valence electrons

3C What is electronegativity?

3C Groundwork resource
Electronegativity

3D What is a physical property of a chemical?

3D Groundwork resource
Physical properties

PRACTICALS

3.2A PRACTICAL: MODELLING	How can modelling be used to understand molecular shapes?	Page 497
3.2B PRACTICAL: MODELLING	How can balloons help us understand VSEPR theory?	pro
3.5 PRACTICAL: CONTROLLED EXPERIMENT	How do intermolecular forces influence the physical properties of covalent molecules?	Page 498

3.1 Covalent compounds

KEY IDEAS

In this topic, you will learn that:

+ a non-metal atom forms a covalent bond when it shares electrons with an atom of the same element or a different element
+ covalent bonds can be single, double or triple
+ there are different ways to represent the molecular formula.

Atoms are the building blocks of all materials, but how the atoms are bonded together is also extremely important. Covalent compounds are formed when non-metal atoms share electrons.

The non-metals make up a very small number of elements in the periodic table (Figure 1). However, when non-metals bond together, they can form a large range of different **molecules** and **compounds**.

molecule
a group of atoms covalently bonded together

compound
two or more atoms bonded together, where the atoms belong to two or more different elements, e.g. O_2 is a molecule, but not a compound; HCl is both a molecule and a compound

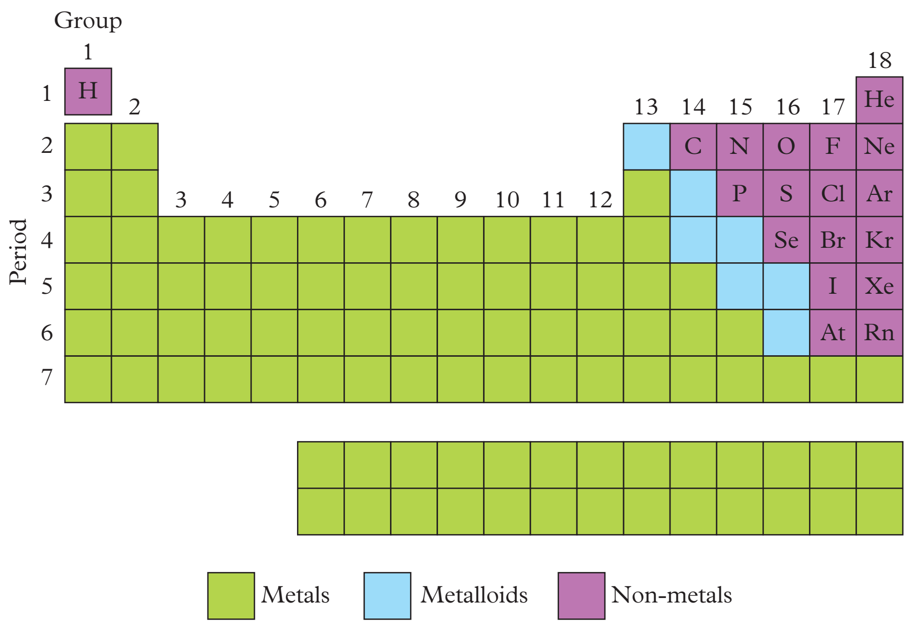


FIGURE 1 The location of non-metals in the periodic table

In this chapter, you will look at the bonds that form between non-metal atoms and the physical properties of the molecules that form from this bonding.

valence shell
the outer shell of an atom where electrons are found

Types of covalent bonding

The noble gases have full **valence shells** and are stable. Other atoms, such as the non-metals, do not have a full outer shell and are therefore not as stable. However, they can complete their outer shell by forming bonds.

octet rule
a rule that states atoms gain or lose electrons to have eight electrons in their valence shell (excluding hydrogen and helium)

Atoms can form bonds with other atoms by donating, accepting or sharing electrons (Figure 2). For most atoms, eight electrons are required for a complete outer shell. This is called the **octet rule**.

Covalent bonds form when two or more non-metal atoms share electrons to complete their octets. By sharing electrons, they become more stable than the separate atoms.

For example, oxygen has six valence electrons in its outer shell (Figure 3). It needs two more to have a full outer shell. Therefore, it forms covalent bonds with other atoms to complete the octet.

In Figure 3, you can see that there are two shared pairs of electrons. These are called **bonding electrons**. Oxygen also has two sets of **non-bonding electrons**, or **lone pairs**, which do not participate in covalent bonding.

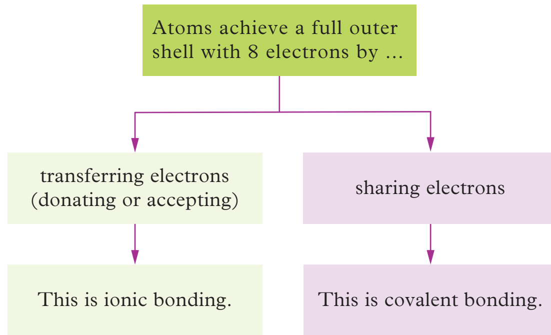


FIGURE 2 To achieve a full (and therefore stable) outer shell, atoms gain, lose or share electrons.

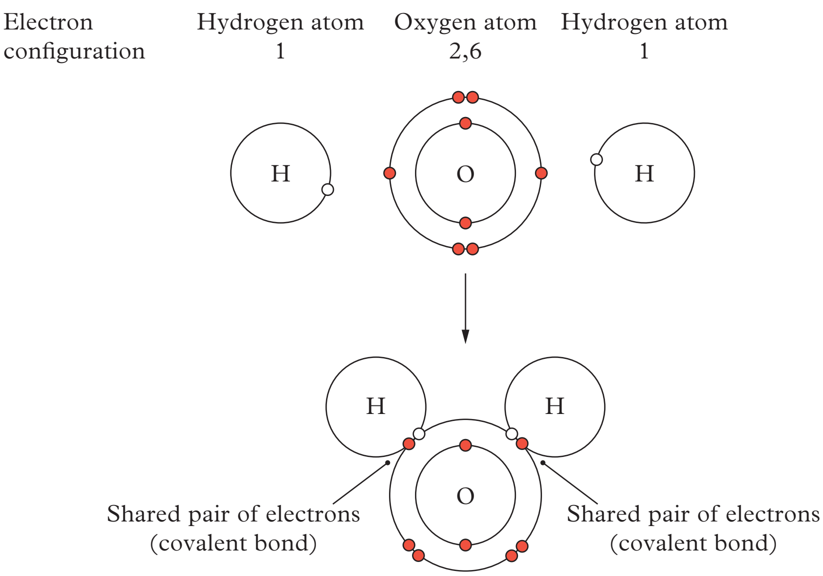


FIGURE 3 Oxygen shares electrons with two hydrogen atoms in order to have eight electrons in its outer valence shell.

covalent bond
a bond that forms between two non-metal atoms when they share one or more pairs of electrons

bonding electrons
valence electrons that are shared with another atom in a covalent bond

non-bonding electrons
valence electrons that are not shared with another atom in a covalent bond

lone pair
a pair of valence electrons that are not shared with another atom in a covalent bond

single covalent bond
a bond that forms between two non-metal atoms when they share two electrons (one pair)

electron pair
a pair of electrons that occupy the same shell and can be part of a bond with another atom or non-bonding as a lone pair

Single covalent bonds

A **single covalent bond** is formed when two electrons (an **electron pair**) are shared between two atoms.

In Figure 4, you can see that the hydrogen atoms each have one electron in their outer shell. When a hydrogen atom shares its single electron with another hydrogen atom, both hydrogen atoms now have two electrons in their outer shell. This sharing of single electrons forms a single covalent bond, making the molecule stable.

Hydrogen is an exception to the octet rule because its valence shell can only have a maximum of two electrons. All other non-metals you will study in VCE Chemistry follow the octet rule for forming covalent bonds.

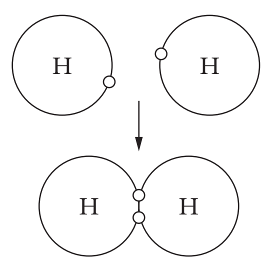


FIGURE 4 Two hydrogen atoms share their single outer shell electrons to form a covalent bond.

Many of the halogens (group 17) form molecules in the same way. For example, chlorine has the electron configuration 2,8,7. This means that each chlorine atom needs one more electron to complete its outer shell and become stable. Chlorine atoms can share electrons with other chlorine atoms (Figure 5) to form molecules of chlorine gas (Cl_2). Both atoms now have full valence shells of eight electrons and the molecule is more stable than the chlorine atoms on their own

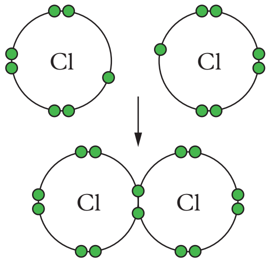


FIGURE 5 Chlorine atoms share an electron pair to form a single covalent bond. Note that only the valence electron shell is shown here.

Double covalent bonds

double covalent bond
a bond that forms between two non-metal atoms when they share four electrons (two pairs)

A **double covalent bond** forms when two atoms share four electrons (two pairs).

Some atoms, such as oxygen, have six valence electrons, meaning that they need two more electrons to complete a full outer shell with eight electrons. They can do this by:

- sharing each single electron separately, i.e. forming two single bonds with two different atoms
- sharing two electrons with another atom with the same number of valence electrons to form a double covalent bond.

In Figure 6, you can see how two oxygen atoms share two pairs of electrons to form a double covalent bond. When two oxygen atoms form a double covalent bond, the molecular formula is O_2, and each oxygen has two sets of non-bonding electrons in their outer shells.

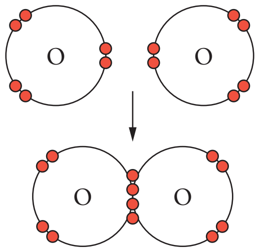


FIGURE 6 Two oxygen atoms sharing two pairs of electrons to form a double covalent bond. Note that only the valence electron shell is shown here.

Triple covalent bonds

triple covalent bond
a bond that forms between two non-metals when they share six electrons (three pairs)

A **triple covalent bond** forms when two atoms share six electrons (three pairs).

A nitrogen molecule contains a triple covalent bond (Figure 7). A nitrogen atom has five electrons in its outer shell so it requires three more electrons to achieve a full valence shell of eight electrons. Each nitrogen atom in the molecule also has a lone pair of electrons in its valence shell. The triple covalent bond formed between the two nitrogen atoms is very strong and makes nitrogen relatively unreactive.

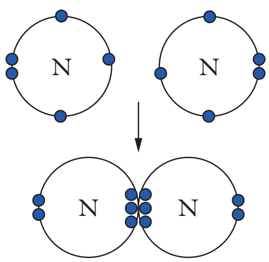


FIGURE 7 Two nitrogen atoms sharing three pairs of electrons to form a triple covalent bond

TABLE 1 Electron pairs are shared to form single, double and triple covalent bonds

Number of electron pairs shared	Type of covalent bond
1	Single
2	Double
3	Triple

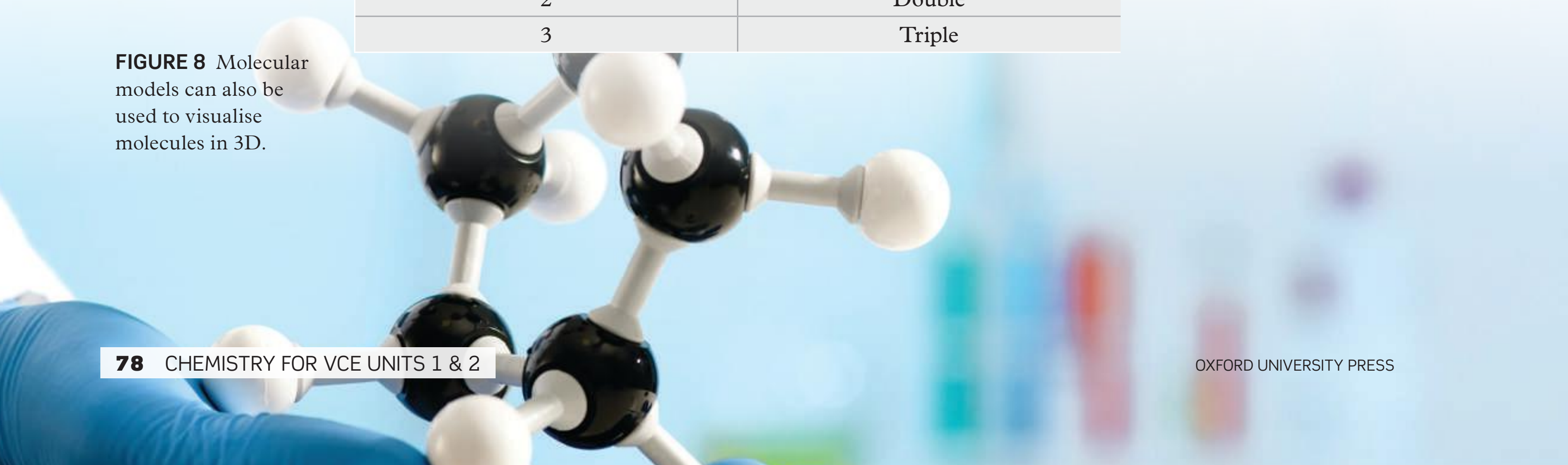


FIGURE 8 Molecular models can also be used to visualise molecules in 3D.

Molecular compounds

All the covalent bonds that you have seen so far have been between two atoms of the same element. However, covalent bonds can also form between atoms of different elements.

Diatomic molecules

When two atoms of the same or different elements join, they form a **diatomic molecule**. H_2, Cl_2, O_2 and N_2 are all diatomic molecules made up of atoms of the same elements. Hydrogen chloride (HCl) is also a diatomic molecule, but it is made up of a hydrogen atom and a chlorine atom (Figure 9). This makes HCl a compound.

Chlorine has seven electrons in its outer shell and hydrogen has one. Because each atom only needs one more electron to complete its outer shell, they can share a pair of electrons between them to form a covalent bond. This forms hydrogen chloride.

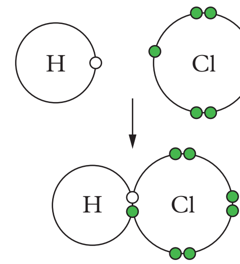


FIGURE 9 A hydrogen chloride (HCl) molecule has a covalent bond between a chlorine atom and a hydrogen atom.

Polyatomic molecules

When more than two atoms of the same or different elements join, they form a polyatomic molecule. Water is a **polyatomic molecule**. It is made up of two hydrogen atoms covalently bonded to one oxygen atom.

Hydrogen atoms each have a single electron in their outer shell and need two electrons to become stable. An oxygen atom has six valence electrons and needs eight to be stable. To form a stable water molecule, the oxygen shares one electron with each hydrogen atom and forms two single covalent bonds.

Polyatomic molecules can also contain double bonds. For example, carbon dioxide is made up of one carbon atom with double covalent bonds to two oxygen atoms (Figure 10).

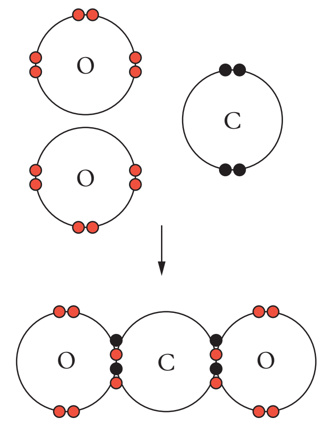


FIGURE 10 Carbon dioxide is a polyatomic molecule with two double covalent bonds.

diatomic molecule
a molecule made of two atoms, which can be of the same element or different elements

polyatomic molecule
a molecule made up of three or more atoms

molecular formula
a way of representing the structure of a covalent molecule that shows the number and type of atoms in the molecule

Representing molecules

As molecules become more complex, it is not reasonable to draw all the electrons and valence shells of every atom. There are different and simpler ways to represent the structures of covalent molecules. The four ways you will look at in this chapter are:

- molecular formula
- structural formula
- Lewis (electron dot) structure
- valence structure.

Molecular formula

The **molecular formula** gives information about the number and type of atoms that make up the molecules. Figure 11 shows the molecular formula of methane.

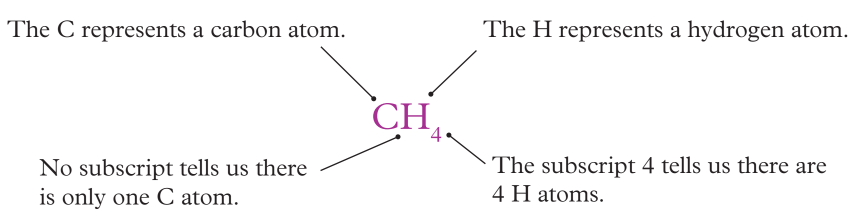


FIGURE 11 The molecular formula of methane (CH_4)

Structural formula

structural formula
a way of representing the structure of a covalent molecule that shows the covalent bonds between the atoms as lines

The **structural formula** shows all the atoms in the molecule and their covalent bonds (represented by a line). It gives information about the number and types of atoms, and what types of covalent bonds are present. Figure 12 shows structural formulas of hydrogen, ammonia and ethene.

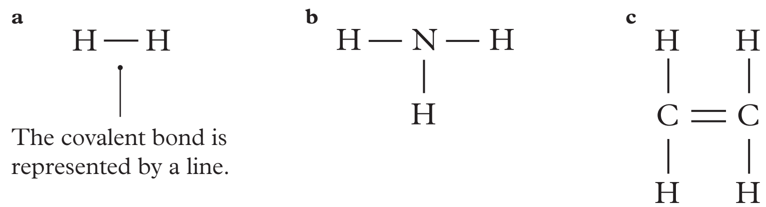


FIGURE 12 The structural formula of **a** hydrogen gas (H_2) shows that it has one single covalent bond; **b** ammonia (NH_3) shows that there are three single bonds; and **c** ethene (C_2H_4) shows there is a double bond present between the two carbon atoms.

Lewis (electron dot) structure

Lewis structure
a way of representing the structure of a covalent molecule that shows all bonding and non-bonding valence electrons as dots or crosses; also called an electron dot structure

Lewis structures (also called electron dot structures) show all the valence electrons in each atom of a molecule. Electrons are represented by crosses and dots. Figure 13 shows only the bonding electrons in hydrogen chloride.

The Lewis structure also shows non-bonding electrons (Figure 14).

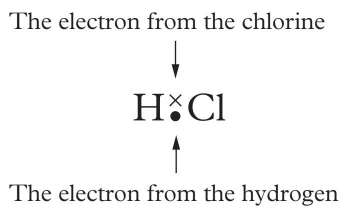


FIGURE 13 The bonding electrons in hydrogen chloride (HCl) are represented by a cross and a dot to show which atoms the electrons come from.

Study tip

Lewis structures do not always include both crosses and dots. Sometimes, they only use dots. Worked example 3.1 shows different ways to draw carbon dioxide.

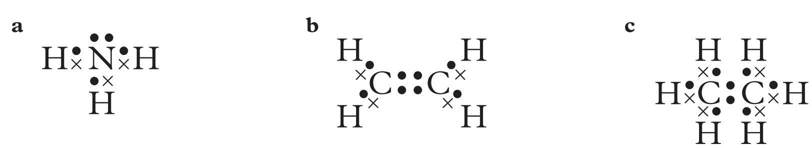


FIGURE 14 The Lewis structures of **a** ammonia (NH_3), **b** ethene (C_2H_4) and **c** ethane (C_2H_6). In the Lewis structure of ammonia, you can see the non-bonding (lone pair) electrons on the nitrogen atom.

Valence structure

valence structure
a way of representing the structure of a covalent molecule that shows the covalent bonds between all atoms as lines, and lone pair electrons as dots or crosses

Another common way to represent molecules is the **valence structure**. This is a simplified Lewis structure with the bonding electrons shown as lines (Figure 15).

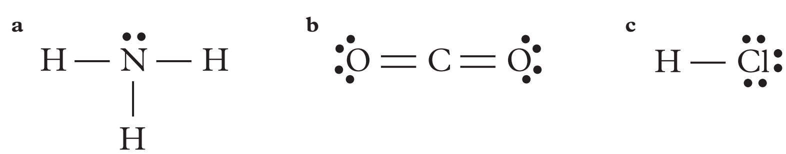


FIGURE 15 The valence structures of **a** ammonia (NH_3), **b** carbon dioxide (CO_2) and **c** hydrogen chloride (HCl). The bonding electrons are shown as lines.

Advantages and disadvantages of the representations

Each way of representing a molecule has its advantages and disadvantages. One key disadvantage of the four representations is that they do not show the size of the atoms in a molecule or the overall molecular shape. The features and limitations of molecular formulas, structural formulas, Lewis structures and valence structures are shown in Table 2.

See how to draw the different representations in Worked example 3.1.

3.1 Worked example
Find me in your obook pro

3.1 Worked example
Video demonstration

TABLE 2 The features and limitations of the different molecular representations

	Information provided			Limitations
	Number and type of atoms	**Type of covalent bonds**	**Non-bonding electrons**	
Molecular formula	✓			Does not show: • how atoms are bonded together • size of atoms or shape of molecule • non-bonding electrons.
Structural formula	✓	✓		Does not show: • size of atoms or shape of molecule • non-bonding electrons.
Lewis structure	✓	✓	✓	Does not show: • size of atoms or shape of molecule • type of covalent bonds as clearly as structural formula. Becomes more cluttered the more atoms there are in a molecule.
Valence structure	✓	✓	✓	Does not show: • size of atoms or shape of molecule.

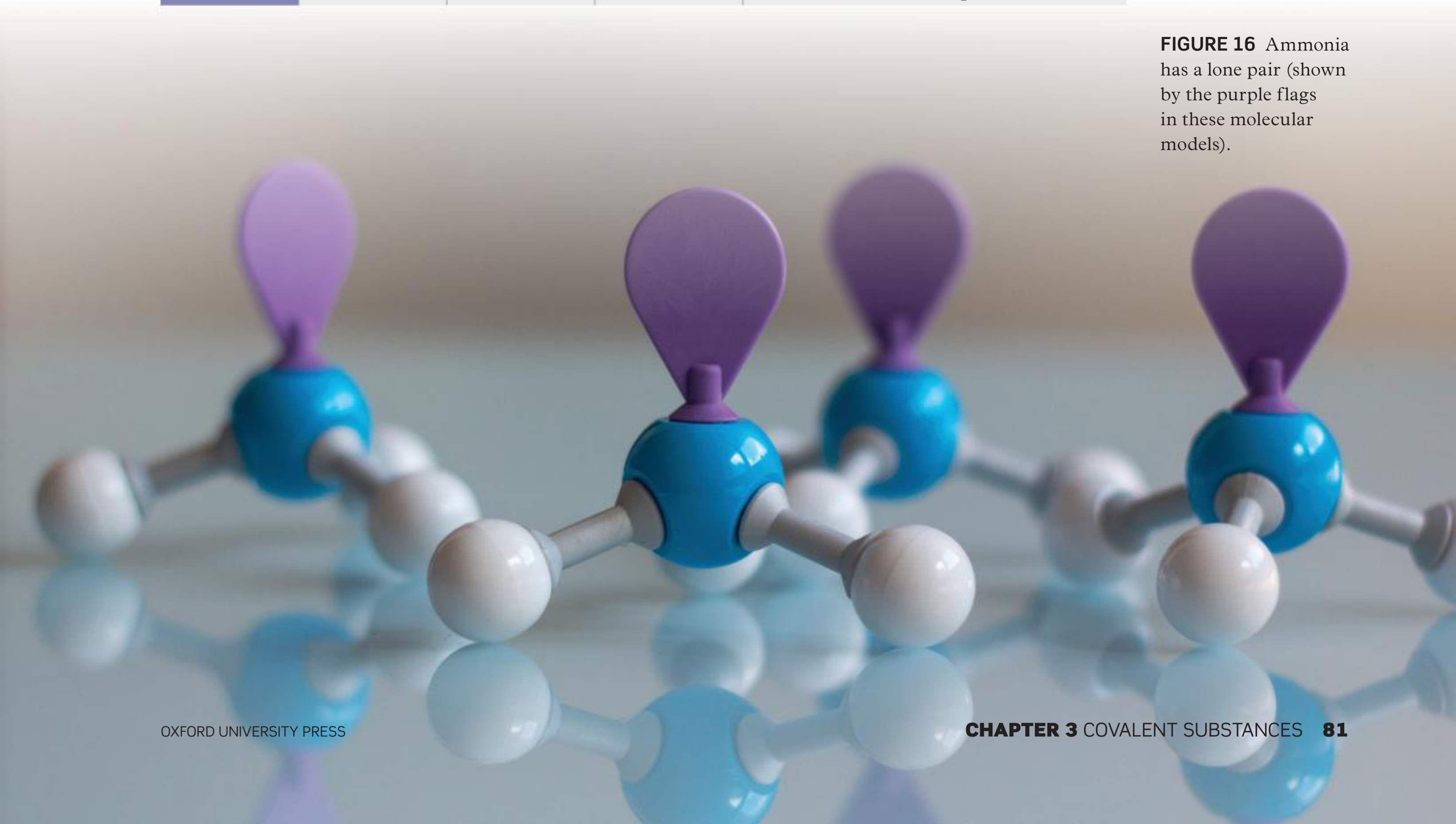


FIGURE 16 Ammonia has a lone pair (shown by the purple flags in these molecular models).

3.1 CHECK YOUR LEARNING

Describe and explain

1 Explain what a covalent bond is.

2 Describe the octet rule.

3 Explain the difference between a diatomic and polyatomic molecule.

4 Identify how many covalent bonds there are in each of these diatomic molecules.

a H_2 b O_2

c Cl_2 d N_2

Apply, analyse and compare

5 Draw the structural formulas of the following molecules.

a H_2O b CCl_4

c C_2H_6 d HF

6 Draw the Lewis structures of the following molecules.

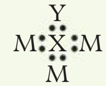
a Cl_2 b CH_4

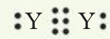
c SO_2 d HCl

7 Write the molecular formulas of the following Lewis structures.

a 

b X:X

c 

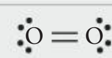
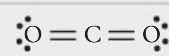
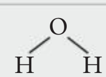
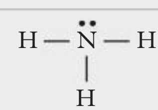
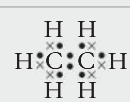
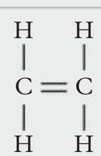
d 

8 Copy and complete the table below by filling in the missing molecular formulas, structural formulas, Lewis structures and valence structures.

9 Compare the Lewis and valence structures for H_2, CH_4, C_2H_6 and C_2H_4. Are there any differences? Why or why not?

Design and discuss

10 Discuss the advantages and disadvantages of the four types of structural representation.

Molecular formula	Structural formula	Lewis structure	Valence structure
	H — H		
			
Cl_2			
			
			
			
			
			
CH_4			
			
			

3.2 Shapes of molecules

KEY IDEAS

In this topic, you will learn that:

- valence shell electron pair repulsion (VSEPR) theory is used to predict molecular shape
- the shape of a covalently bonded molecule depends on its bonding and non-bonding electron pairs
- there are four different molecular shapes: linear, bent, pyramidal and tetrahedral.

molecular shape
the arrangement of atoms in an individual covalently bonded molecule

Molecular shape plays an important role in how molecules react and interact with other molecules. Many of the physical and chemical properties of covalent molecules result from their shape.

Molecular shape

valence shell electron pair repulsion (VSEPR) theory
a theory that states that the valence electron pairs (either bonding or non-bonding) arrange themselves to be as far apart as possible

Molecular shape refers to the arrangement of atoms in an individual, covalently bonded molecule. We can work out the molecular shape of a molecule by using **valence shell electron pair repulsion (VSEPR) theory**.

Valence shell electron pair repulsion theory

VSEPR theory focuses on the geometry of electrons. It states that:

- valence electrons are grouped in pairs (including bonding electrons and non-bonding electrons)
- electron pairs are arranged around the central atom(s) of the molecule
- all electron pairs in an atom repel each other
- to minimise the negative charges between them, the electron pairs are as far apart as possible.

electron-dense area
a region of a molecule where there are one or more pairs of bonding or non-bonding electrons; for example, a lone pair, or electrons in a single, double or triple covalent bond

These electron pairs create **electron-dense areas**, which affect the shape of a molecule. Let's look at how VSPER theory applies to methane (Figure 1).

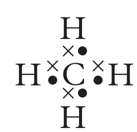


FIGURE 1 The Lewis structure of methane (CH_4).

We know that in methane:

- a carbon atom is surrounded by four hydrogen atoms
- the stable carbon atom has eight valence electrons (four valence electron pairs)
- each hydrogen atom is stable with its two valence electrons.

bond angle
the angle between adjacent bonds from the same atom

ball-and-stick model
a way of visualising the 3D structure of a molecule using balls to represent atoms and sticks to represent bonds

The shape of the methane molecule is determined by its four bonding pairs of electrons. Because of their negative charges, the electron pairs try to get as far away from each other as possible. They do this by arranging themselves around the central carbon atom at equal distances apart from each other.

Because molecules are three-dimensional structures, the two-dimensional Lewis structure is only helpful up to a point. To see the exact **bond angle** and distances between atoms, you can use the **ball-and-stick model** (Figure 2). This shows that the methane molecule is tetrahedral.

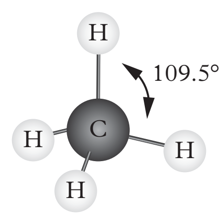


FIGURE 2 The ball-and-stick model of methane (CH_4). The hydrogens are arranged so that the bonding angles are 109.5°.

3.2A PRACTICAL: MODELLING

How can modelling be used to understand molecular shapes? Go to page **497**.

3.2B PRACTICAL: MODELLING

How can balloons help us understand VSEPR theory? Go to your **Student obook pro**.

Lone pairs

Not all pairs of electrons are part of a covalent bond. Some are non-bonding pairs, or lone pairs, of electrons. You can see this in hydrogen chloride (HCl) (Figure 3).

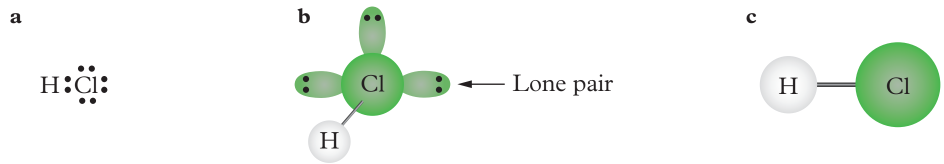


FIGURE 3 **a** The Lewis structure of hydrogen chloride (HCl) with three lone pairs on the chlorine, and **b** the molecular shape that this creates, shown using the ball-and-stick model. **c** The molecule has a linear shape.

Study tip

We talk about the electrons as pairs because each pair of electrons is one single covalent bond.

The three non-bonding pairs of electrons and one bonding pair (between the hydrogen and chlorine atoms) arrange themselves at equal angles around the central chlorine atom to form a tetrahedral shape. This means there are four electron-dense areas. However, because the lone pairs are not bonds, the hydrogen chloride molecule has a linear shape. This is one of the four different molecular shapes we will look at in this chapter, which include:

- tetrahedral
- pyramidal
- bent or V-shaped
- linear.

Study tip

You do not need to remember the bond angles for your exams but is useful to remember the shapes and that the electron pairs are as far apart as possible.

Tetrahedral molecules

A **tetrahedral** molecule has a triangular pyramid shape. Many of the molecules we have looked at so far have electrons arranged in a tetrahedral shape. However, a molecule's shape can only be called tetrahedral when it has four electron-dense areas around a central atom that are also bonding pairs of electrons.

Methane is a tetrahedral molecule. As you saw in Figure 2, the central carbon has no lone pairs. This means that all the electrons can be as far from the other electron pairs as possible and form this tetrahedral shape.

tetrahedral
arranged in a triangular pyramid shape; the central atom is bonded to four atoms

Pyramidal molecules

A **pyramidal** molecule has a triangular shape. It is basically the tetrahedral shape without the top atom. An example of this is ammonia (Figure 4).

pyramidal
arranged in a triangular shape; the central atom is bonded to three atoms with one lone pair

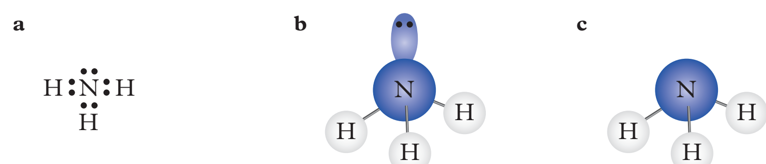


FIGURE 4 Ammonia (NH_3) is a pyramidal molecule, as shown by its **a** Lewis structure and **b** ball-and-stick model. **c** Overall, the molecule has a pyramidal shape.

The nitrogen in ammonia has a full shell of eight valence electrons. Three electron pairs are bonded to hydrogen atoms and one pair is non-bonding. The shape they form around the nitrogen is a tetrahedral shape because of the four electron-dense areas. However, when viewed without the lone pair of electrons, ammonia has a pyramidal shape.

Bent molecules

Water is the best example of a **bent**, or V-shaped, molecule. It is this shape that gives water many of its properties. Water also has four electron-dense areas around the central oxygen atom: two lone pairs and two bonding pairs of electrons. You can see how they arrange

bent
arranged in a V-shape; the central atom is bonded to two atoms

themselves around the central oxygen atom in Figure 5. Because the lone pairs are ignored when discussing the molecule's shape, water is said to have a bent or V-shape.

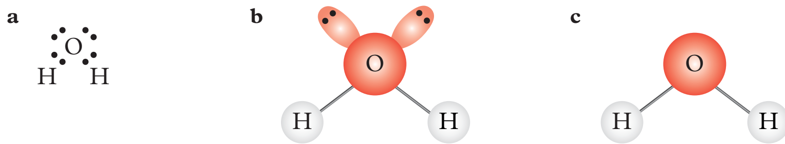


FIGURE 5 The V-shape of a water molecule, as shown by its **a** Lewis structure and **b** ball-and-stick model. **c** Overall, the molecule has a bent shape.

The flat V-shape is formed when the two lone pairs push the hydrogen–oxygen bonds closer, reducing the bond angle.

Linear molecules

The atoms in a **linear** molecule, such as hydrogen chloride, are arranged in a straight line. All molecules made up of two atoms are linear. Some molecules containing three or four atoms are also linear, depending on the number of electron-dense areas. For example, carbon dioxide (CO_2) contains three atoms and is linear in shape (Figure 6).

linear
arranged in a straight line; the central atom is bonded to one or two atoms

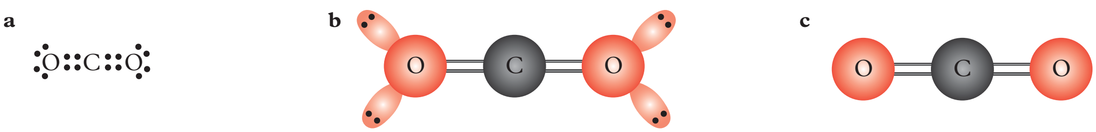


FIGURE 6 Carbon dioxide (CO_2) has a linear shape, as shown by its **a** Lewis structure and **b** ball-and-stick model. **c** Overall, the molecule has a linear shape.

The central carbon in the carbon dioxide molecule has four bonding pairs of electrons. However, because they are double covalent bonds, the two pairs of electrons form two electron-dense areas. They stay together but repel all other paired electrons. This makes carbon dioxide linear in shape.

Many gases are small linear molecules. In Figure 7, you can see the linear shape of O_2 and N_2 gases. Each oxygen atom in the O_2 molecule has four pairs of electrons: two bonding and two lone pairs. As in carbon dioxide, the electrons in the double bond stay together and the two bonding pairs on each oxygen atom arrange themselves as far away as possible from the other valence electrons. This is the same for N_2. The triple-bonded electrons stay together and the lone pair on each nitrogen atom is positioned as far as possible from the triple bond.

A summary of the covalent molecular shapes is shown in Table 1.

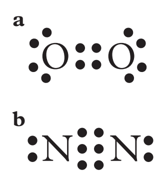
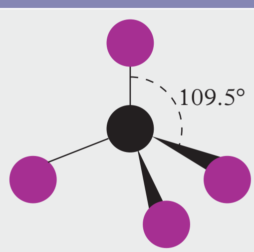
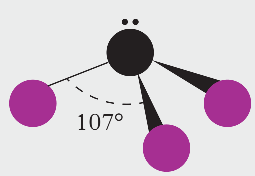
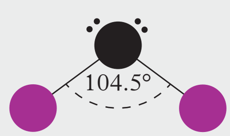
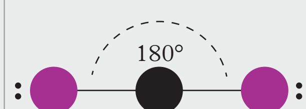


FIGURE 7 Gaseous **a** oxygen (O_2) and **b** nitrogen (N_2) have linear shapes.

TABLE 1 Molecular shapes of covalent molecules

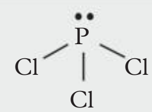
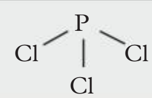
Name	Tetrahedral	Pyramidal	Bent	Linear
Molecular shape	 109.5°	 107°	 104.5°	 180°
Number of electron-dense areas	4	4	4	2
Number of lone pairs	0	1	2	2
Example	CH_4	NH_3	H_2O	CO_2

3.2 WORKED EXAMPLE

IDENTIFYING MOLECULAR SHAPES

Determine the molecular shape of PCl_3.

Solution

Think	Do
Step 1: Use the periodic table to determine how many valence electrons each atom has and how many more are needed to make a full valence shell.	P has five electrons in its outer shell, so it needs three electrons to complete its valence shell. Cl has seven electrons in its outer shell, so each chlorine needs one electron to complete its valence shell.
Step 2: Identify the type of covalent bonds needed between the atoms to make full valence shells.	Cl can only form a single covalent bond with P because it can only share one electron.
Step 3: Draw the Lewis structure of PCl_3.	
Step 4: Count the number of bonds and lone pairs from the central P atom and determine the overall shape of the molecule with all electron pairs.	There are three bonds and one lone pair on the central P atom. The electrons are arranged in a tetrahedral shape for maximum separation. 
Step 5: Determine the shape of the molecule based on maximum spacing of all the electron pairs and without any lone pairs.	The shape of PCl_3 is pyramidal. 

3.2 CHECK YOUR LEARNING

Describe and explain

1 Describe the following molecular shapes.

a Linear b Bent

c Pyramidal d Tetrahedral

2 Explain VSEPR theory and how it is used to determine molecular shapes.

3 Describe the role of non-bonding electron pairs in a molecule's shape.

Apply, analyse and compare

4 Determine the molecular shapes of the following molecules.

a F_2 b PH_3

c CO_2 d HI

e CF_4

5 Draw the Lewis structures of the following molecules.

a CCl_4 b Cl_2

c NH_3 d H_2S

e CS_2

6 Determine and draw the molecular shapes for the molecules in Question 5.

Design and discuss

7 O_2 and N_2 have a double and a triple bond, respectively. Both molecules are linear.

'A molecule with double or triple bonds can have any shape other than linear.' Evaluate this statement, using examples to explain your reasoning.

3.3 Polar and non-polar characteristics

KEY IDEAS

In this topic, you will learn that:

+ the polarity of a covalent bond depends on how atoms in the bond share electrons
+ bonds are polar if the atoms have different electronegativities
+ the shape of a molecule, its non-bonding electrons, and the direction of its dipoles affect its polarity.

polarity
the separation of electric charge in a bond or molecule that depends on the electronegativity of its atoms

Electron pairs in covalent bonds are not always shared equally. This is because the different non-metals in a molecule have different electronegativities.

Understanding electronegativity is important because it controls polarity. **Polarity** can be used to explain the properties of covalent compounds.

Electronegativity

Electronegativity is a measure of a non-metal atom's ability to attract valence electrons.

In Chapter 2, you looked at the trend of electronegativity in the periodic table. You might remember that the electronegativity of the elements increases as you go from left to right (excluding the noble gases) and decreases as you move down the periodic table (Figure 1). Therefore, the most electronegative elements are the non-metals in the top right corner of the periodic table.

Electronegativity increases →

Electronegativity decreases ↓

H 2.2																	He
Li 1.0	Be 1.6											B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0	Kr 2.9
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	*	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 2.0	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

Lanthanoids *	La 1.1	Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.1
Actinoids **	Ac 1.1	Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3

FIGURE 1 Electronegativity increases from left to right and decreases going down the periodic table. The non-metals are outlined in purple.

All the non-metals have higher electronegativities than the other elements (metals and metalloids) (see Chapter 2). The more electronegative the atom, the more it attracts electrons to its nucleus. Therefore, it has a bigger share of the electrons in the covalent bond.

This means that two covalently bonded non-metals with different electronegativities share the electron pair with different strengths of attraction.

In Figure 2, you can see that in Cl_2 gas the two chlorine atoms share the bonding electrons equally because of their equal electronegativity. This makes it a **non-polar covalent bond**. However, in hydrogen chloride (HCl), the chlorine atom has a much higher electronegativity than hydrogen, so it attracts the shared electron pair towards its nucleus more strongly. This gives the HCl molecule an electron-rich region and an electron-poor region. The bond between the hydrogen and chlorine is therefore a **polar covalent bond**.

non-polar covalent bond
a covalent bond in which electrons are shared equally; the two bonded atoms have equal electronegativity

polar covalent bond
a covalent bond in which electrons are not shared equally; the two bonded atoms have different electronegativities

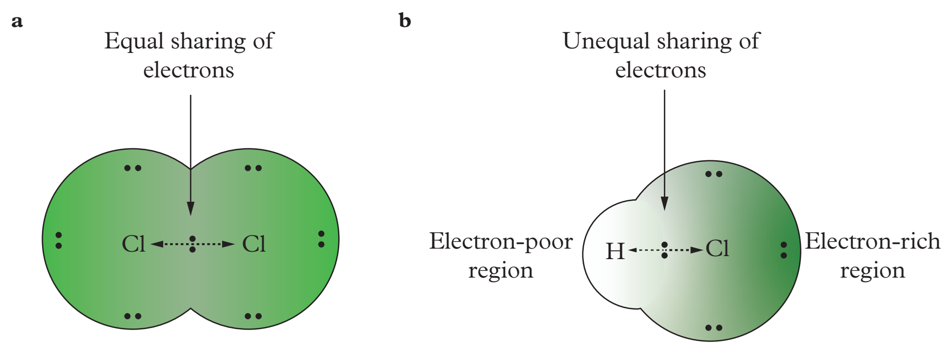


FIGURE 2 Sharing of electron pairs in **a** chlorine (Cl_2) and **b** hydrogen chloride (HCl).

Polarity

Polarity describes how electric charge is separated in a molecule. Polarity depends on the molecular shape and the polarity of the bonds. Generally:

- **non-polar** molecules are **symmetrical** with equal sharing of electrons between atoms
- **polar** molecules are **asymmetrical**, having either lone pairs of electrons on a central atom or atoms with different electronegativities.

non-polar
not having a positive and a negative end

symmetrical
made up two halves that are mirror images of each other

polar
having a positive end and a negative end

asymmetrical
opposite of symmetrical, made up of parts that are not mirror images of each other

overall dipole
the presence of positive and negative charges at opposite ends of a molecule; unbalanced sharing of electrons

Non-polar molecules

Non-polar molecules have an equal distribution of charge on all ends of the molecule. This means that they have no **overall dipole** (unbalanced sharing of electrons).

If atoms in a molecule have the same electronegativity:

- the atoms have the same 'pulling power'
- electrons are shared equally
- electrons are located (on average) halfway between atoms
- charge across the molecule is even (non-polar).

Diatomic non-polar molecules

In diatomic molecules made up of identical atoms – such as gaseous hydrogen (H_2), oxygen (O_2) and nitrogen (N_2) – each of the atoms has the same electronegativity. In these molecules, the atoms share the electron pair(s) equally. Therefore, they are non-polar molecules.

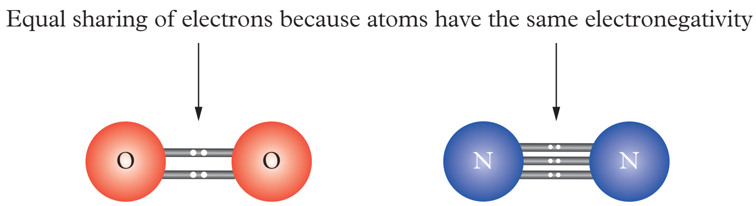


FIGURE 3 In non-polar diatomic molecules, such as oxygen (O_2) and nitrogen (N_2), the atoms share the electron pair equally.

dipole
an uneven separation of charge; formed when a bond or a molecule has a partial positive charge (δ+) at one end and a partial negative charge (δ–) at the opposite end

Polyatomic molecules

Polyatomic molecules that are symmetrical are also non-polar. This is because no one atom in the molecule has an unequal 'pull' on the electron pair.

For example, methane is a tetrahedral molecule that has a central carbon surrounded equally by four hydrogen atoms (Figure 4). Carbon is slightly more electronegative than hydrogen. This means that the electron pairs are more strongly attracted to the carbon than to the hydrogen atoms, giving carbon a more negative charge.

This slight difference in charge creates a **dipole** (or partial charge) and is represented by the Greek letter delta (δ). The carbon has a partial negative charge (δ–) and the hydrogen atoms each have a more positive charge (δ+). This can also be shown using a dipole arrow, which is an arrow with a plus symbol at one end.

Because the methane molecule is symmetrical, the slight dipoles cancel each other out, making the overall methane molecule non-polar. Another example is carbon dioxide (CO_2). The bonding electrons are pulled towards the more electronegative oxygen atoms. This gives the oxygen atoms partial negative charges while the central carbon has a partial positive charge. Again, because the molecule is symmetrical, the dipoles cancel out to make carbon dioxide non-polar.

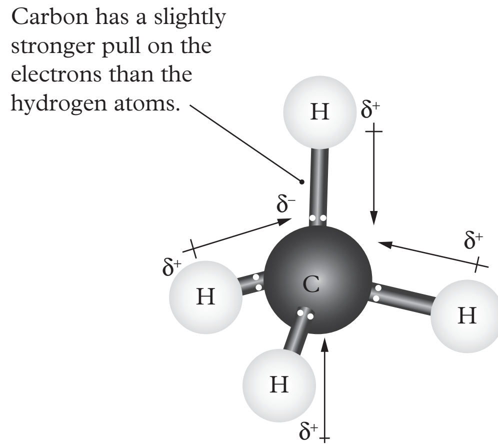


FIGURE 4 Methane is a symmetrical non-polar molecule. All the dipoles are equal and cancel each other out perfectly, making it non-polar.

Polar molecules

A polar molecule exists when the electrons are shared unequally between the atoms and the shape of the molecule is asymmetrical. The molecule has an overall dipole.

If the atoms in a molecule have different electronegativities:

- the bonding electrons stay closer to the more electronegative atom
- one end of the bond has a greater negative charge
- charge across the molecule is uneven (i.e. polar).

Linear molecules

Some linear diatomic molecules are polar, such as hydrogen fluoride (HF) (Figure 5).

The great difference in electronegativity between hydrogen and fluorine means that the fluorine atom has a much stronger pull on the electrons. The fluorine atom has a greater negative charge and therefore is the negative end of the dipole (δ–). The hydrogen atom has a slight positive charge (δ+) because it has a smaller share of the electrons in the bond.

The greater the difference in electronegativity between the two atoms, the more polar the bond is. For example, the hydrogen fluoride bond (H–F) is more polar than the hydrogen chloride bond (H–Cl) because fluorine has a higher electronegativity than chlorine.

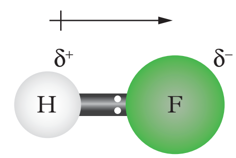


FIGURE 5 Hydrogen fluoride (HF) is a polar molecule because the electron pair is shared unevenly. Fluorine has a greater electronegativity than hydrogen, so it pulls the electrons closer to its nucleus.

Study tip

A bond is considered polar when the difference in electronegativity between atoms is more than 0.4.

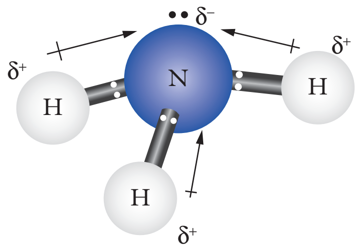


FIGURE 6 Ammonia is a polar molecule.

Non-linear asymmetrical molecules

Polyatomic molecules that are asymmetrical are also polar because their dipoles do not cancel each other out. For example, ammonia (NH_3) is a pyramidal molecule (Figure 6) with a nitrogen atom bonded to three hydrogen atoms, and a non-bonding pair of electrons.

Nitrogen has a higher electronegativity than hydrogen, so it pulls the electrons away from the hydrogen atoms, which makes the three bonds polar. The molecule is asymmetrical, so the dipoles do not cancel each other out. This makes ammonia a polar molecule.

Molecular shape and the effect on polarity

Symmetry and shape have a large effect on the polarity of molecules. You will notice that the bonds in carbon dioxide are polar and have dipoles because of the difference in electronegativity between carbon and oxygen. However, these dipoles orientate in the opposite directions, so they cancel each other out (Figure 7). By contrast, water is a polar molecule because the dipoles do not cancel each other out.

In Table 1, you can see some different examples of each molecular shape and its symmetry and polarity.

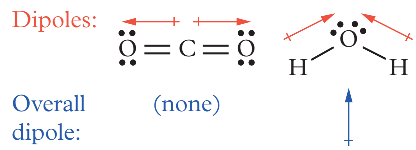


FIGURE 7 In carbon dioxide (CO_2), the dipoles cancel each other out, so the molecule is non-polar. In water (H_2O), the dipoles do not cancel out, so the molecule is polar.

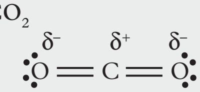
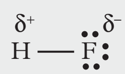
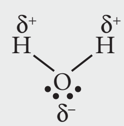
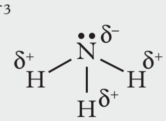
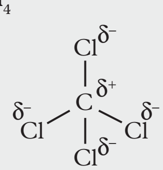
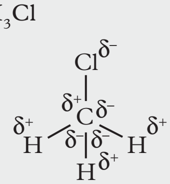
Study tip

A general rule of thumb is that symmetrical molecules are non-polar and asymmetrical molecules are polar. However, there are some exceptions. Make sure to carefully compare the electronegativity of the bonded atoms.

Study tip

When you want to determine the dipoles of a bond, remember to assign δ+ to the element with lowest electronegativity and δ− to the element with highest electronegativity.

TABLE 1 The effect of molecular shape and symmetry on polarity

Shape	Example	Symmetry	Polarity
Linear	CO_2 	Symmetrical	Non-polar
	HF 	Asymmetrical – electrons are shared unevenly	Polar
V-shaped	H_2O 	Asymmetrical	Polar
Pyramidal	NH_3 	Asymmetrical	Polar
Tetrahedral	CCl_4 	Symmetrical	Non-polar
	CH_3Cl 	Asymmetrical – the single chlorine atom and three hydrogen atoms make this asymmetrical	Polar

Compare bond and overall molecule polarity in Worked examples 3.3A and 3.3B.

3.3A WORKED EXAMPLE

COMPARING BOND POLARITY

Compare the polarity of the bonds in the molecules NO and CO.

Solution

Think	Do
Step 1: Find the electronegativity of each atom from Figure 1.	N = 3.0 O = 3.4 C = 2.6
Step 2: For each molecule, subtract the lower electronegativity from the higher electronegativity.	NO = O – N = 3.4 – 3.0 = 0.4 CO = O – C = 3.4 – 2.6 = 0.8
Step 3: Identify the bond that has the greater difference in electronegativity.	CO has a greater difference in electronegativity than NO.

3.3B Worked example
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3.3B Worked example
Video demonstration

3.3 SKILL DRILL

Analysing and evaluating polarity data

Key science skill: Analyse and evaluate data and investigation methods

The polarity of a solvent is indicated by its **polarity index**. For example, pentane is a non-polar substance and has a polarity index of 0.0; water is a polar substance and has a polarity index of 10.0.

The polarity indexes of three solvents (methanol, acetonitrile and toluene) are determined experimentally by measuring their resistance when voltage is applied. The results are shown in Table 2 along with the theoretical polarity indexes.

Practise your skills

1 Calculate the mean experimental polarity index for each solvent.

2 Choose the correct graph style to present the means you calculated in Question 1. Make sure you include the following elements: title, labels, units, legend (if applicable), line of best fit (if applicable).

3 Evaluate the results, by discussing the accuracy and precision of the measurements.

4 Identify any potential outliers.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

TABLE 2 The experimental and theoretical polarity indexes of methanol, acetonitrile and toluene.

Solvent	Experimental polarity index			Theoretical polarity index
Methanol (CH_4O)	5.1	5.3	4.9	5.1
Acetonitrile (C_2H_3N)	5.4	5.2	6.5	5.8
Toluene (C_7H_8)	3.7	3.8	3.8	2.4

polarity index
a measure of a substance's polarity determined by the resistance when a voltage is applied

3.3 CHECK YOUR LEARNING

Describe and explain

1 Describe the effect that molecular shape has on the polarity of a molecule.

2 Identify and explain which bond has the highest polarity: O–H or S–O?

Apply, analyse and compare

3 Some covalent bonds are shown below.

- **i** N–F
- **ii** H–C
- **iii** O–S
- **iv** P–O
- **v** Cl–P
- **vi** N–C

a For each pair of atoms, identify which atom has the highest electronegativity.

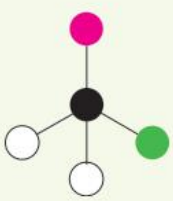
b For each pair of atoms, determine the dipoles of the bond.

c Determine which of the bonds has the highest polarity.

d Determine which of the bonds has the lowest polarity.

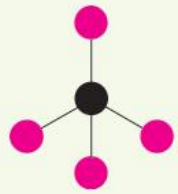
4 Analyse the following general molecular diagrams and determine if the molecules are polar or non-polar. Different colours represent atoms with different electronegativities.

a

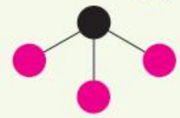


b

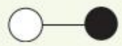
c



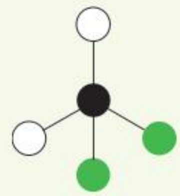
d



e



f



5 Draw the following molecules with the correct molecular shape, then determine the polarity of each molecule.

a HCN

b CH_3OH

c H_2S

d SO_3

e PCl_3

Design and discuss

6 Discuss the difference between a polar bond and a polar molecule. Why can a non-polar molecule have polar bonds? Explain your reasoning with an example.

7 BF_3 is shown below. It is a trigonal planar molecule. Discuss whether the molecule is polar or non-polar.

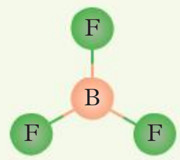


FIGURE 8 Polar and non-polar solvents are used in chemistry.

3.4 The relative strength of bonds

KEY IDEAS

In this topic, you will learn that:

+ covalent molecules can be attracted to each other through dispersion forces, dipole–dipole attractions and hydrogen bonding
+ intramolecular forces occur within molecules and intermolecular forces occur between molecules.

intramolecular bond
a bond within a molecule

metallic bond
a chemical bond that results from the electrostatic attraction between positive metal cations and negative delocalised electrons.

ionic bond
a chemical bond that results from the electrostatic attraction between a positive metal cation that has lost electrons and a negative non-metal anion that has gained electrons

intermolecular force
an attraction between molecules

So far in this chapter, you have learnt about covalent bonding. Covalent bonds are a type of **intramolecular bond**; **metallic** and **ionic bonds** are other types of bonds that you will look at in Chapters 4 and 5.

Covalent bonds are a very strong type of intramolecular bond. Because electrons are shared between atoms, it takes a lot of energy to break covalent bonds.

Intermolecular forces

You know that water (H_2O) is a covalent molecule. It is an oxygen atom covalently bonded to two hydrogen atoms. But a glass of water is not just one water molecule. It is hundreds of thousands of millions of water molecules. The molecules are not covalently bonded to each other. If they were, the water would probably be a solid and you would not be able to drink it.

Instead, individual water molecules are attracted to each other by **intermolecular forces**. The forces between water molecules are a lot weaker (about 100 times) than the covalent bonds between the hydrogen and oxygen atoms. This is why you can drink the liquid water.

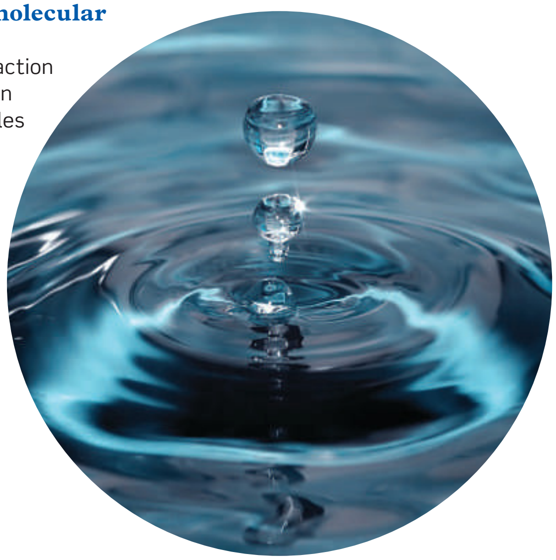


FIGURE 1 Liquid water contains many intermolecular bonds.

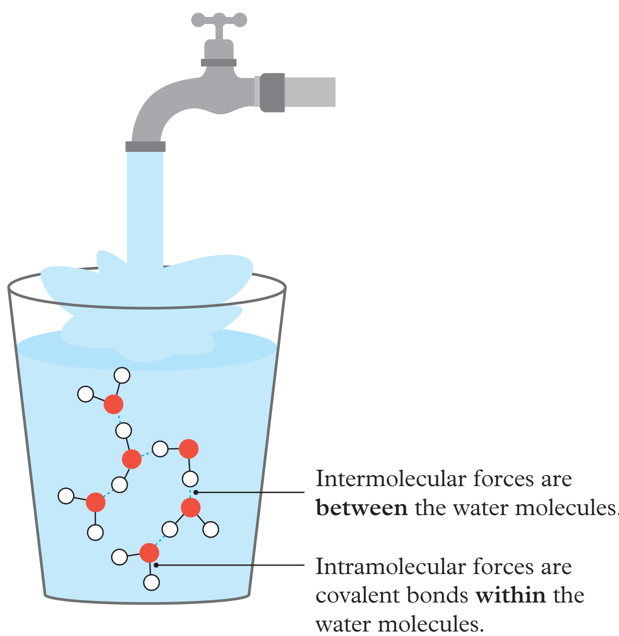


FIGURE 2 Water has intramolecular forces within the covalent molecules and intermolecular forces between the molecules.

To picture intermolecular and intramolecular forces, imagine you were asked to build a tower of blocks using Super Glue and Blu Tack.

You use the Super Glue to attach one H block to one F block. The Super Glue represents a covalent bond. Because it's Super Glue, it requires a lot of strength to separate the H–F blocks. Once you have several H–F blocks, you use the Blu Tack to join them together (Figure 3a). The Blu Tack represents the intermolecular forces between the H–F blocks.

Your little sister walks into the room and pushes your tower over. H–F blocks separate from other H–F blocks. Unsurprisingly, Blu Tack is not strong enough to hold them together (Figure 3b). Intermolecular forces can be overcome more easily than intramolecular bonds because they are weaker. The Super Glue holds. Like intramolecular bonding, it is very strong, and the H and F blocks are hard to break apart.

Study tip

Intramolecular bonding is within molecules (e.g. covalent bonding). Intermolecular forces are between molecules (e.g. dispersion forces, dipole–dipole attractions and hydrogen bonding).

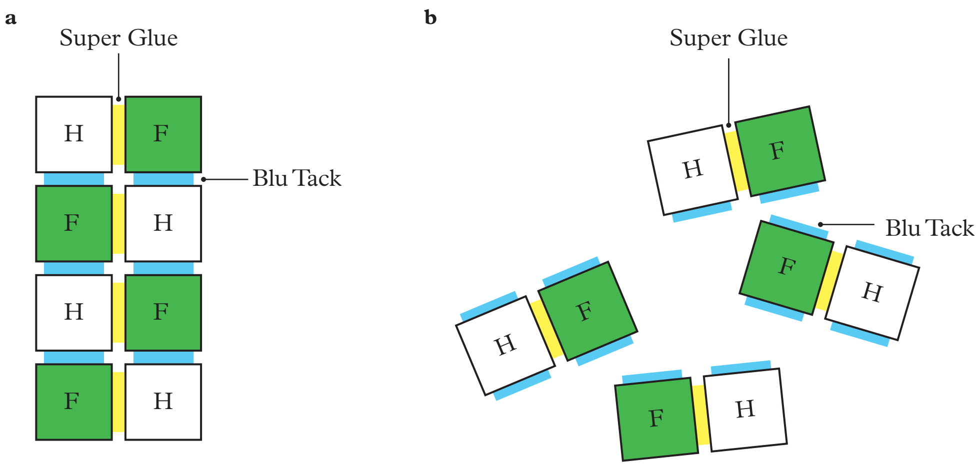


FIGURE 3 **a** A tower made of H and F blocks held together by Super Glue (yellow) and Blu Tack (blue). **b** The tower breaks because the Blu Tack (intermolecular forces) is not strong enough to withstand it being pushed over. The individual H–F blocks (molecules) do not break apart because the Super Glue (intramolecular bonds) is strong enough to hold them together.

There are three types of intermolecular forces:

- dispersion forces
- dipole–dipole attraction
- hydrogen bonding.

They have different strengths and give different properties to molecules. The type of intermolecular forces present between covalent molecules depends on the atoms present and the polarity and shape of the molecules.

FIGURE 4 The building blocks represent different atoms.

Dispersion forces

Dispersion forces are the weakest intermolecular force. They exist between all covalent molecules, regardless of their polarity and size. They occur because the shared electrons within the molecular bonds are constantly and randomly moving to form **temporary dipoles**, or **dipole moments**.

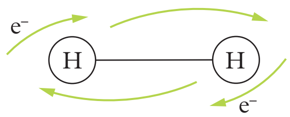
You can explore how dispersion forces form in Figure 5, which uses hydrogen gas (H_2) as an example.

The size and the shape of the molecules affect the strength of the dispersion forces between them. You will learn about how this affects the properties of covalent molecules in Topic 3.5.

dispersion force
an intermolecular force that results from attraction between temporary dipoles in polar and non-polar molecules

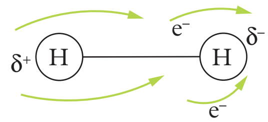
temporary dipole (or dipole moment)
when a bond gains a temporary negative charge at one end and a temporary positive charge at the other end

1



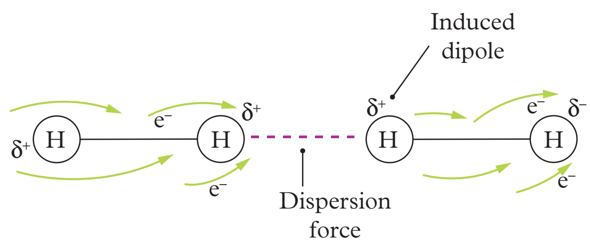
The shared electrons in H_2 move randomly around the bond. Overall, they spend the same amount of time at each of the hydrogen atoms.

2



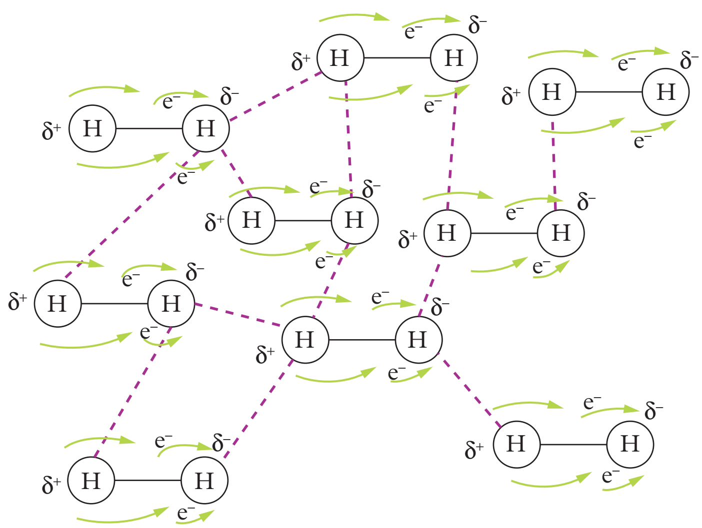
Electrons can gather closer at one end to create a temporary negative charge at one end and temporary positive charge at the opposite end. This is called a temporary dipole (dipole moment).

3



When a H_2 molecule with a dipole moment gets close to another H_2 molecule, it creates a dipole in the second H_2. Electrons in the first H_2 molecule repel the electrons in the second H_2 and cause an induced dipole.

4



As a consequence, H_2 molecules with dipoles induce dipoles in many other nearby H_2 molecules. The attraction between the temporary dipoles is called a dispersion force.

FIGURE 5 Dispersion forces form between temporary dipoles in covalent hydrogen gas (H_2) molecules.

Dipole–dipole attractions

When a molecule is polar, it has a permanent dipole. One end of the molecule is always slightly positive (δ^+) and the other always slightly negative (δ^-).

dipole–dipole attraction an intermolecular force that forms as a result of attraction between permanent dipoles in polar molecules

Because of their permanent dipoles, polar molecules can have **dipole–dipole attractions** or forces between them. For hydrogen chloride (HCl), the negative end of one dipole (the chlorine atom) attracts the positive end of a dipole of a nearby molecule (the hydrogen atom). This is shown in Figure 6.

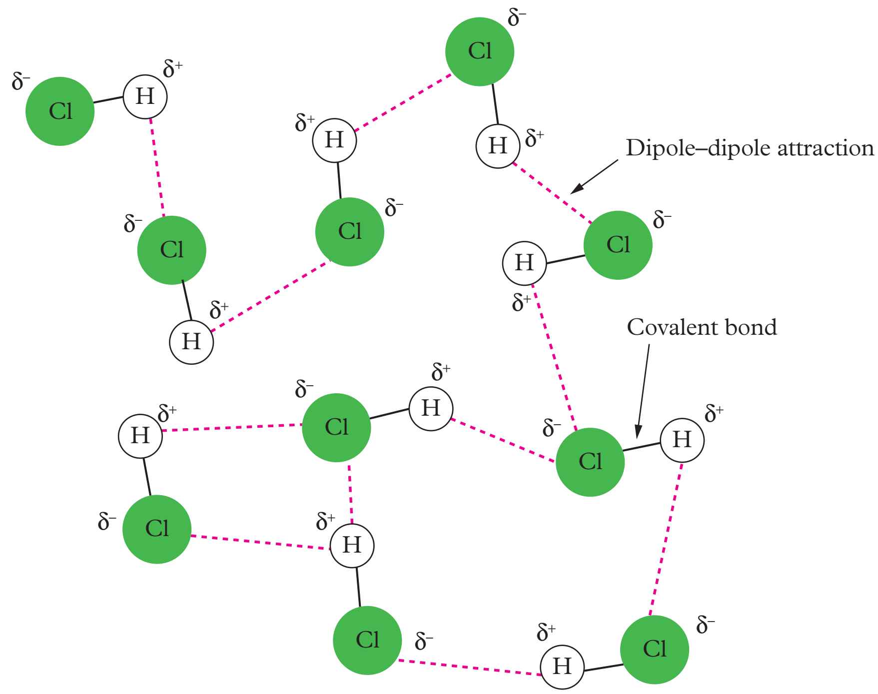


FIGURE 6 The dipole–dipole attractions between the permanent dipoles in hydrogen chloride (HCl) molecules

Study tip

The more polar a molecule, the stronger the dipole–dipole attractions are.
A molecule has a higher polarity when:
- there is a big difference in electronegativities between the atoms
- the molecule has large asymmetry.

The strength of the dipole–dipole attractions depends on the polarity of the molecule. The more polar the molecule (because of a large difference in electronegativity or a large molecular asymmetry), the stronger the dipole–dipole attractions between the molecules are. This affects the properties of polar molecules, which you will explore in Topic 3.5.

The dipoles in hydrogen chloride are small, so the strength of the dipole–dipole forces are relatively weak. Despite this, dipole–dipole attractions are stronger intermolecular forces than dispersion forces.

Hydrogen bonding

Hydrogen bonds are a much stronger type of dipole–dipole force that occur in very special circumstances. To hydrogen bond, molecules must have a hydrogen atom covalently bonded to a fluorine, oxygen or nitrogen atom.

When attached to one of these small, highly electronegative elements, the single electron on the hydrogen is drawn away from the hydrogen atom. The hydrogen gains a strong partially positive charge. It can strongly attract the negatively charged lone pair (non-bonding) electrons on neighbouring F, O and N atoms (Figure 7). This makes hydrogen bonding stronger than dipole–dipole forces.

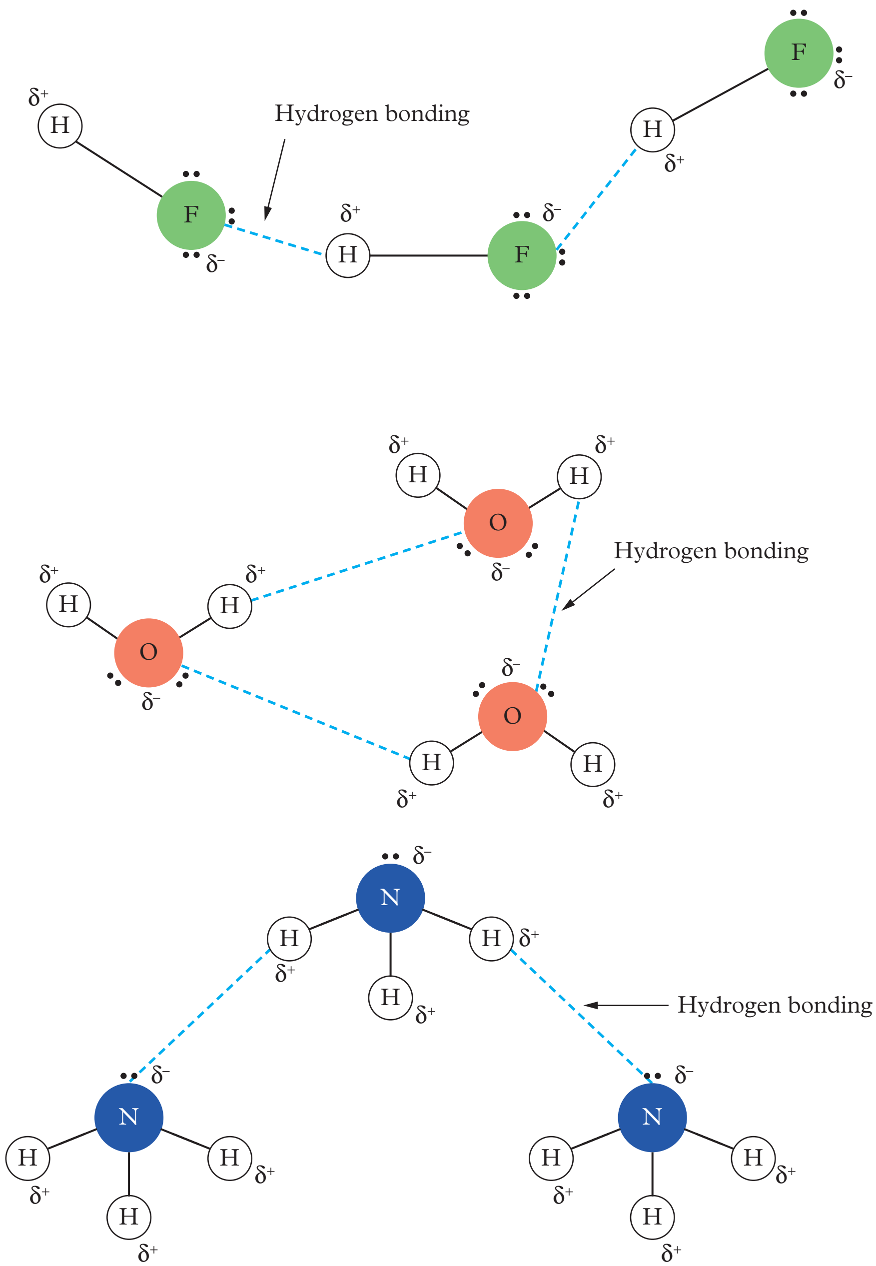


FIGURE 7 Hydrogen bonding between **a** hydrogen fluoride molecules, **b** water molecules and **c** ammonia molecules

Hydrogen bonds can form between a hydrogen atom on one molecule and any molecule that has a fluorine, oxygen and nitrogen, even if the second molecule has no hydrogen.

hydrogen bond
an intermolecular force between a hydrogen atom that is connected to a fluorine, oxygen or nitrogen atom in one molecule, and a fluorine, oxygen or nitrogen atom on a nearby molecule, which has at least one lone pair of electrons

Study tip

Hydrogen bonding only occurs when a hydrogen is attached to the atoms that are FON!
That is: fluorine, oxygen and nitrogen.

For example, water can form hydrogen bonds with an aldehyde such as formaldehyde. Formaldehyde on its own is a gas because it only has dipole–dipole attractions between the molecules (Figure 8a). Although it contains both O and H in its structure, they are not directly bonded. Formaldehyde therefore cannot undergo hydrogen bonding with itself.

However, the oxygen in formaldehyde (with its lone pairs of electrons) can form hydrogen bonds with highly positive (δ^+) hydrogen atoms in water molecules (Figure 8b). This mixture then becomes useful as an embalming fluid.

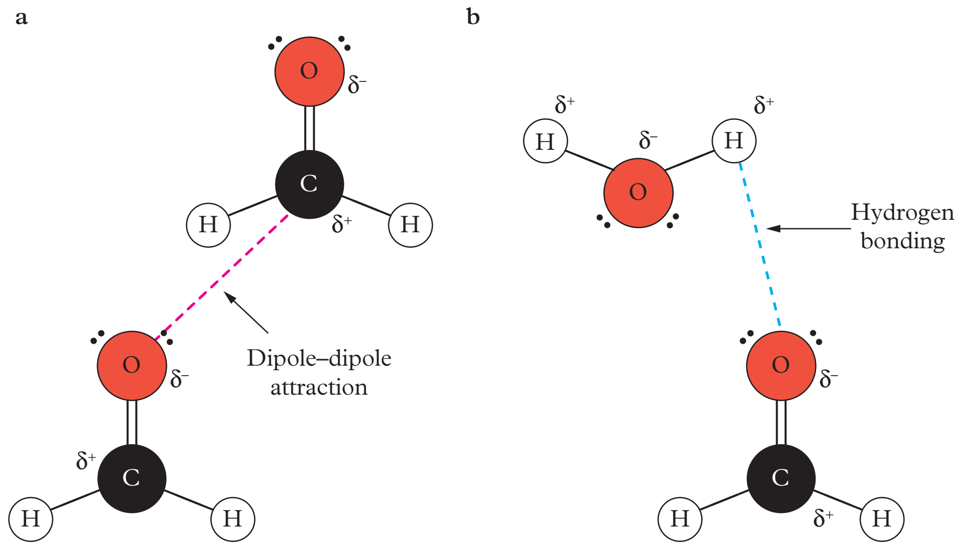
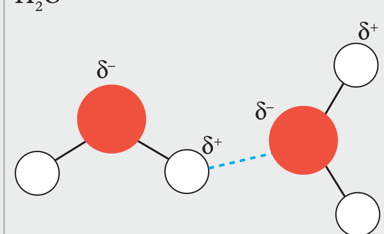
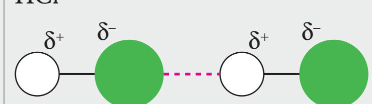
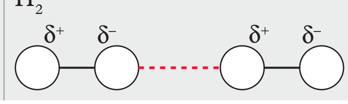


FIGURE 8 **a** Formaldehyde molecules only have dipole–dipole attractions with other formaldehyde molecules. **b** However, formaldehyde can form hydrogen bonds with water because the oxygen in formaldehyde has two lone pairs of electrons.

Comparing intermolecular force strength

The three types of intermolecular force that occur between covalent molecules have different strengths. This is summarised in Table 1.

TABLE 1 Comparing intermolecular force strength

Intermolecular force	Relative strength	Present in	Example
Hydrogen bonding	Strongest	Molecules that contain an H bonded to F, O or N	H_2O 
Dipole–dipole attraction	Medium	Polar molecules	HCl 
Dispersion force	Weakest	All molecules and atoms	H_2 

Study tip

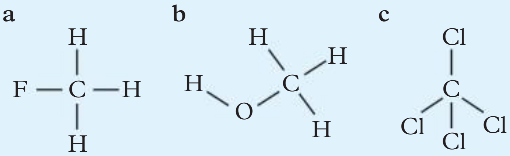
The relative strengths of intermolecular forces are: hydrogen bonding > dipole–dipole attraction > dispersion force.

3.4 WORKED EXAMPLE

DETERMINING THE INTERMOLECULAR FORCES PRESENT

Analyse each of the following molecules and determine the strongest intermolecular forces that are present between each of these molecules

a b c



Solution

Use the questioning flow chart in Figure 9 to help you.

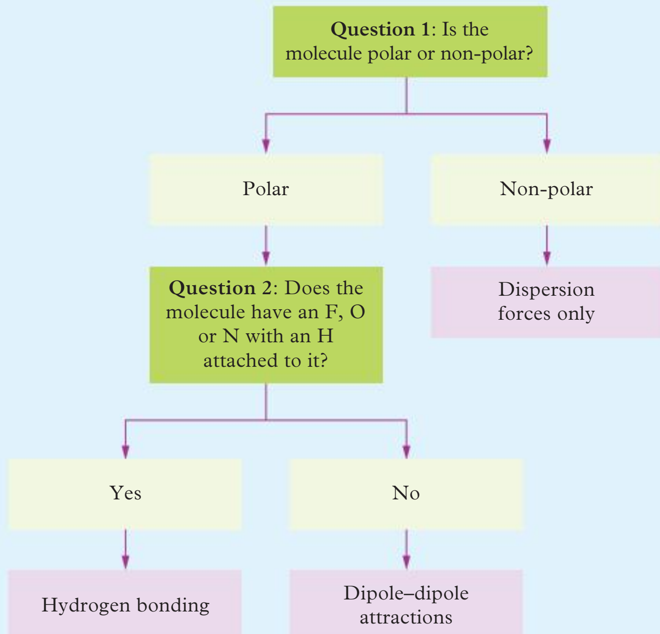


FIGURE 9 A questioning flow chart for intermolecular forces

Think	Do
Step 1: For each molecule, ask the first question: Is it polar or non-polar?	**a** Polar **b** Polar **c** Non-polar
Step 2: If it is non-polar, there are only dispersion forces present. If it is polar, ask the next question: Does the molecule have an F, O, N with an H attached to it?	**a** Has an F but there is no H attached, so no **b** Has an O with an H attached, so yes **c** Only dispersion forces are present
Step 3: Identify the strongest intermolecular force present.	**a** Dipole–dipole attraction is the strongest intermolecular force present. **b** Hydrogen bonding is the strongest intermolecular force present. **c** Dispersion forces are the strongest intermolecular force present.

3.4 CHECK YOUR LEARNING

Describe and explain

1 Describe the circumstances required for each of the following intermolecular forces.

- **a** Dispersion forces
- **b** Dipole–dipole attractions
- **c** Hydrogen bonding

2 Identify which of Cl_2, HCl or CH_3Cl has the strongest dipole–dipole attraction between its molecules.

Apply, analyse and compare

3 Analyse each of the following molecules and determine the strongest intermolecular forces present between them.

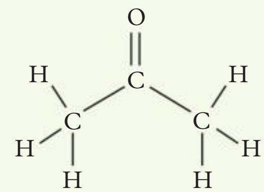
- **a** HCN
- **b** $CHCl_3$
- **c** H_2
- **d** CO_2
- **e** CH_3OH

4 Compare the substances in Question 3 and put them in order of strongest to weakest intermolecular forces.

5 Compare the three intermolecular forces and explain why the strengths of the intermolecular forces are different.

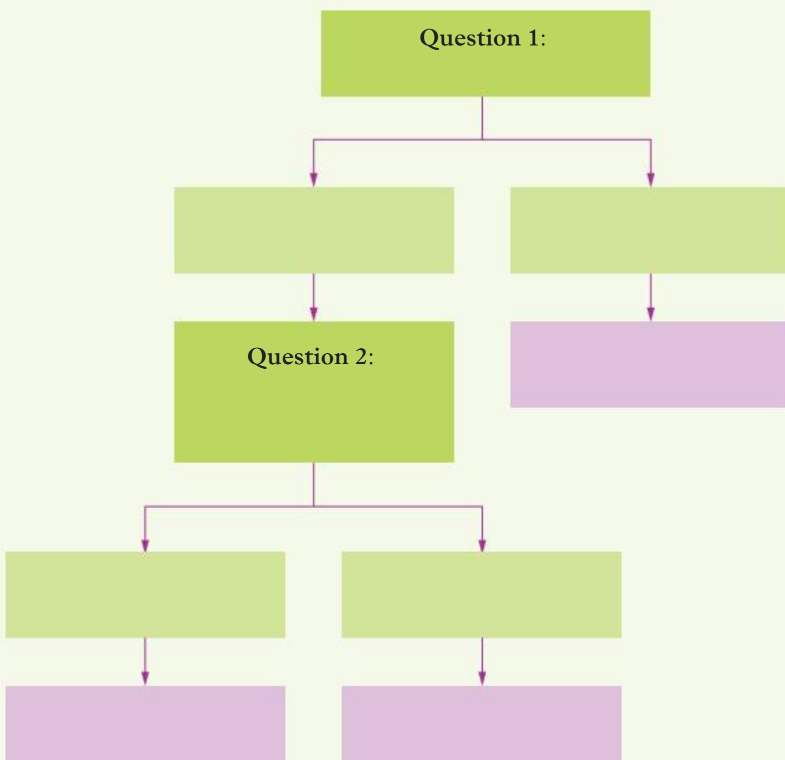
Design and discuss

6 Propanone molecules cannot hydrogen bond with each other, but they can form hydrogen bonds with water. Discuss why this can occur.



7 In Worked example 3.4, there is a questioning flow chart to help determine the strongest intermolecular forces present.

Design a set of questions for the flow chart that will enable you to determine the polarity of a molecule *before* determining the intermolecular forces. Then add a set of questions that will enable you to determine the strength of the intermolecular forces *after* they have been identified. A template has been provided for you below.



3.5 Physical properties of molecular substances

KEY IDEAS

In this topic, you will learn that:

+ the physical properties of covalent substances are determined by the structure of the substance and the intermolecular forces present.

physical property
a characteristic that can be observed or measured without changing the identity of the substance

Covalent molecules have a variety of different **physical properties**. These are characteristics that can be observed or measured without changing the identity of the substance. The physical properties of covalent molecules that you will look at in this chapter are:

- boiling point – the temperature at which a liquid becomes a vapour (gas)
- melting point – the temperature at which a solid becomes a liquid
- electrical conductivity – the ability of a substance to allow charged particles to move through it.

Physical properties are determined by the size, polarity and intermolecular forces in a covalent molecule.

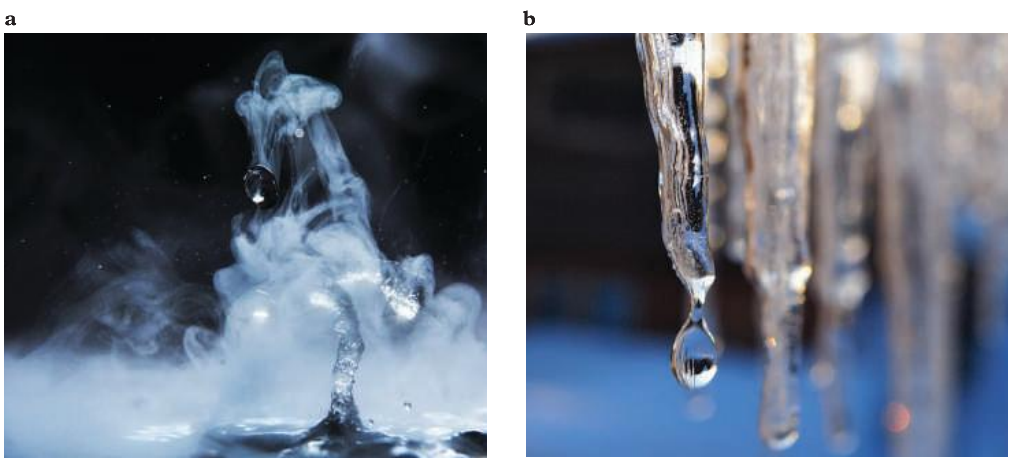


FIGURE 1 Covalent molecules have different physical properties, including **a** boiling point and **b** melting point.

Physical properties of covalent molecules

boiling point
the temperature at which a substance changes state from liquid to gas

melting point
the temperature at which a substance changes state from solid to liquid

Covalent molecules have relatively low **boiling points** and **melting points** compared with metals and ionic compounds, which you will look at in Chapters 4 and 5. Although covalent bonding is a very strong form of intramolecular bonding, it is the intermolecular forces that determine the physical properties of covalent molecules.

Effect of intermolecular forces on melting and boiling points

Melting point and boiling point depend on the energy required to overcome intermolecular forces. When a molecule melts, its intermolecular forces weaken. When it boils, its intermolecular forces break. Since intermolecular forces have different strengths, the types of forces present are important.

3.5
PRACTICAL:
CONTROLLED EXPERIMENT

How do intermolecular forces influence the physical properties of covalent molecules? Go to page **498**.

Molecules that form hydrogen bonds have higher melting and boiling points than molecules that only have dipole–dipole attractions. They are typically solids or liquids at room temperature. Molecules with only dispersion forces have the lowest melting and boiling points and are usually gases.

You can see this in the graph in Figure 2, where three molecules with similar molar mass have significantly different melting and boiling points due to the intermolecular forces present.

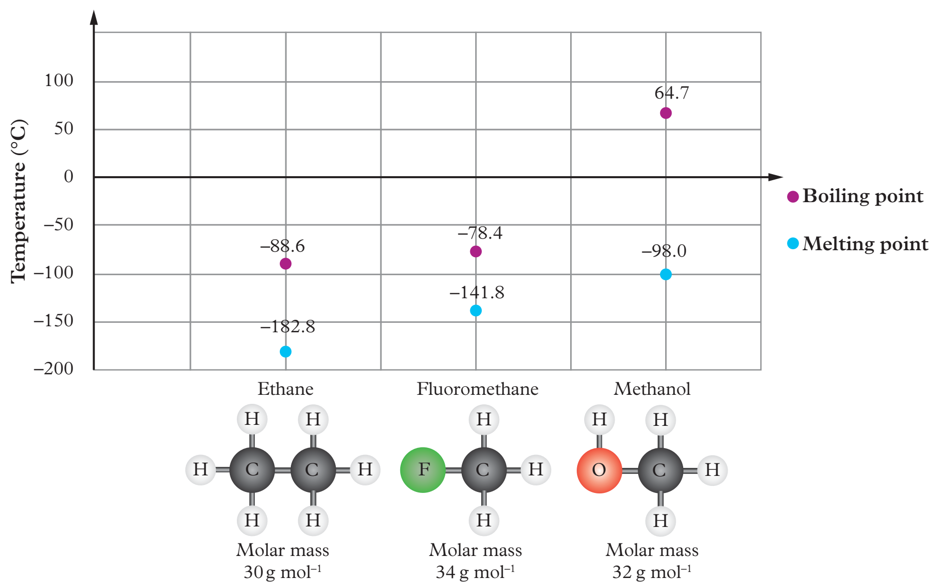


FIGURE 2 The effect of intermolecular forces on the melting and boiling points of ethane, fluoromethane and methanol

The strong hydrogen bonds between oxygen and hydrogen atoms in different methanol molecules need a lot of energy to break apart, so the boiling point and melting point of methanol is much higher than that of fluoromethane or ethane.

Ethane, which is non-polar and only has dispersion forces present, has a much lower melting point and boiling point than the other two molecules. This is because it does not require a lot of energy to break the weak dispersion forces. This also means that ethane is a gas at room temperature, whereas methanol is a liquid.

In terms of strength, dipole–dipole attractions are between hydrogen bonds and dispersion forces. It takes slightly more energy to break the dipole–dipole attractions between fluoromethane molecules than it does to break the dispersion forces between ethane molecules. Fluoromethane is a gas at room temperature.

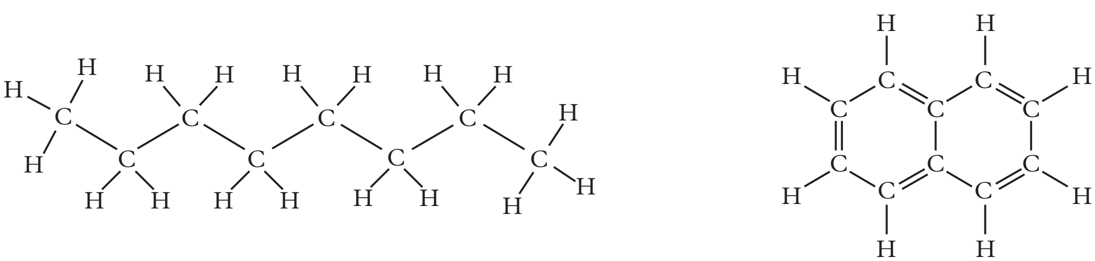
FIGURE 3 Covalent molecules have different boiling points.

Some more examples of boiling points are shown in Table 1. Generally, a molecule that has a high boiling point also has a high melting point.

TABLE 1 Some physical properties of small covalent molecules

Molecule	Molar mass (g mol^{-1})	Strongest intermolecular force present	Boiling point (°C)	State at room temperature
Hydrogen (H_2)	2.0	Dispersion forces	–252.9	Gas
Nitrogen (N_2)	28.0	Dispersion forces	–195.8	Gas
Oxygen gas (O_2)	32.0	Dispersion forces	–183	Gas
Hydrogen chloride (HCl)	36.5	Dipole–dipole attractions	–85.05	Liquid
Water (H_2O)	18.0	Hydrogen bonding	100	Liquid
Butanol ($C_4H_{10}O$)	74.0	Hydrogen bonding	117.7	Liquid

Although polarity and intermolecular forces are good indicators of melting and boiling points, not every non-polar molecule is a gas at room temperature. For example, at room temperature, octane (a component of petrol) is a liquid and naphthalene (used in moth balls) is a solid.



Octane (liquid) **Naphthalene (solid)**

FIGURE 4 The non-polar molecules octane and naphthalene are liquid and solid at room temperature, respectively.

Effect of molecular size on melting and boiling points

All molecules have dispersion forces between them, regardless of their polarity. The strength of the dispersion forces increases as the size of the molecule increases.

Consider the butanol molecule. Although butanol is polar and can form hydrogen bonds, it has a longer non-polar chain than methanol. This means that there are many electrons, and temporary dipoles are formed more easily than molecules with shorter non-polar chains. The more dipoles there are, the greater the attraction between butanol molecules and the stronger the dispersion forces between them. This contributes to the high melting and boiling points of butanol, compared with a shorter chain molecule such as methanol (Figure 5).

Study tip

Melting and boiling points of covalent molecules increase as intermolecular forces increase in strength: hydrogen bonding > dipole–dipole attractions > weak dispersion forces

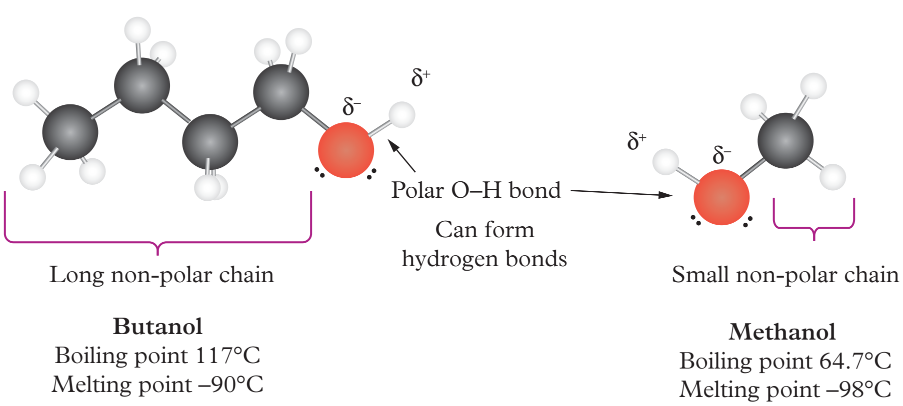


FIGURE 5 Butanol has a longer non-polar chain than methanol, making its melting and boiling points higher.

Study tip

The longer the molecule within a homologous series, the higher the boiling point.

In Table 2, you can see that larger non-polar hydrocarbons have higher boiling points than smaller non-polar hydrocarbons. They also have higher melting points. This is because it takes more energy to overcome the dispersion forces between the larger molecules.

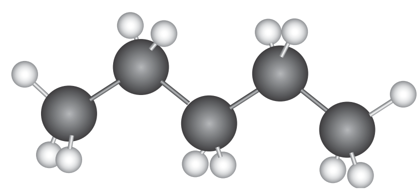
TABLE 2 The boiling points of non-polar hydrocarbons increase as the size of the molecule increases

Molecule	Molar mass (g mol^{-1})	Boiling point (°C)
Methane (CH_4)	16.0	–161.6
Butane (C_4H_{10})	58.0	–1
Octane (C_8H_{18})	114.0	125.6
Dodecane ($C_{12}H_{26}$)	170.0	216.2

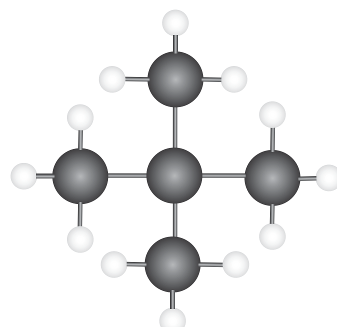
Effect of molecular shape on melting and boiling points

The shape of a molecule also affects the strength of the dispersion forces. Long straight-chain molecules have stronger dispersion forces than smaller, more compact molecules with the same number of electrons.

For example, pentane and 2,2-dimethylpropane have the same molecular formula, C_5H_{12} (Figure 6). They have the same number of electrons available to form temporary dipoles. However, because pentane is longer and less compact, it has more points where it can interact with other pentane molecules and create more, stronger, dispersion forces. The compact shape of 2,2-dimethylpropane gives it a much lower boiling point.



Pentane
Boiling point 36.1°C



2,2-Dimethylpropane
Boiling point 9.5°C

FIGURE 6 Pentane has a higher boiling point than 2,2-dimethylpropane because it has a less compact shape.

3.5 CHALLENGE

Linking ideas together: polarity, shape and boiling point

Compare each pair of molecules i–vi by drawing their structures and determine which:

a is the most polar

b has the highest boiling point.

- **i** Methane (CH_4) and dichloromethane (CH_2Cl_2)
- **ii** Sulfur dioxide (SO_2) and carbon dioxide (CO_2)
- **iii** Carbon disulfide (CS_2) and sulfur difluoride (SF_2)
- **iv** Nitrogen trichloride (NCl_3) and oxygen dichloride (OCl_2)
- **v** Water (H_2O) and hydrogen peroxide (H_2O_2)
- **vi** Methanol (CH_3OH) and propanol ($CH_3CH_2CH_2OH$)

Electrical conductivity of covalent substances

Electrical conductivity is the ability of a substance to allow charged particles to pass through it. For a substance to conduct electricity, it must have charged particles present, such as in ionic and metallic substances, which you will learn about in Chapters 4 and 5. Covalent compounds do not have charged particles, so they generally do not conduct electricity.

There are a few exceptions to this. Covalently bonded acids can **ionise** in water and form charged protons (H^+ ions) and negative ions that can then conduct electricity. One example is sulfuric acid (Figure 7).

electrical conductivity
the ability to allow charged particles to move through a substance

ionise
form charged ions

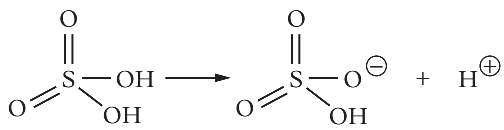


FIGURE 7 Sulfuric acid ionises to form a negatively charged hydrogen sulfate ion and a positively charged proton.

Graphite, which is a covalent lattice structure, has free-moving electrons that can carry charges and conduct electricity. In Chapter 9, you will look at some polymers that also have free electrons that can carry a charge. But, overall, covalent substances do not conduct electricity.

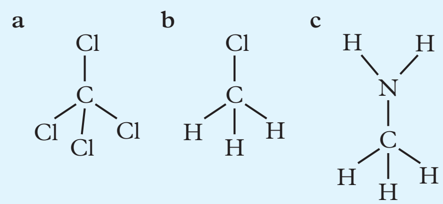
This property of covalent substances makes them useful **electrical insulators**, materials that prevent the movement of charges particles. One example is plastic, which is used to coat wires and prevent 'loss' of charges travelling through them.

electrical insulator
a material that does not allow the movement of charged particles through it

3.5 WORKED EXAMPLE

PREDICTING THE PHYSICAL PROPERTIES OF COVALENT COMPOUNDS

Compare the three covalent substances a–c and determine the order of boiling points of the molecules from highest to lowest.



Solution

Think	Do
Step 1: Determine the polarity of each of the molecules.	**a** Non-polar **b** Polar **c** Polar
Step 2: Use the questioning flow chart from Worked example 3.4 to determine the strongest intermolecular forces present for each molecule.	**a** Dispersion forces **b** Dipole–dipole attractions **c** Hydrogen bonding
Step 3: Determine the order of boiling points of the molecules from highest to lowest. Molecules with hydrogen bonding have the highest boiling point, followed by dipole–dipole attractions and then dispersion forces.	Final answer: c > b > a

3.5 CHECK YOUR LEARNING

Describe and explain

1 Explain why most covalent molecules cannot conduct electricity.

2 Match the boiling points to the molecules:
Molecules: Cl_2, H_2O, CH_4, HCl
Boiling points (°C): –161.6, –85.0, –34.6, 100

3 Explain why F_2 has a boiling point of –188°C, but HF has a boiling point of 19.5°C.

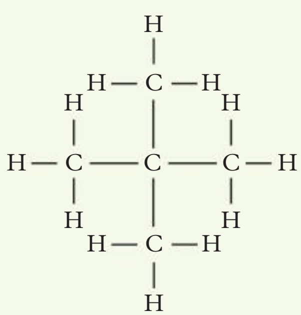
Apply, analyse and compare

4 Compare the terms 'melting point' and 'boiling point'.

5 Compare each set of molecules and order them from highest boiling point to lowest boiling point.

a CO_2, HF, HCN, O_2

b **i**



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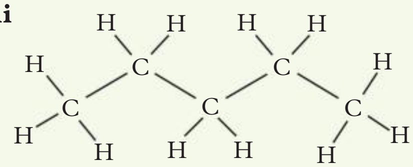

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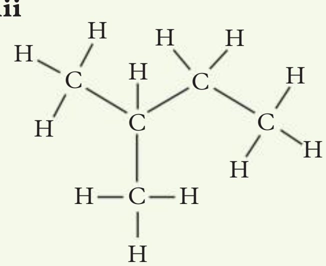
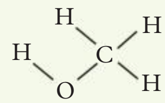


c **i**

ii



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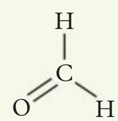

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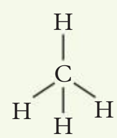


Design and discuss

6 Discuss the trend in boiling points of the following halogens (group 17 atoms) bonded to hydrogen.

Molecule	Boiling point (°C)
HF	19.5
HCl	–85.05
HBr	–66
HI	–35.36

3.6 The structure and bonding of diamond and graphite

KEY IDEAS

In this topic, you will learn that:

- diamond and graphite are both made from covalently bonded carbon atoms – they are different carbon allotropes
- the properties of diamond and graphite are directly linked to their chemical structures.

Carbon is a unique element because it can form four covalent bonds with a variety of other atoms. It can also form covalent bonds with other carbon atoms.

allotrope
a different structural form of an element

When carbon bonds with itself, the different forms it takes are called carbon **allotropes**. Each carbon allotrope has different properties. In this topic, you will learn about two allotropes of carbon – diamond and graphite.

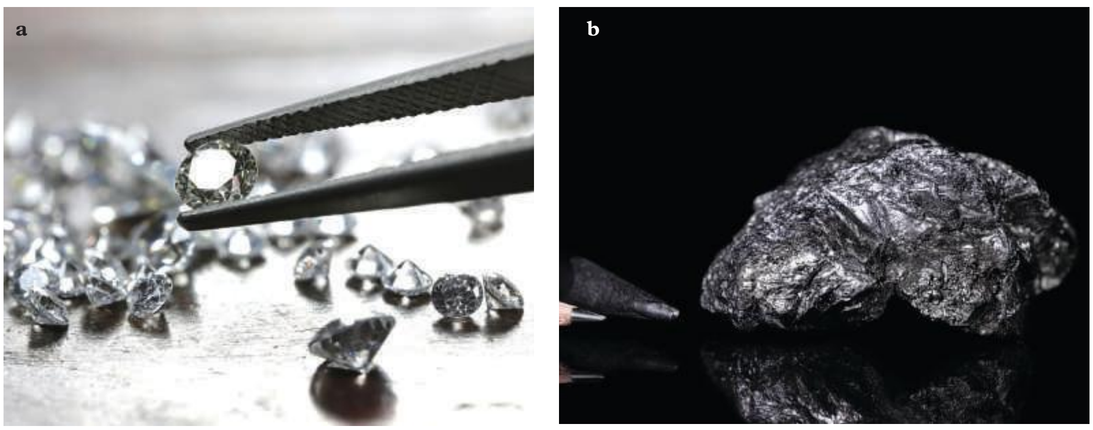


FIGURE 1 **a** Diamond and **b** graphite are allotropes of carbon.

Diamond

Diamonds are beautiful and can be found in many different sizes, shapes and colours. They form naturally underground over 1–4 billion years, at very high temperatures and pressures.

Structure of diamond

lattice
an interlaced structure consisting of regular repeated atoms or molecules

Diamond is made up of only carbon atoms. Each carbon atom is covalently bonded to four other carbon atoms in a covalent network **lattice**, as shown in Figure 2.

You will remember from Topic 3.2 that electron pairs position themselves as far away from each other as possible (according to VSEPR theory). This is why the structure of diamond is tetrahedral. The repeating arrangement makes diamond the hardest naturally occurring substance on Earth.

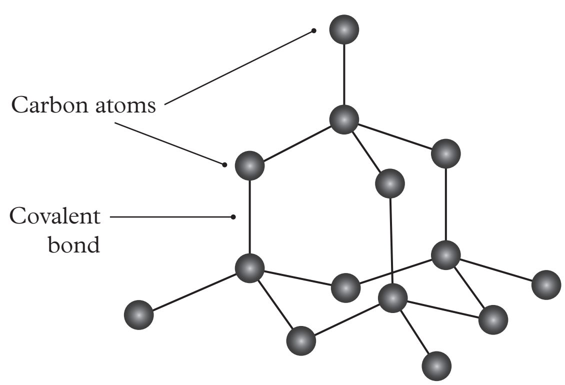


FIGURE 2 The structure of diamond

FIGURE 3 Diamonds are added to industrial tools to harden their surfaces and allow better grinding, drilling, cutting and polishing of hard materials.

Although the carbon atoms in diamond are all covalently bonded to one another, diamond is not considered to be a molecule. Instead, it is a network solid in which the number of carbon atoms varies depending on the size of the diamond. You cannot have a discrete diamond molecule, so diamond also does not have intermolecular bonding. Only the strong covalent bonds are present, giving diamond many unique properties.

Properties of diamond

Diamond is best known for its extremely high melting point. This means it is very difficult to break the covalent bonds to turn diamond from a solid into a liquid. Like all other covalent molecules, it is the structure of diamond that dictates this property and others (Table 1).

TABLE 1 The properties of diamond and the aspects of its structure which dictate them

Property	Why does it have this property?
Very hard	The entire structure consists of only very strong covalent bonds.
Rigid structure	Each electron pair is positioned as far from other electrons as possible to create a lattice structure that breaks rather than bends.
High melting point	Only having covalent bonds means that it takes an extremely large amount of energy to break them. At about 4000°C, diamond **sublimes** (changes from a solid to gas) rather than melts.
High thermal conductivity	Heat can pass through without disrupting the strong covalent bonds.
Does not conduct electricity	All the electrons are bonded and held tightly between the atoms, so they are not free to move and cannot conduct electricity.

sublime
to change state directly from a solid to a vapour

Applications of diamond

Apart from the obvious use of diamonds in jewellery, the properties of diamonds make them useful in various industries for drilling, cutting, polishing and grinding. As a gemstone, diamonds are valued for their colour, size and clarity. In industry, diamonds are used for their hardness and heat conductivity. Table 2 links the properties of diamonds to various applications.

TABLE 2 The properties of diamond mean that it has many applications

Property	Application	Description
Very hard	Industry	• Crushed into small particles, which are embedded in saw blades, drill bits and grinding wheels • Increases the cutting, grinding or drilling ability of these tools so they can be used on much harder materials
	Low-friction micro bearings	• Extremely resistant to abrasion and therefore durable • Reduces friction and wear in moving parts
High thermal conductivity	High-performance electronics	• Used to conduct heat away from sensitive parts within high-performance electronics and computers • Keeps heat-sensitive areas cool
Rigid structure	Speaker domes	• Increases the quality of diamond speaker domes • Can vibrate rapidly without deforming when sound waves are applied

Read Real-world chemistry 3.6 to learn about growing diamonds in the laboratory for industrial purposes.

3.6 REAL-WORLD CHEMISTRY

Laboratory-grown diamonds

Natural diamonds take up to four billion years to form under natural conditions – 140 km below the Earth's mantle, and under very high pressures and temperatures. Laboratory-grown diamonds are chemically and optically identical to natural diamonds. They are also produced under high temperatures and pressures, and are also made only of carbon … but they take only a few weeks to grow. That's much better than four billion years!

Natural diamonds are also often impure and can be flawed or weak because of structural defects. However, these flaws can make diamonds beautiful in jewellery. An example is the Hope diamond, one of the most recognisable diamonds in the world (Figure 4). It is blue because of an impurity, boron, in its structure. It is an attractive gem, but useless for purposes where the strength of a diamond is needed.

FIGURE 4 The Hope diamond is blue because it contains boron. This makes it beautiful, but not very strong.

Diamonds grown in the laboratory are always pure carbon, so they are reliably strong and beautiful and useful in industry, as well as the jewellery market. Laboratory-grown diamonds should not be confused with simulated diamonds, such as cubic zirconia, zircon and white sapphire. These are not chemically similar to natural or laboratory-grown diamonds; they are just manufactured to look like a diamond.

Growing a diamond can be done in three steps.

1 Place a seeding diamond – a tiny diamond crystal that will be the blueprint for the carbon to start growing on – in a high-pressure vacuum chamber.

2 Heat the chamber to 3000°C and add methane (CH_4) and hydrogen (H_2) gas to form plasma. The methane will donate carbon atoms to the seeding diamond.

3 Allow a few weeks for the diamond to grow!

This process forms a high-quality pure carbon diamond, with identical chemical and optical qualities to natural diamonds. In many cases, they are also more pure than natural diamonds and far stronger. This makes them perfect for industries that require hard diamonds, such as electronics, optics, sensors, lasers and computers.

Apply your understanding

1 Explain the similarities in structures between laboratory-grown and natural diamond.

2 Describe why laboratory-grown diamonds can be stronger than natural diamonds.

3 Discuss the benefits of using laboratory-grown diamonds in industry.

4 Compare the chemical structures of laboratory-grown diamonds and natural diamonds.

FIGURE 5 A real diamond and a laboratory-grown diamond – can you tell the difference?

Graphite

Graphite is an allotrope of carbon with a very different structure and properties from diamond. Where diamond is very hard and non-conductive, graphite is soft and conducts electricity. Like diamond, graphite is formed under the Earth's surface over millions of years and under high temperatures and pressures.

Structure of graphite

Like diamond, graphite is made only of carbon, but each carbon atom is covalently bonded to three other carbon atoms to form a covalent layer lattice (Figure 6). Because each atom only has three covalent bonds, there is one lone unpaired electron for every carbon atom. This **delocalised electron** is not attached to any particular carbon or bond but is free to move about.

delocalised electron
an electron that is not linked to a particular atom or single covalent bond; it is free to move within the structure

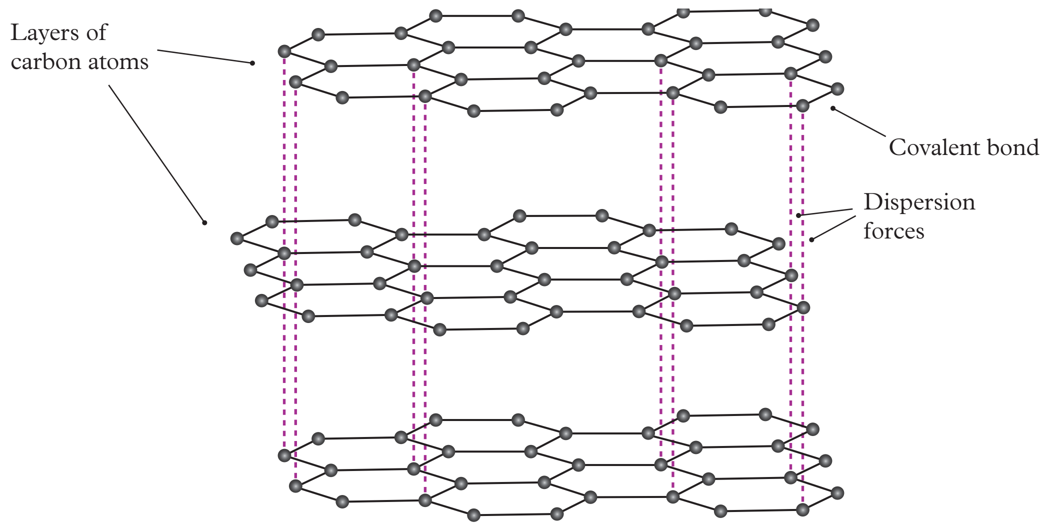


FIGURE 6 The structure of graphite

The layers are held together by dispersion forces created by the delocalised electrons. The electrons cannot move *between* sheets or interact with electrons in other sheets, but they can move freely *throughout* a layer because they are not fixed to one particular carbon atom.

You will remember that covalent bonding is very strong, whereas dispersion forces are not. This gives graphite many of its unique properties.

Properties of graphite

The properties of graphite are dictated by its structure. These are summarised in Table 3.

TABLE 3 The properties of graphite

Property	Why does it have this property?
Slippery	The covalent layers can move and slip across each other because of dispersion forces between them. This is like a pack of cards: each card is strong, but the cards slip and move over each other.
Soft	The dispersion forces between the layers make graphite soft.
High melting point	The strength of the covalent bonds within the layers means that it takes a huge amount of energy to break them. At about 3600°C, graphite sublimes to a gas rather than melts.
High thermal conductivity	Heat can pass through the layers without disrupting the covalent bonds.
Conducts electricity	The delocalised electrons are free to move and can carry a charge throughout the layers.

Applications of graphite

You probably use graphite every day. Pencils contain graphite. This was one of the first uses of graphite and how it got its name, from the Greek work *graphene*, which means 'writing'. Because of the dispersion forces between the layers, graphite can slip off the pencil tip and adhere to the paper.

You might have used golf, fishing, tennis or cycling equipment that has graphite fibre. Graphite can be added to rubber and polymers, which are then woven into fibres and used to enhance the strength of the composite materials. You might also know this as carbon fibre.

Graphite is also used as a lubricant because it is soft and slippery. The layers of graphite slide over one another to reduce the friction between the moving parts of a machine. Locksmiths also use graphite as a lubricant for small intricate spaces.

Because graphite conducts electricity, it is also often used as the electrodes in dry cells or electrolysis. You will learn more about how graphite is used for these purposes in Unit 3, and you will probably use carbon electrodes in many practicals.

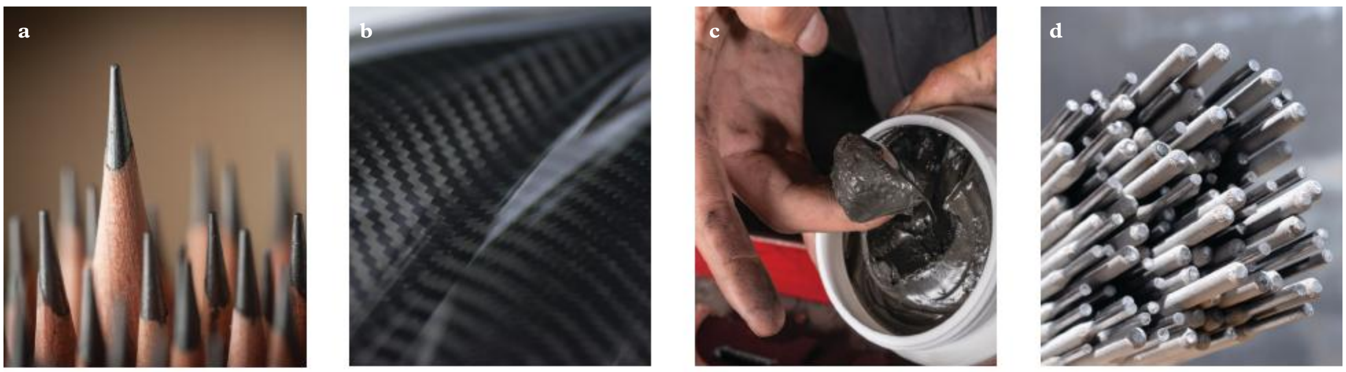


FIGURE 7 Graphite is used in **a** pencils, **b** carbon fibres, **c** lubricants and **d** welding electrode rods.

3.6 CHECK YOUR LEARNING

Describe and explain

1 Explain why carbon can form so many compounds.

2 Explain why the melting points of diamond and graphite are so high.

Apply, analyse and compare

3 Compare the following properties of diamond and graphite, and explain how the difference in structure dictates the property.

- **a** Hardness
- **b** Ability to conduct electricity
- **c** Thermal conductivity

Design and discuss

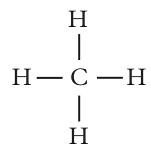
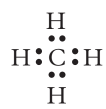
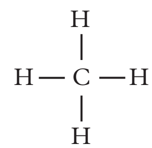
4 Discuss the properties that explain why:

- **a** diamond is used in electronics as a thermal conductor
- **b** graphite is used as a lubricant.

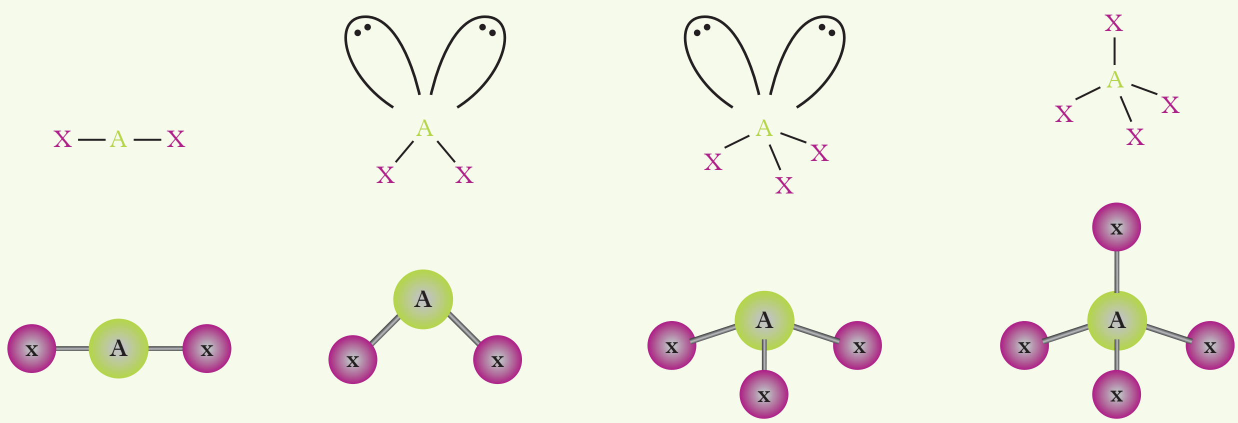
CHAPTER 3 Review

Chapter summary

3.1
- Non-metals form covalent bonds by sharing electrons.
- Covalent bonds can be single, double or triple.
- The structure of molecules can be represented by:

Molecular formula	Structural formula	Lewis (electron dot) structure	Valence structure
CH_4			

3.2
- The shape of covalent molecules is determined by the electron pairs.
- VSEPR theory states that electron pairs, both bonded and non-bonding, will be as far apart as possible around the central atom.
- The shapes of common covalent molecules:



Linear Bent (V-shaped) Pyramidal Tetrahedral

3.3
- A difference in electronegativity between the atoms in a bond makes the bond polar.
- In general, symmetrical molecules are non-polar and asymmetrical molecules are polar.
- The larger the difference in electronegativity between the atoms in a bond, the greater the polarity of the bond.

3.4
- Covalent bonds are an intramolecular force within a molecule.
- Hydrogen bonding, dipole–dipole attractions and dispersion forces are intermolecular forces between molecules.
- Strength of intermolecular forces: hydrogen bonding > dipole–dipole attraction > dispersion force.
- Dispersion forces are caused by temporary dipoles.
- Dipole–dipole attractions are formed from permanent dipoles.
- Hydrogen bonding is a type of dipole–dipole attraction that only happens when hydrogen is bonded to a fluorine, oxygen or nitrogen atom with lone-pair electrons.

3.5

- The stronger the intermolecular forces, the higher the boiling and melting point of a substance.
- The larger the molecule, the stronger the dispersion forces are between the molecules.
- Covalent molecules generally do not conduct electricity because all the electrons are fixed within the covalent bonds or atoms structure.

3.6

- Diamond and graphite are both allotropes of carbon.
- Diamond has a covalent network lattice.
- Graphite is a covalent layer lattice.
- The properties of diamond and graphite are dictated by their chemical structures.

Property	Diamond	Graphite
Hardness	Very hard	Soft and slippery
Melting point	High	High
Thermal conductivity	High	High
Electrical conductivity	Cannot conduct	Can conduct

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept: ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• represent covalent molecules using Lewis (electron dot) structures, structural formulas and molecular formulas, including for hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene	☐	☐	☐	Go back to Topic 3.1
• explain how the shape of molecules is determined by the repulsion of electron pairs according to VSEPR theory, including: linear, bent, pyramidal and tetrahedral	☐	☐	☐	Go back to Topic 3.2
• identify whether a covalent molecule is polar or non-polar from its molecular shape	☐	☐	☐	Go back to Topic 3.3
• describe the different types of intermolecular forces, including dispersion forces, dipole–dipole attraction and hydrogen bonding	☐	☐	☐	Go back to Topic 3.4
• describe the relative strengths of intramolecular bonding and forces	☐	☐	☐	Go back to Topic 3.4
• explain how the structure of covalent molecules affects their physical properties, including: melting and boiling points, and non-conduction of electricity	☐	☐	☐	Go back to Topic 3.5
• explain how the structure and bonding of diamond and graphite affects their physical properties, including: heat conductivity, electrical conductivity, and hardness	☐	☐	☐	Go back to Topic 3.6

Revision questions

Multiple choice

1 Two elements, J and K, have the following electronic configurations:

J: 2,8,4 and K: 2,8,6

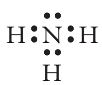
Identify the molecular formula for the compound formed between J and K.

A J_2K
B JK_2
C JK
D J_2K_3

2 Identify the correct electron dot diagram from the following options.

A 

B 

C 

D 

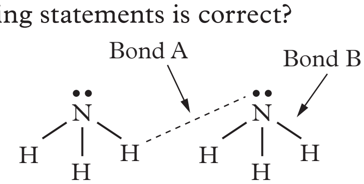
3 The structure of CO_2 is:

A linear.
B V-shaped.
C tetrahedral.
D pyramidal.

4 Which of the following molecular compounds has the weakest intermolecular forces?

A NH_3
B H_2
C O_2
D HCl

5 Consider the diagram. Which of the following statements is correct?



A Bond A is a dipole–dipole attraction and bond B is a weak dispersion force.
B Bond A is a weak dispersion force and bond B is a hydrogen bond.
C Bond A is a covalent bond and bond B is a dipole–dipole attraction.
D Bond A is a hydrogen bond and bond B is a covalent bond.

6 Which of the following is a property of a covalent molecular compound?

A Relatively low melting point
B Malleable and ductile
C High melting point
D Conducts electricity when molten but not when solid

7 A water molecule is polar because:

A the molecule is linear with no dipoles.
B hydrogen and oxygen have the same electronegativity.
C hydrogen has a lower electronegativity than oxygen and the water molecule is non-linear.
D hydrogen has a higher electronegativity than oxygen and the water molecule is non-linear.

8 The structure of CH_2Cl_2 is:

A linear.
B V-shaped.
C tetrahedral.
D pyramidal.

9 Which is the correct order of molecules, from highest to lowest boiling point?

A CH_3Cl, HCl, CH_3OH, H_2, O_2
B H_2, O_2, HCl, CH_3Cl, CH_3OH
C CH_3OH, HCl, CH_3Cl, H_2, O_2
D CH_3OH, CH_3Cl, HCl, O_2, H_2

10 A molecule is a gas at room temperature and liquefies at –196°C. What type of intermolecular forces are present between molecules of this gas?

A No bonding
B Strong covalent bonding
C Weak bonding due to dispersion forces
D Weak bonding due to dipole–dipole attractions

Short answer

Describe and explain

11 Describe the three different intermolecular forces that can be present between covalent molecules.

12 Explain the role of non-bonding electrons in the shapes of molecules.

13 Draw the Lewis (electron dot) structures for:

- **a** HBr
- **b** CH_3Cl
- **c** CS_2
- **d** F_2O

14 Explain why more energy is required to break a double covalent bond than a single covalent bond.

15 Explain how differences in electronegativity leads to intermolecular forces between molecules.

16 Graphite and naphthalene both have covalent bonding. Explain why graphite can conduct electricity whereas naphthalene cannot.

FIGURE 1 Naphthalene is used in moth balls.

17 Describe the role of diamond as an electrical insulator.

18 Explain how molecular shape can influence the boiling point of covalent compounds.

Apply, analyse and compare

19 Determine the shape of:

- **a** CO_2
- **b** Cl_2
- **c** CH_2Cl_2
- **d** HF
- **e** PCl_3
- **f** N_2
- **g** HCN
- **h** NH_2Cl

20 For each molecule in Question 19, determine the:

- **a** polarity
- **b** strongest intermolecular force present between them.

21 Consider the following pairs of covalently bonded atoms.

i N–H **ii** C–H **iii** O–H
iv H–H **v** Cl–O

For each pair of covalently bonded atoms:

- **a** identify which atom in the pair has the highest electronegativity
- **b** determine the partial charges of each atom in the bond.

Determine which of the bonds i–v has the:

- **c** highest polarity
- **d** lowest polarity.

22 Consider the following covalent molecules:

N_2, O_2, HCN, H_2S, HCl, C_2H_4, NH_3

Identify any molecule that:

- **a** has a double bond
- **b** has a triple bond
- **c** has a V-shaped structure
- **d** is non-polar
- **e** has dipole–dipole attractions
- **f** has the highest boiling point.

23 A student claims that diatomic molecules can only have dispersion forces present between molecules.

Evaluate the statement to determine whether it is correct. Justify your reasoning with examples.

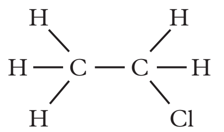
24 Arrange the following molecules in order from highest to lowest boiling points.

a Propanone ($CH_3C{=}OCH_3$), ethanol (CH_3CH_2OH), propane ($CH_3CH_2CH_3$)

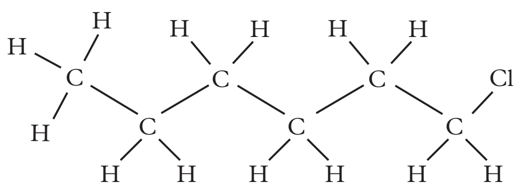
b Br_2, Cl_2, I_2

c H_2, Cl_2, CO_2, N_2

25 The structures of chloroethane and chlorohexane are shown. Compare the boiling points by discussing the intermolecular forces that are present.



Chloroethane
Boiling point 12.3°C



Chlorohexane
Boiling point 135°C

26 Compare these three molecules that can form hydrogen bonds: HF, H_2O and NH_3.

a Draw the molecular structure of each; show lone pair electrons.

b Identify the molecule that has the highest polarity.

27 Compare the following pairs of molecules and explain the difference in boiling points, with reference to intermolecular forces.

FIGURE 2 Boiling point is a physical property.

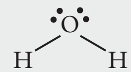
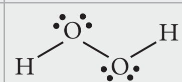
a $CHCl_3$ (61°C) and $CHBr_3$ (150°C)

b HF (20°C) and HCl (–85°C)

c Br_2 (59°C) and ICl (97°C)

28 Compare the structure (shape and intermolecular forces) and physical properties (including boiling point) of H_2, HCl and CCl_4.

29 The boiling point and structures of water and hydrogen peroxide are shown in the table.

Molecule	Structure	Boiling point (°C)
Water (H_2O)		100
Hydrogen peroxide (H_2O_2)		150.2

a Draw the Lewis structure for each molecule.

b Identify the intramolecular forces within each molecule.

c Identify the intermolecular forces between each molecule.

30 Compare each pair of molecules by drawing their structures:

i Nitrogen trifluoride (NF_3) and phosphorus trifluoride (PF_3)

ii Boron trihydride (BH_3) and ammonia (NH_3)

iii Chloromethane (CH_3Cl) and dichloromethane (CH_2Cl_2)

iv Hydrogen (H_2) and oxygen (O_2)

v Chlorine (Cl_2) and phosphorus trichloride (PCl_3)

vi Hydrogen fluoride (HF) and hydrogen chloride (HCl)

vii Methane (CH_4) and methanol (CH_3OH)

Determine which of each pair:

a is the most polar

b has the highest boiling point.

31 Compare temporary dipoles and permanent dipoles and contrast the type of intermolecular force that results from each.

Design and discuss

32 Discuss how the hydrogen bonding between water molecules allows water beetles to walk on water.

FIGURE 3 Hydrogen bonding in water allows water beetles to walk on water.

33 Discuss why diamond and graphite have much higher melting points than other covalent molecules.

FIGURE 4 Diamond has a very high melting point.

34 The table shows the boiling points of different covalent molecules.

Hydrogen halogen molecule	Boiling point (°C)	Halogen molecule	Boiling point (°C)
H–F	19.5	F–F	–188
H–Cl	–85.05	Cl–Cl	–34.6
H–Br	–66	Br–Br	58.8
H–I	–35.36	I–I	184.4

Evaluate the following statements.

a I_2 has a higher boiling point than HI.

b HF has a higher boiling point than F_2.

c HF has the highest boiling point of all the hydrogen halogen molecules.

d There is a trend in the boiling points of the halogens.

35 Discuss why the boiling point of water is 100°C, but temperatures of more than 1000°C are required to separate the hydrogen and oxygen atoms in water.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 4

Reactions of metals

KEY KNOWLEDGE

→ the common properties of metals (lustre, malleability, ductility, melting point, heat conductivity and electrical conductivity) with reference to the nature of metallic bonding and the existence of metallic crystals

→ experimental determination of a reactivity series of metals based on their relative ability to undergo oxidation with water, acids and oxygen

→ metal recycling as an example of a circular economy where metal is mined, refined, made into a product, used, disposed of via recycling and then reprocessed as the same original product or repurposed as a new product

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FIGURE 1 Bismuth is a post-transition metal that forms a complex silvery white crystal lattice. Oxidised bismuth crystals can display an iridescent tinge.

GROUNDWORK

In Chapter 4, you will learn about the bonding and properties of metals. You will also learn about how metals react with water, acids and oxygen and how metals can be recycled.

This chapter will build on concepts you have already learnt in Year 9 or 10 Chemistry. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

4A What are some common metallic properties?

4A Groundwork resource
Metallic properties

4B What happens to the charge of a neutral atom after it loses an electron?

4B Groundwork resource
Ions

4C Why is sodium such a reactive metal?

4C Groundwork resource
Reactivity

4D Why are metals considered to be useful materials?

4D Groundwork resource
Application of metals

PRACTICALS

4.2 PRACTICAL: CONTROLLED EXPERIMENT	How can we determine the reactivity series for metals?	Page 500
4.3 PRACTICAL: FIELDWORK	How is metal recovered from scrap metal?	pro

4.1

Properties of metals

KEY IDEAS

In this topic, you will learn that:

+ a metallic bond is a strong electrostatic attraction between positive cations and delocalised electrons
+ the strength of the metallic bond is responsible for a metal's uses, and its physical and chemical properties
+ common metallic properties include lustre, high melting point, high density, malleability, ductility, heat conductivity and electrical conductivity.

Of the 118 elements in the periodic table, 95 are recognised as metals. Many metallic elements have properties in common because of their structure and the unique way that they bond. The properties of metals have led to their frequent use by society for thousands of years. In this topic, we will look at the nature of metallic bonding, the existence of metallic crystals and common metallic properties.

Metallic bond model

Because of their low ionisation energies, metal atoms tend to easily lose their valence (outer shell) electrons. When atoms of a metal are compacted together, each atom's valence electrons become attracted to the positive centre of nearby metal atoms. Metallic bonds are the attraction force between these electrons and surrounding nuclei. These bonds are responsible for holding the metal together.

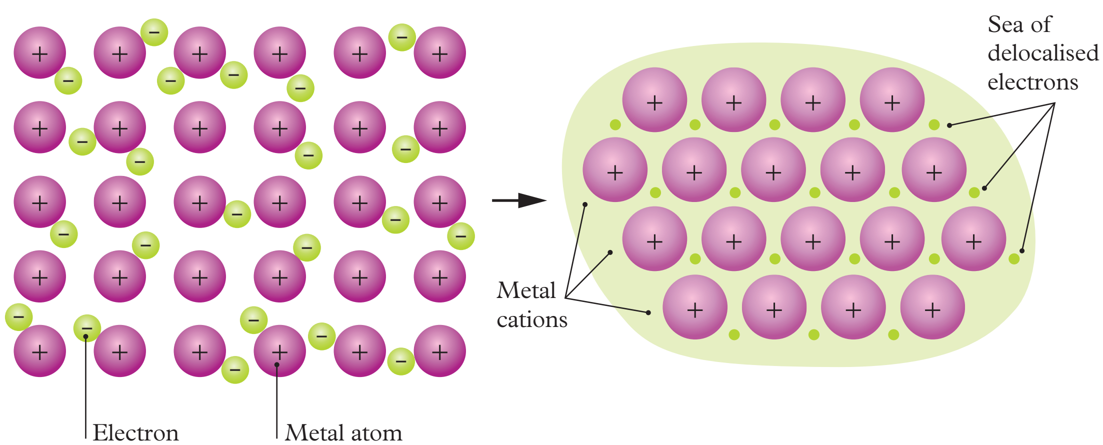
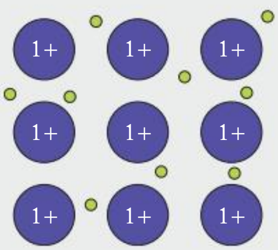
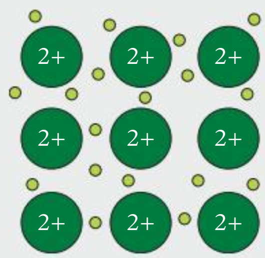
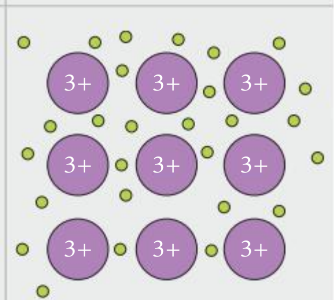


FIGURE 1 Metallic bonds are a result of attractive forces between delocalised electrons and positive metal cations.

The attraction between valence electrons and surrounding positive nuclei outweighs the repulsion between nearby valence electrons. This allows the valence electrons of metal atoms to flow from atom to atom in the metal. Since electrons are not confined to their original atom, these free-moving electrons are called delocalised electrons. Solid metals are sometimes described as having a 'sea' of delocalised electrons, which move at random (see Figure 1).

When valence electrons of a metal atom delocalise, the atom is left with more protons than electrons and becomes a positive metal cation. The charge of the cation depends on how many valence electrons have been delocalised from each atom. This is shown in Table 1.

TABLE 1 A summary of metallic bonding in potassium (K), magnesium (Mg) and aluminium (Al)

	K	Mg	Al
Number of valence electrons that delocalise per atom	1	2	3
Charge of cation when electrons are delocalised	1+	2+	3+
Visual example of metallic bonding of metallic ions			

Forces of attraction between the metal cations and the electrons moving between them are much stronger than the repulsion forces between surrounding metal cations. The delocalised electrons act like Super Glue and hold the metal cations together.

Metallic bond strength

Metallic bond strength is determined by the number of protons, number of delocalised electrons and ionic radii of the cation. The more protons there are, the stronger the interaction between cations and delocalised electrons and the stronger the metallic bond. The more delocalised electrons there are, the more the sea of electrons pulls on protons in the nucleus of nearby cations and the stronger the metallic bond. The smaller the ionic radius, the more closely packed atoms can be and the stronger the metallic bond.

Limitations of metallic bond model

As you will see later in this topic, the metallic bond model can be used to explain many common properties of metals. However, the model is limited in its ability to explain the:

- range of melting temperatures and densities of different metals
- differences in electrical conductivity between metals
- ferromagnetic nature of some metals
- relative densities of metals.

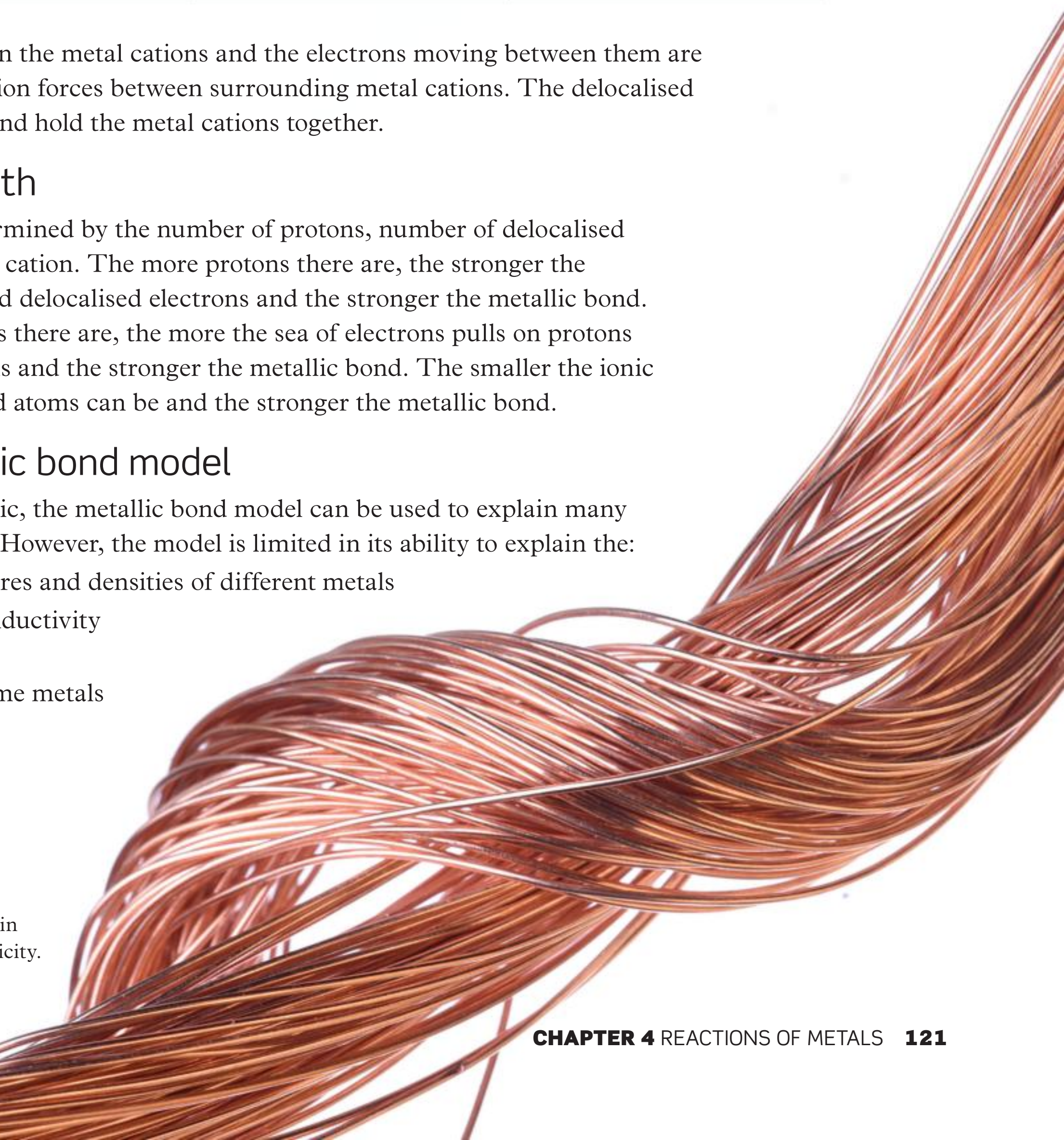


FIGURE 2 The delocalised electrons in copper wire allow it to conduct electricity.

Metallic crystals

Metal solids consist of tightly packed solid crystals. The three-dimensional arrangement of atoms can be seen by viewing a crystal through an electron microscope.

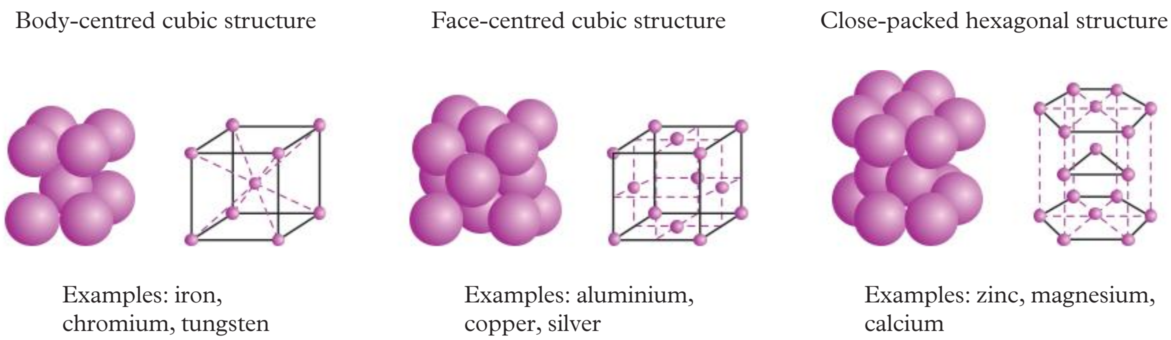


FIGURE 3 Common crystal arrangements of metals

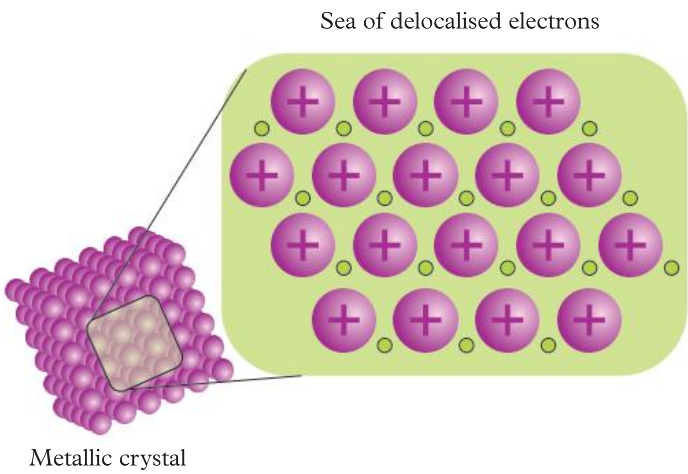


FIGURE 4 Crystal grains are lattices of cations surrounded by delocalised electrons.

Not all the atoms in a piece of metal are arranged in a regular way. Metal solids are made up of large numbers of small crystals, but the arrangement of individual crystals is random. Any piece of metal is made up of several **crystal grains**, which are regions of perfect regularity (see Figure 4). Each individual crystal grain is a lattice of cations surrounded by delocalised electrons.

Grain boundaries are the point where one crystal meets another. At grain boundaries. the regular lattice is disrupted, and atoms become misaligned (Figure 5). The way a metal behaves depends on the size of these crystals and the way that they are arranged.

crystal grain
an organised lattice of metal cations

grain boundary
the point at which different crystal grains meet in a metal

Study tip

In metals with smaller grain boundaries, the cations move over short distances, have many dislocations and are less malleable than in metals with larger grain boundaries.

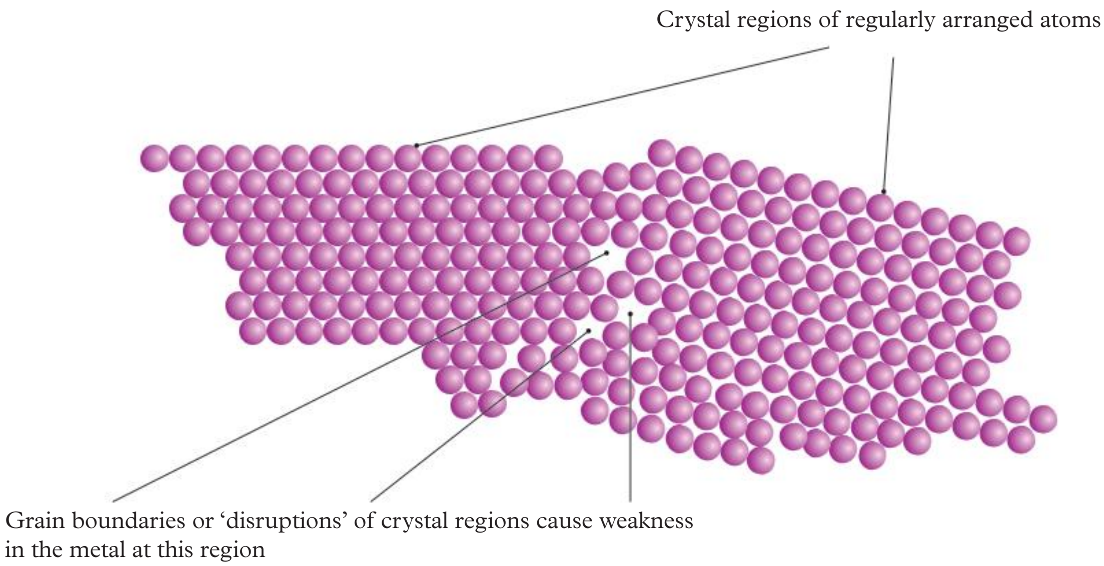


FIGURE 5 Metal crystals at grain boundaries

Properties and uses of metals

The metallic bond model and lattice structure of metals can be used to explain the properties and uses of most metals. Common properties among metals include:

- lustre
- high melting points
- high density and hardness
- malleability and ductility
- heat conductivity
- electrical conductivity.

Lustre

Most metals have a lustrous or 'shiny' appearance due to delocalised electrons within the metallic lattice. When exposed to light, these electrons vibrate and reflect light back, resulting in a shiny appearance (see Figure 6). This is why metals are often used in jewellery and in reflective surfaces such as mirrors and sculptures.

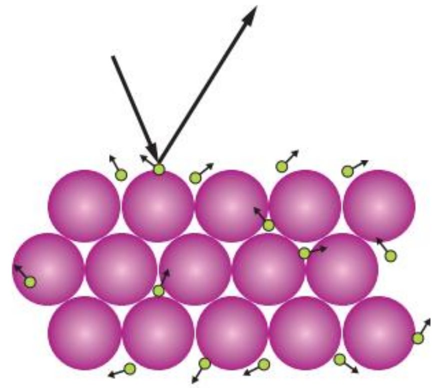


FIGURE 6 Delocalised electrons reflect light.

High melting points

Metallic bonds are very strong because of the force of attraction between the metal cations and the delocalised electrons. They have high melting points because a large amount of energy is needed to break these bonds (see Figure 7).

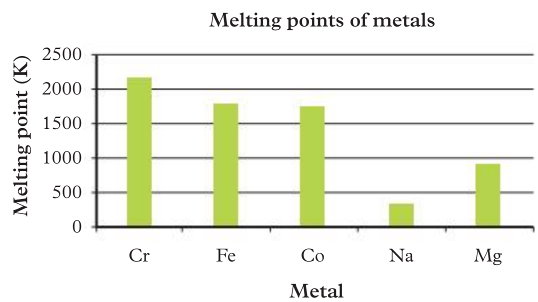


FIGURE 7 Metals have high melting points: the melting points of chromium (Cr), iron (Fe) cobalt (Co), sodium (Na) and magnesium (Mg).

Density and hardness

Metals are hard and usually have high densities at room temperature (with mercury as an exception). Metallic lattices are very tightly packed because of the strong electrostatic attraction between metal cations and the delocalised electrons. This close packing of lattices is why metals have relatively high densities. The **hardness**, high densities and high melting points of metals are useful in construction materials for buildings, bridges, transportation vehicles, machinery and appliances.

Malleability and ductility

Metals are **malleable**, meaning they can be bent or pressed. They can also be drawn into wires, which means they are **ductile**. These properties are due to delocalised electrons, which allow metal cations to slide over each other when a force is applied without breaking metallic bonds. The strong electrostatic attractive forces between cations and delocalised electrons keep metallic bonds in place (see Figure 8).

hardness
the ability to withstand deformation and scratching

malleable
can be shaped or hammered without breaking

ductile
can be drawn into wire

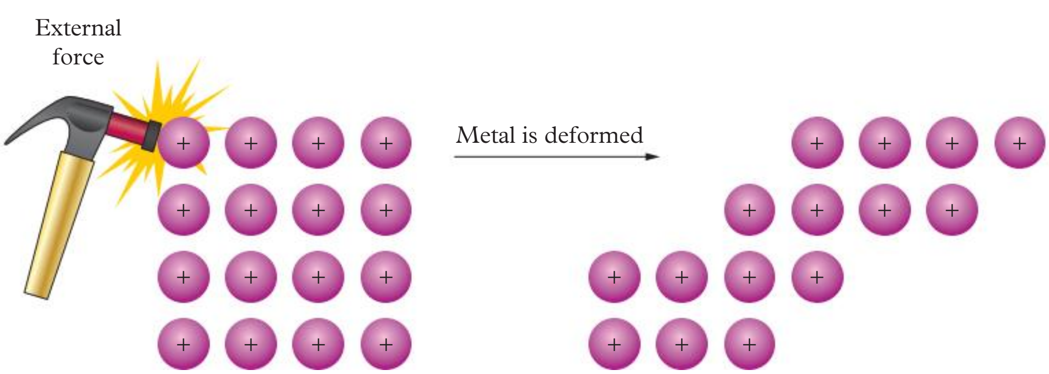


FIGURE 8 The delocalised electrons allow metal cations to shift when a force is applied.

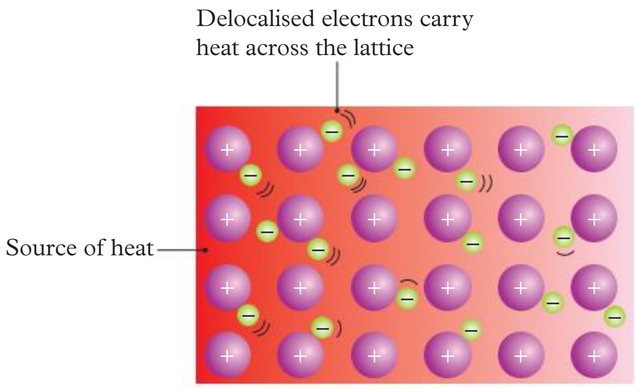


FIGURE 9 Delocalised can transfer heat energy throughout the lattice.

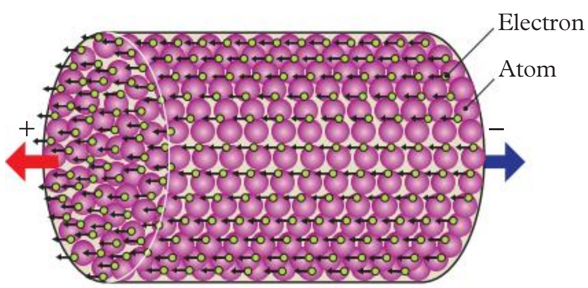


FIGURE 10 Delocalised electrons can carry electrical charge.

Heat and electrical conductivity

Metals are good conductors of heat and electricity because of the delocalised electrons within the metallic lattice. These electrons can vibrate and transfer kinetic energy (heat) throughout the lattice. Delocalised electrons can also send an electrical charge throughout the lattice. When a metal is connected into a circuit, all electrons move towards the positive terminal. This is why metals are often used as heating elements in appliances and electrical wires.

s-, d- and p-block metals

The properties of the metals in groups 1 and 2 of the periodic table (s-block) vary. The alkali metals (group 1) are soft whereas some of the alkaline metals (group 2) are hard. Apart from magnesium and beryllium, s-block metals are too reactive for any structural use, except as minor components of alloys with lead.

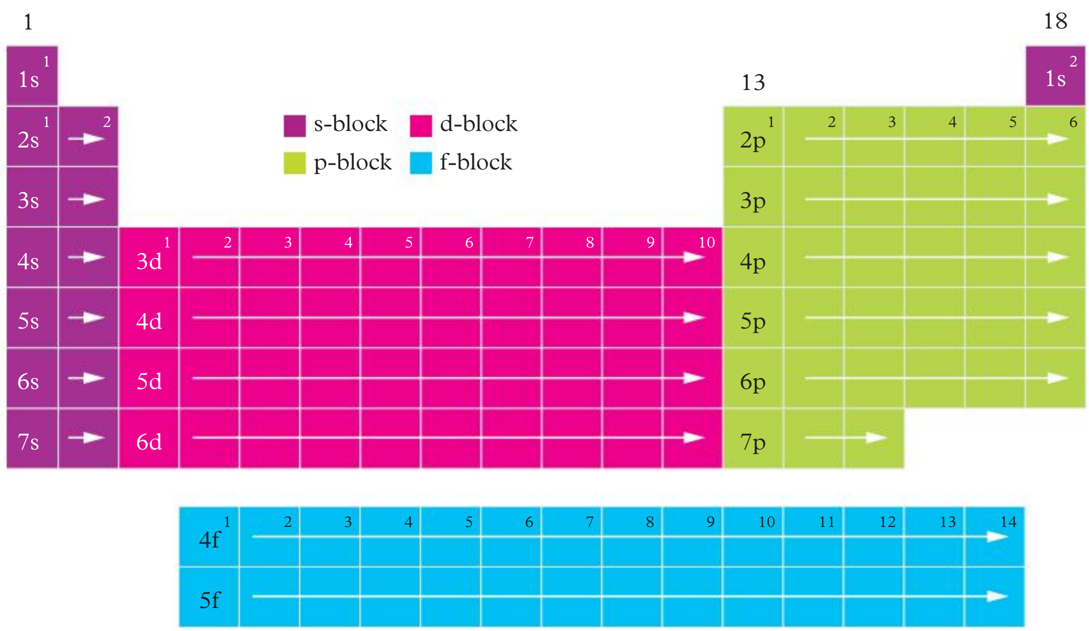


FIGURE 11 The blocks of the periodic table

Transition metals are located in the d-block (groups 3–12) of the periodic table. These metals form metallic ions with different charges, as a result of their partially filled d-orbitals. Their d-orbitals also enable them to form different coloured compounds. In general, the transition metals are denser, stronger, harder and have higher melting points (excluding mercury).

The metals in p-block (groups 13–16) are ductile and malleable but are mechanically weaker and usually have lower melting points than the transition metals. This is because p-block metals do not form ions with different charges. Because of their proximity to the semi-metals and non-metals, p-block metals are also not as reactive as the other metals mentioned above.

4.1 SKILL DRILL

Presenting melting point data

Key science skill: Analyse and evaluate data and investigation methods

Table 2 shows the melting points of various metals.

TABLE 2 The melting points of lead (Pb), nickel (Ni), gold (Au), copper (Cu) and tungsten (W)

Metal	Melting point (°C)
Pb	163
Ni	1452
Au	1063
Cu	1083
W	3399

Practise your skills

1 Graph the data presented in Table 2 (remember to include a title and labelled axis).

2 Suggest why the melting point of lead is much lower than that of the other metals.

3 Suggest why the melting point of tungsten is much higher than that of the other metals.

Need help analysing data? See Topic 1.7 (page 19).

4.1 CHECK YOUR LEARNING

Describe and explain

1 Define ‘metallic bonding’.

2 Write the electron subshell configuration of titanium and identify the charge that titanium metal atoms have when their valence electrons delocalise.

3 Explain the following in terms of the metallic bonding model.

- **a** You almost burn your hand on the end of a metal spoon that has been sitting in your hot cup of coffee for a few minutes.
- **b** When you touch a piece of wood on a cool day, it feels warm, but if you touch a piece of metal, it feels cold.
- **c** When you crush a steel can for recycling by stamping on it, it collapses down but does not fall apart or shatter.
- **d** It is dangerous to touch a person who has been electrocuted with anything metallic if they are still in contact with the live wires.

Apply, analyse and compare

4 Goldsmiths create gold ornaments either by casting (pouring molten gold into moulds) or by beating/hammering gold sheet into shape.

- **a** What properties of gold are being used in these two processes?
- **b** What information does this give us about gold particles?

5 Use your understanding of how proton number, delocalised electrons and ionic radius influence metallic bond strength to contrast the strengths of metallic bonds in lithium and calcium.

Design and discuss

6 Research the three metals that you currently make the most use of in your everyday life. Discuss your chosen metals’ unique properties that allow them to be used in these ways.

7 Draw a labelled diagram that shows metallic bonding taking place in solid calcium. Include information about the charge of the cations and the number of electrons delocalised.

4.2 Reactivity series of metals

KEY IDEAS

In this topic, you will learn that:

+ the chemical reactivity of metals depends on their position in the periodic table and how easily their valence electrons can be removed
+ a metal reactivity series can be determined empirically by their reactions with water, acid and oxygen.

Reactivity of metals

The chemical reactivity of metals depends on their relative positions in the periodic table and how easily their valence electrons can be removed. A summary of periodic table trends is given in Table 1. For specific details on these trends revisit Chapter 2.

TABLE 1 A summary of periodic table trends

Trend	Moving down a group	From left to right across a period
Metallic character	Increases	Decreases
Reducing strength of metals (ability to lose electrons)	Increases	Decreases
First ionisation energy (amount of energy in its gaseous state to lose an electron)	Decreases	Increases
Electronegativity (ability of a nucleus to attract electrons towards itself)	Decreases	Increases
Reactivity	Increases	Decreases

FIGURE 1 The reaction of iron with nitric acid, producing nitrogen dioxide gas

Understanding trends in the periodic table allows us to make predictions about the reactivity of metals. The reactivity of metals increases down a group and decreases from left to right across the table. Metals in groups 1 and 2 of the periodic table are very reactive and readily lose their valence electrons to form metallic ions in a chemical reaction. The reactivity of transition metals is generally lower than that of group 1 and 2 metals and varies greatly because of the different positions of transition metals. Metals in groups 13–16 are the least reactive.

Reactions of metals

reactivity series an analytical arrangement of metals from lowest to highest reactivity

A **reactivity series** is a ranking of substances from most to least reactive. Scientists have collected and used experimental data on how different metals react with water, acids and oxygen to determine the metal reactivity series. A reactivity series of common metals is shown in Figure 2.

4.2
PRACTICAL: CONTROLLED EXPERIMENT

How can we determine the reactivity series for metals?
Go to page **500.**

Most reactive	Metal	Reaction with water	Reaction with acids	Reaction with oxygen
	Potassium	Violent reaction with cold water	Explosive reaction with dilute acids	Reacts with oxygen at ordinary temperatures to form oxides
	Sodium Lithium Calcium	Strong reaction with hot water/steam Moderate reaction with cold water	Strong but less vigorous reaction with dilute acids	
	Magnesium	Slow reaction with cold water, strong reaction with hot water and steam	Reacts vigorously with dilute acids	Reacts with oxygen on heating to form oxides (apart from Al, which reacts at ordinary temperatures)
	Aluminium Zinc Iron	Moderate reaction with water and steam	Reacts moderately with dilute acids	
	Tin	Does not react with water		
	Lead		Reacts with concentrated acids	
	Copper Silver Gold Platinum		Does not react with dilute acids	Does not react with oxygen even with strong heat
Least reactive				

FIGURE 2 Reactions with water, acids and oxygen can determine a metal reactivity series.

Reactions with water

Given their high reactivity, it is no surprise that group 1 metals react readily with water. Group 2 metals are less reactive in water than group 1 metals. The transition metals range from reacting slowly with water to being unreactive, depending on the specific metal.

When metals do react water at room temperature, the products formed are hydrogen gas and a solution of a metal hydroxide, or hydrogen gas and a (partially soluble) metal oxide. This can be represented by the following equations.

$$\text{metal (s)} + \text{water (l)} \rightarrow \text{hydrogen (g)} + \text{metal hydroxide (aq)}$$
$$\text{metal (s)} + \text{water (l)} \rightarrow \text{hydrogen (g)} + \text{metal oxide (aq)}$$

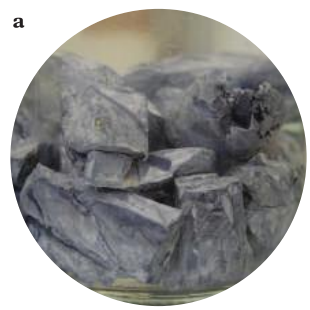
For example, when solid strontium reacts with water, it produces hydrogen gas and strontium hydroxide solution. This can be written as:

$$\text{strontium (s)} + \text{water (l)} \rightarrow \text{hydrogen (g)} + \text{strontium hydroxide (aq)}$$
or
$$Sr(s) + 2H_2O(l) \rightarrow H_2(g) + Sr(OH)_2(aq)$$

When solid aluminium reacts with water, it produces hydrogen gas and aqueous aluminium oxide. This can be written as:

$$\text{aluminium (s)} + \text{water (l)} \rightarrow \text{hydrogen (g)} + \text{aluminium oxide (aq)}$$
or
$$2Al(s) + 3H_2O(l) \rightarrow 3H_2(g) + Al_2O_3(aq)$$

a



b

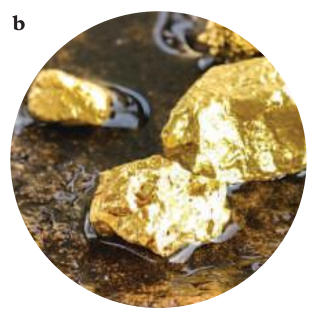


FIGURE 3 **a** The group 1 metal potassium is so reactive that it needs to be stored in mineral oil. **b** The transition metal gold is highly unreactive.

Bubbles forming on the metal surface indicates the slow formation of hydrogen gas. Addition of an acid–base indicator called phenolphthalein can help detect the formation of a metal hydroxide because it will turn pink in the presence of hydroxide ions (a base).

Reactions with acids

When metals react with an acid solution at room temperature, heat energy is released and metals dissolve in solution. The products formed are hydrogen gas and a soluble metal salt. This can be represented by the following equation:

$$\text{metal (s)} + \text{acid (aq)} \rightarrow \text{hydrogen (g)} + \text{metal salt (aq)}$$

For example, when solid zinc reacts with nitric acid, it produces hydrogen gas and soluble zinc nitrate. This can be written as:

$$\text{zinc (s)} + \text{nitric acid (aq)} \rightarrow \text{hydrogen (g)} + \text{zinc nitrate (aq)}$$

or

$$Zn(s) + 2HNO_3(aq) \rightarrow H_2(g) + Zn(NO_3)_2(aq)$$

The presence of hydrogen gas can be confirmed by holding a lit splint (or match) over the reaction vessel, which produces an audible 'pop'.

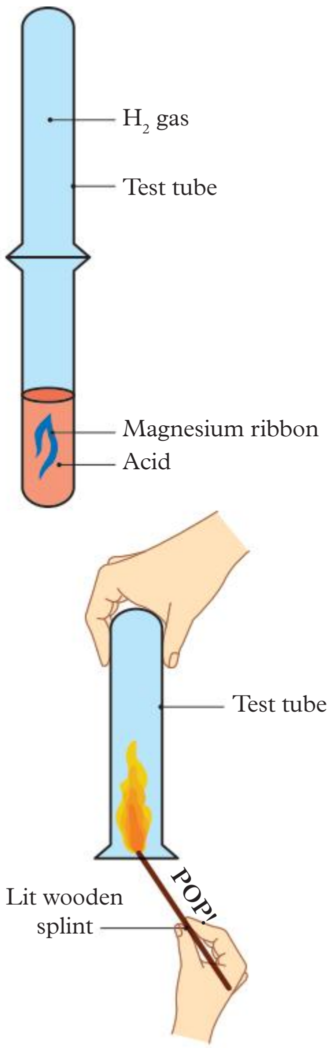


FIGURE 4 Hydrogen gas can be confirmed by holding a lit splint near a reaction vessel. If a 'pop' sound is produced, then hydrogen gas is present.

Reactions with oxygen

When metals react with oxygen gas, a solid metal oxide is formed. This can be represented by the following equation:

$$\text{metal (s)} + \text{oxygen (g)} \rightarrow \text{metal oxide (s)}$$

For example, solid magnesium reacts with oxygen gas to produce magnesium oxide. This can be represented as:

$$\text{magnesium (s)} + \text{oxygen (g)} \rightarrow \text{magnesium oxide (s)}$$

or

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

All group 1 metals react rapidly with oxygen. Group 2 metals also readily react with oxygen but not quite as quickly as group 1 metals. Transition metals are much slower to react with oxygen than group 1 and 2 metals. For example, the formation of rust (iron oxide), when the transition metal iron is exposed to oxygen and water, takes a long time.

FIGURE 5 Exposure to oxygen and water over a long time will cause an iron nail to form iron oxide (rust).

4.2 Challenge
Find me in your obook pro

See if you can explain the advantage of using alloys in Challenge 4.2.

4.2 WORKED EXAMPLE

BALANCING CHEMICAL EQUATIONS

Construct the balanced chemical equation for the reaction of solid calcium metal with an aqueous solution of hydrochloric acid.

Think	Do
Step 1: Write the correct chemical formulas for the reactants and products.	$Ca + HCl \rightarrow CaCl_2 + H_2$
Step 2: Tally atoms on both sides of the arrow.	$Ca + HCl \rightarrow CaCl_2 + H_2$ **Left side** / **Right side** Ca: 1 / 1 H: 1 / 2 Cl: 1 / 2
Step 3: Put numbers (coefficients) in front of the formulas. These numbers multiply all the atoms in the formula.	$Ca + 2HCl \rightarrow CaCl_2 + H_2$
Step 4: Check your work by re-counting all types of atoms on both sides of the arrow.	$Ca + 2HCl \rightarrow CaCl_2 + H_2$ **Left side** / **Right side** Ca: 1 / 1 H: 2 / 2 Cl: 2 / 2
Step 5: If the states of the reactants and products are known, then add the appropriate symbols.	$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

Study tip

Remember that all chemical equations must obey the law of conservation of mass. This means that the chemical bonds between the reactants and products are broken and rearranged, so that the number of atoms involved remains unchanged.

4.2 CHECK YOUR LEARNING

Describe and explain

1 Explain how the relative chemical reactivity of metals is determined.

2 Describe how the presence of hydrogen gas can be confirmed from the reactions of metals with water and acid solutions.

Apply, analyse and compare

3 Compare the relative reactivities of caesium, scandium and tin with water, acids and oxygen.

Design and discuss

4 Research the long-term effects of metal being exposed to the environment and/or atmosphere and discuss this with your classmates.

5 In laboratories, group 1 metals are usually stored in oil. Discuss why group 1 metals are stored this way and why it would be unsafe to store them otherwise.

4.3 Metal recycling

KEY IDEAS

In this topic, you will learn that:

+ transition from a linear economy towards a circular economy is a strategy for more sustainable development
+ metal recycling is an example of a circular economy.

The use of natural resources (including metal ores) was once based on a linear economy model, where resources were processed to make products that were discarded after use. A simple representation of this is shown in Figure 1.

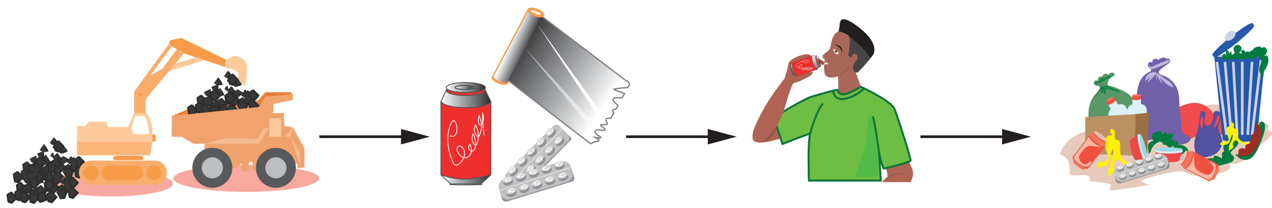


FIGURE 1 The aluminium industry was once based on a linear economy.

Today, a more sustainable way of natural resource usage involves transitioning to a circular economy model, which is a continuous cycle of use and reuse of resources. In this topic, we will examine the lifecycle of aluminium to understand how metal recycling contributes to a circular economy.

Life cycle of aluminium

non-ferrous scrap
scrap metal that does not contain iron

Aluminium, copper, lead, nickel, tin and zinc are among many base metals that are referred to in the recycling industry as **non-ferrous scrap**. These materials have an extensive range of uses and maintain their core chemical properties through repeated recycling and reprocessing. This makes metals such as aluminium infinitely recyclable and important to maintaining sustainability in terms of resource conservation.

The lifecycle of aluminium from extraction to manufacturing, distribution, disposal and recycling is an example of a metal based within a circular economy.

Extraction and refinement

Bauxite is the natural resource from which aluminium is extracted. It contains a mixture of aluminium oxides, hydroxides, clay minerals and other insoluble materials. Aluminium minerals found in bauxite include gibbsite ($Al(OH)_3$), boehmite (γ-AlO(OH)) and diaspore (α-AlO(OH)).

After bauxite ores have been mined, they are refined by the Bayer process. This process involves several stages, including milling, digestion, removal of silicon dioxide, clarification, precipitation and calcination. At the end of the process, alumina can be collected and used for aluminium production.

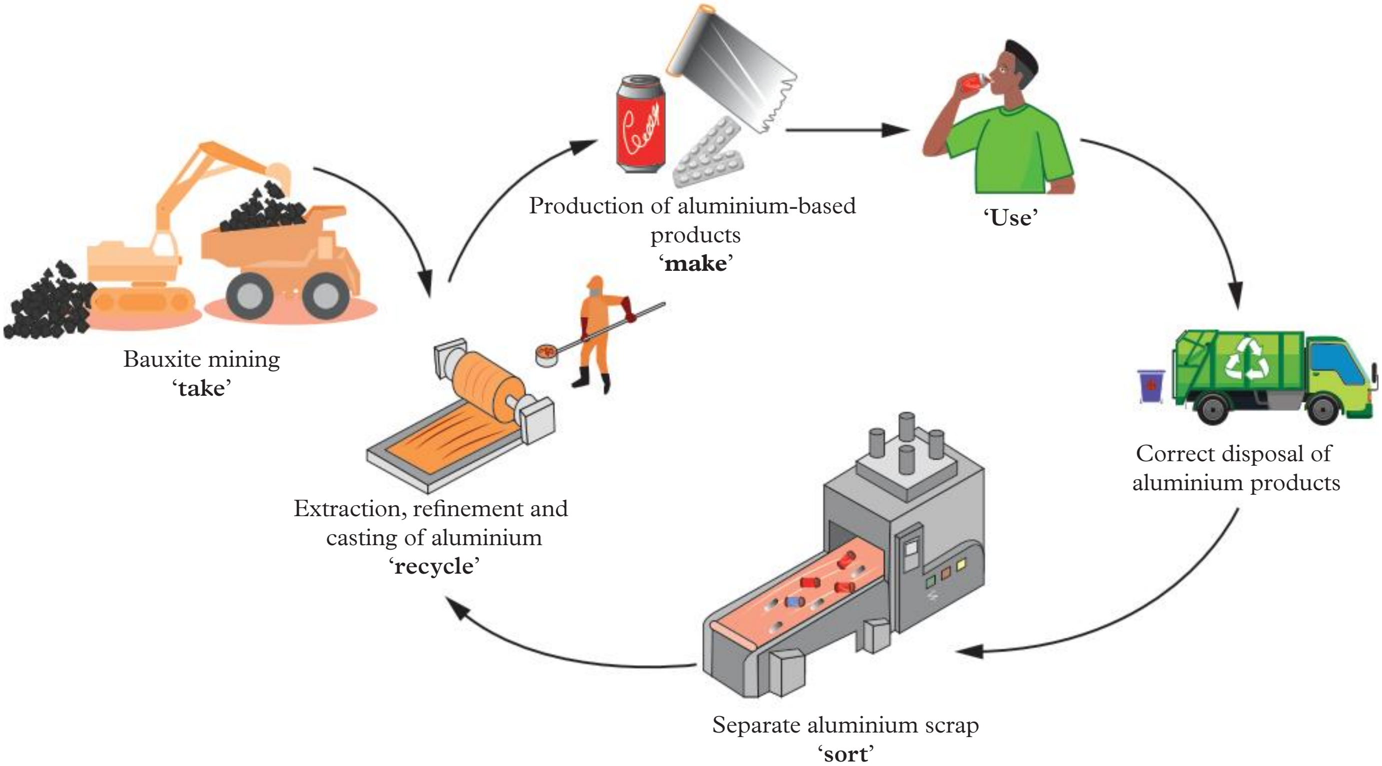


FIGURE 2 The recycling of aluminium is an example of a circular economy.

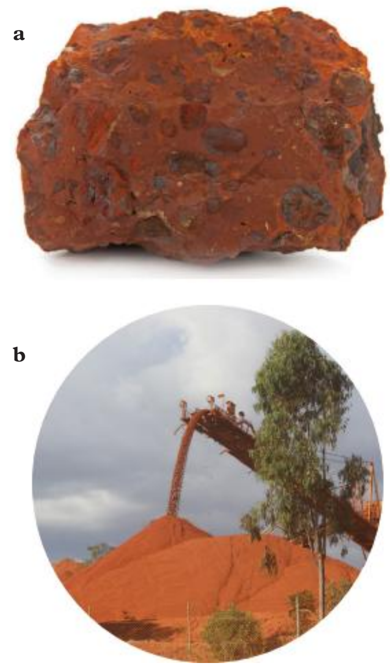


FIGURE 3 **a** Bauxite ore, and **b** mining bauxite at Weipa, Queensland

Milling and digestion

At the milling stage, the bauxite is washed and crushed to reduce particle size and increase available surface area for the digestion stage. Calcium hydroxide ($Ca(OH)_2$) and caustic soda (sodium hydroxide, NaOH) are added at the mills to make a pumpable slurry. During the digestion stage, aluminium minerals are dissolved to form a sodium aluminate saturated solution. Silicon dioxide is also removed from the slurry.

Gibbsite dissolves in the following reaction:

$$Al(OH)_3 + Na^+ + OH^- \rightarrow Al(OH)_4^- + Na^+$$

Boehmite and diaspore dissolve in the following reaction:

$$AlO(OH) + Na^+ + OH^- + H_2O \rightarrow Al(OH)_4^- + Na^+$$

Silicon dioxide is removed according to the following reaction:

$$2NaOH + SiO_2 \rightarrow Na_2SiO_3 + H_2O$$

Clarification

In the clarification stage, bauxite residue solids are separated from the sodium aluminate saturated solution. Bauxite residue is then washed so caustic soda can be recovered and reused in the digestion process.

Precipitation

In the precipitation process, the aluminium oxide is recovered by crystallisation and cooling of the saturated sodium aluminate solution. This results in the formation of small aluminium trihydroxide ($Al(OH)_3$) crystals, which grow and cluster to form larger crystals.

The formation of aluminium trihydroxide can be expressed as:

$$Al(OH)_4^- + Na^+ \rightarrow Al(OH)_3 + Na^+ + OH^-$$

The crystals formed in the precipitation stage are classified into size ranges. Large crystals are prepared for the calcination stage and fine crystals are returned to the start of the precipitation stage.

Calcination

Large $Al(OH)_3$ crystals are fed into calciners and burned at temperatures up to 1100°C to remove moisture and form alumina (Al_2O_3).

The formation of alumina can be expressed as:

$$2Al(OH)_3 \rightarrow Al_2O_3 + 3H_2O$$

A summary of the refining process is shown in Figure 4.

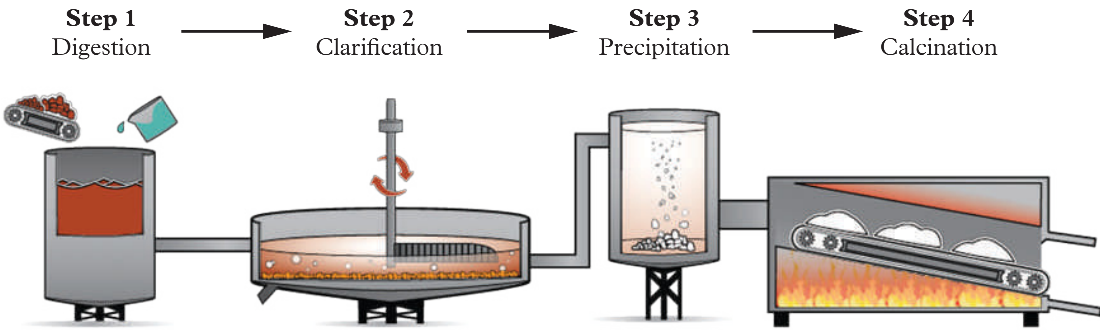


FIGURE 4 Stages of refining aluminium hydroxide

Smelting and casting

The main method of smelting aluminium used today is the Hall–Heroult process. In this process a large electric current is passed through a molten mixture of cryolite, alumina and aluminium fluoride to obtain pure, liquid aluminium metal.

Direct current is fed into a line of electrolytic cells connected in series called a 'potline'. Each pot is a large carbon-lined metal container, forming the negative electrode (cathode) in the cell.

The cell contains molten cryolite (Na_3AlF_6), maintained at a temperature of 960–980°C, in which aluminium oxide powder (Al_2O_3) is dissolved. Aluminium fluoride (AlF_3) is added to the solution to maintain optimal reaction conditions and lower the electrolyte's freezing point. Large carbon blocks are suspended in the solution and serve as the positive electrode (anode).

FIGURE 5 An aluminium smelting pot raised above a potline

FIGURE 6 A carbon anode being placed into the molten mixture

The electrical current passes from the carbon anodes via the bath to the carbon cathode cell lining. The current then passes to the anode of the next pot in series. The reactivity of aluminium means that significant energy (in the form of electricity) is required to split aluminium oxide into aluminium and oxygen.

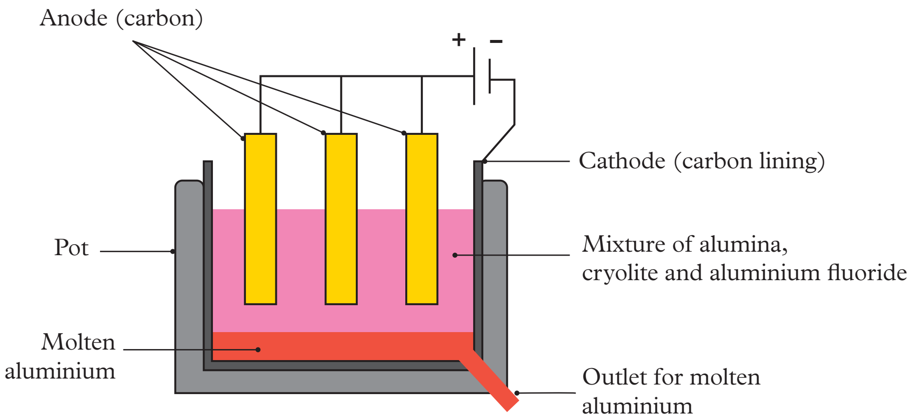


FIGURE 7 The aluminium smelting process

As the electric current passes through the solution, the dissolved aluminium oxide is split into molten aluminium (Al) and oxygen (O_2). Oxygen reacts with the carbon in the anode blocks to form carbon dioxide.

The reaction can be expressed as:

$$2Al_2O_3 + 3C \rightarrow 4Al + 3CO_2$$

Molten aluminium sinks to the bottom of the cell, while the gaseous by-products form at the top. The aluminium is drained from the pot in a process called tapping and transported to dedicated casting operations. Aluminium can then be alloyed and cast into ingots, billets, T-bars, rolled coils for cans and other bulk products.

Semi-fabrication and products

Molten aluminium (when cooled) is classified as a 'soft' metal. The durability and strength of this aluminium can be increased by the addition of (or alloying of aluminium with) other elements (silicon, manganese, copper, magnesium or zinc) before it is processed into different products. These products include:

- castings and cast products for the automotive industry
- extrusions to specific shapes for tools and other items used in construction, transportation and machines
- rolled products, including flat sheet, coiled sheet, plate and foil, of different thicknesses and tempers for the packaging, food and beverage, construction and transportation industries
- milled products, which are non-standard semi-fabricated products that have specific shapes or forms made to order. They include sheets, plates, foil, extruded products, drawing stock, wire, pigments and powder, forgings and impact extrusion products.

Recycling and repurposing

Approximately 75% of produced aluminium is currently still in use and can be continuously recycled. Recycling aluminium saves up to 95% of the energy required to manufacture new aluminium metal. The key steps involved in recycling and repurposing aluminium involve collecting scrap, sorting scrap, crushing, remelting, casting and repurposing.

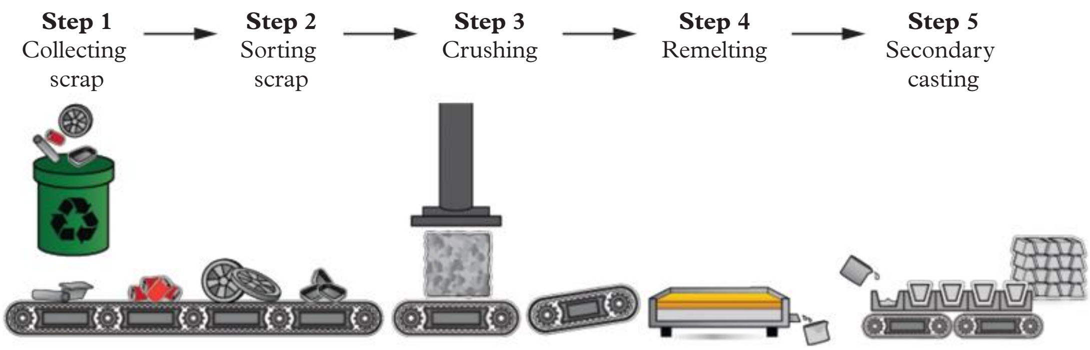


FIGURE 8 Aluminium recycling stages

Collecting scrap

new scrap
scrap metal that comes from the manufacturing of metal products

old scrap
used scrap metal that has been collected from consumers

New scrap is scrap taken from the manufacturing and/or fabrication of aluminium products, up to the point where they are sold to the final consumer. For example, off-cuts of aluminium sheet or extrusions are considered new scrap and may be safely recycled by aluminium smelters if the composition is known.

Old scrap has been used by the consumer and discarded or disposed of. For example, used beverage cans, window frames, electrical cabling and car cylinder heads are considered old scrap. Aluminium smelters are unable to safely accept this old scrap because its composition is unknown and it may be contaminated.

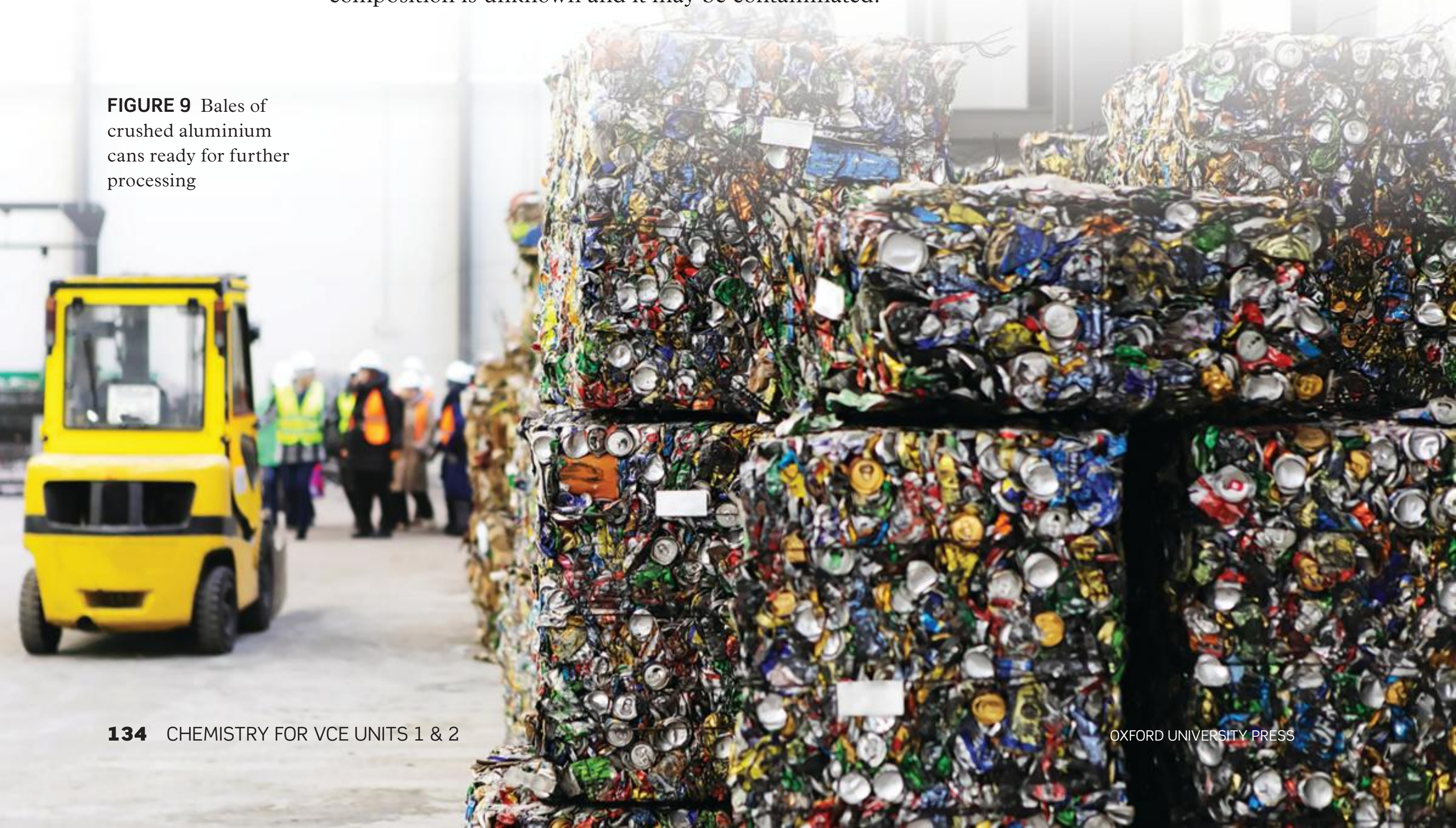


FIGURE 9 Bales of crushed aluminium cans ready for further processing

Sorting scrap

Scrap is sorted by grouping all coated (painted or lacquered) aluminium together and grouping all uncoated aluminium together. Paper, plastic and other non-aluminium recycling is removed during this process.

Crushing

Sorted aluminium scrap is crushed and compacted into large bales (see Figure 9), to reduce the cost of freight, storage and handling.

Remelting

Uncoated scrap is washed and loaded directly into a large furnace called a remelter, where it is re-heated at high temperatures into a molten form. If the scrap aluminium is coated, it is further processed through a gas-fired rotary furnace to remove any remaining coating; it is then washed and transferred to the remelter.

Casting

The molten aluminium scrap is cast at a temperature of just over 700°C to form ingots. After casting, ingots are ready for transport and repurposing.

Repurposing

The recast scrap aluminium ingots are transported or distributed to the semi-fabrication and product manufacturing centres mentioned in the previous section, ready to be remade into more useful specific consumer and industrial products, continuing the aluminium circular economy.

Explore how lithium is recycled in Australia in Real-world chemistry 4.3.

Study tip

Circular economies are not just limited to metal recycling. Other industry examples are human excrement for fertiliser and biogas, fashion, animal feed in agriculture, food waste, plastic and glass packaging.

4.3 Real-world chemistry

Find me in your obook pro

4.3 CHECK YOUR LEARNING

Describe and explain

1 Describe the key processes involved in the:
   a production of aluminium
   b recycling of aluminium.

2 Explain with the use of chemical equations the process of smelting aluminium oxide to form aluminium metal.

3 Identify whether a used soft-drink can is considered new scrap or old scrap.

4 Explain why aluminium is considered infinitely recyclable.

Apply, analyse and compare

5 Compare new scrap with old scrap in the metal recycling process.

6 Determine which type of aluminium product (castings, extrusions, milled products or rolled products) would have been used to create an:
   a aluminium-based oven tray
   b aluminium wire
   c aluminium car frame.

Design and discuss

7 A student made the following statement: ‘Aluminium is not the only material recovered for repurposing in the aluminium recycling process’. Identify whether this statement is correct or incorrect. Justify your answer by using information provided throughout this topic.

CHAPTER

4 Review

Chapter summary

4.1
- Metallic bonding occurs through the electrostatic attraction between delocalised electrons and metal cations.
- Common metallic properties such as lustre, high density, malleability, ductility, electrical conductivity and heat conductivity can be explained by the metallic bond model.
- Solid metals are made from metallic crystals with organised arrangements of atoms.

4.2
- The reactivity of different metals relates to their position on the periodic table and the periodic table trends that apply.
- A reactivity series of metals can be determined through observing metal reactions with water, acids and oxygen.
- When a metal reacts with water, the products are hydrogen gas and a solution of a metal hydroxide or partially soluble metal oxide.
- When a metal reacts with an acid, the products are hydrogen gas and a metal salt.
- When a metal reacts with oxygen, the product is a solid metal oxide.

4.3
- The recycling of metals (such as aluminium) contributes to a circular economy.

Key formulas

Metal reactions with water	metal (s) + water (l) → hydrogen (g) + metal hydroxide (aq) or metal (s) + water (l) → hydrogen (g) + metal oxide (aq)
Metal reaction with acids	metal (s) + acid (aq) → hydrogen (g) + metal salt (aq)
Metal reaction with oxygen	metal (s) + oxygen (g) → metal oxide (s)

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• define metallic bonding and metallic crystals	☐	☐	☐	Go back to Topic 4.1
• describe how the metallic bonding model can explain common metallic properties	☐	☐	☐	Go back to Topic 4.1
• explain how periodic table trends apply to the reactivity of metals	☐	☐	☐	Go back to Topic 4.2
• explain and determine how metals will react with water, acids and oxygen	☐	☐	☐	Go back to Topic 4.2
• understand how metals must undergo several processes such as mining, refinement and casting before they can be used commercially	☐	☐	☐	Go back to Topic 4.3
• explain how the recycling of metals contributes to a circular economy	☐	☐	☐	Go back to Topic 4.3

Revision questions

Multiple choice

1 The structure and bonding of solid sodium metal at room temperature can be best described as a:

A network lattice of closely packed sodium atoms, held together by strong ionic bonds.

B lattice of Na_2 molecules.

C network lattice of sodium ions, held together by a 'sea' of electrons.

D lattice of Na_2 molecules held together by weak bonds.

2 Which of the following properties is not characteristic of metals?

A Good conductors of heat and electricity

B Shiny

C Hard and brittle

D Relatively dense

3 The property of metals that enables them to conduct an electric current is:

A the outer electrons of metals are not firmly bound to the atom.

B metal atoms are better suited than non-metal atoms to be charge carriers.

C ions in the metal can move freely through the metallic lattice.

D metal atoms are not firmly bonded to each other as non-metallic atoms are.

4 The term 'malleable' is applied to a metal that:

A can be hammered into thin sheets.

B resists corrosion.

C can be drawn out to form a thin wire.

D can be used to form alloys.

5 When a force is applied to a metal:

A the layers of metal cations move relative to each other within the delocalised sea of electrons.

B metallic bonds are broken.

C electrons and ions move together in layers.

D the delocalised sea of electrons can be squashed.

6 When a metal is not conducting electricity:

A electrons are associated with specific atoms.

B electrons are free to move throughout the lattice, in all directions.

C electrons don't move.

D cations are free to move throughout the lattice, in all directions.

7 Aluminium, being strong and light, is used to produce:

A aircraft bodies.

B cooking utensils.

C foils.

D all of the above.

8 Copper wire can be bent easily without breaking. The best explanation of this is that:

A the forces between copper atoms are weak, allowing the copper atoms to easily move around.

B slight changes in the positions of copper atoms do not break the metallic bonds because they are equally strong in all directions.

C copper atoms are strongly bonded in layers, but there are only weak bonds between the layers.

D copper atoms are arranged in flat molecules that freely slide over each other, allowing the material to be bent.

9 Which of these metals is the least reactive?

A Silver

B Gold

C Platinum

D Mercury

10 Sodium metal melts at 98°C and potassium melts at 64°C. The melting point of potassium is lower than that of sodium because:

A the potassium atoms are smaller than the sodium atoms so less heat energy is required to make the potassium atoms slide over each other.

B the sodium atoms lose more valence electrons than potassium atoms do when they form the metal so the attractive forces in solid potassium are weaker.

C diatomic potassium molecules are held together in the solid by weaker dispersion forces than those between the diatomic sodium molecules.

D the potassium ions are larger than the sodium ions so the delocalised valence electrons attract the potassium ions less strongly.

Short answer

Describe and explain

11 Which properties of metals would be the most important for the following situations?

a Emergency huts for accommodation after a cyclone disaster

b The heating coil in an electric hot water system

c A telescope mirror

d An ore drill bit for grinding the face of a rock

12 Consider the following metals.

Fe, Mg, K, Au, Pb, Cs

a Place the metals in order from most reactive to least reactive.

b List the symbols of *two* metals that are likely to react vigorously with water.

c Which metal is least likely to form an oxide layer?

13 Define 'first ionisation energy'. Explain how the first ionisation energy of a metal can help determine how reactive that metal is.

14 A description of the metallic bonding model says that 'metals consist of a regular, three-dimensional crystal lattice of cations, surrounded by a sea of delocalised electrons'.
   - **a** What evidence is there for a:
     - **i** regular three-dimensional lattice?
     - **ii** sea of delocalised electrons?
   - **b** Why do metals contain metal cations?

15 In terms of the metallic bonding model, explain why metals are:
   - **a** generally hard, with high melting points
   - **b** good electrical conductors
   - **c** ductile and malleable.

16 Identify which United Nations Sustainable Development Goal(s) are addressed through the recycling of metals. Explain your answer.

Apply, analyse and compare

17 Using appropriate examples from the life cycle of aluminium (in Topic 4.3), compare a linear economy and a circular economy.

18 The melting temperatures of various metallic elements are shown in the table.

Element	Melting temperature (°C)
Lithium	180
Sodium	98
Iron	1540
Chromium	1890
Copper	1080
Zinc	420

   - **a** Suggest whether the data in the table supports the general trend of main group metals having lower melting temperatures than transition metals.
   - **b** State another property that is usually different in transition metals compared with main group metals.
   - **c** State one physical property that is exhibited by both the main group metals and the transition metals.
   - **d** Using the metallic bonding model, explain the property outlined in your choice for part c.

Design and discuss

19 Discuss how aluminium recycling is an example of a circular economy.

20 Draw a diagram that show the arrangement of copper ions and electrons in a metal lattice.

21 Design an experiment you could use to determine a reactivity series for metals.

22 Research a metal other than aluminium that can be extracted and/or recovered from discarded consumer goods. Create a flow chart of the methods used for its extraction and where it is repurposed.

23 Draw a diagram that shows how delocalised electrons allow metals to:
   - **a** conduct electricity
   - **b** conduct heat
   - **c** have a shiny appearance.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 5

Reactions of ionic compounds

KEY KNOWLEDGE

→ the common properties of ionic compounds (brittleness, hardness, melting point, difference in electrical conductivity in solid and molten liquid states) with reference to the nature of ionic bonding and crystal structure

→ deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions (NH_4^+, OH^-, NO_3^-, HCO_3^-, CO_3^{2-}, SO_4^{2-} and PO_4^{3-})

→ the formation of ionic compounds through the transfer of electrons from metals to non-metals, and the writing of ionic compound formulas, including those containing polyatomic ions and transition metal ions

→ the use of solubility tables to predict and identify precipitation reactions between ions in solution, represented by balanced full and ionic equations including the state symbols: (s), (l), (aq) and (g)

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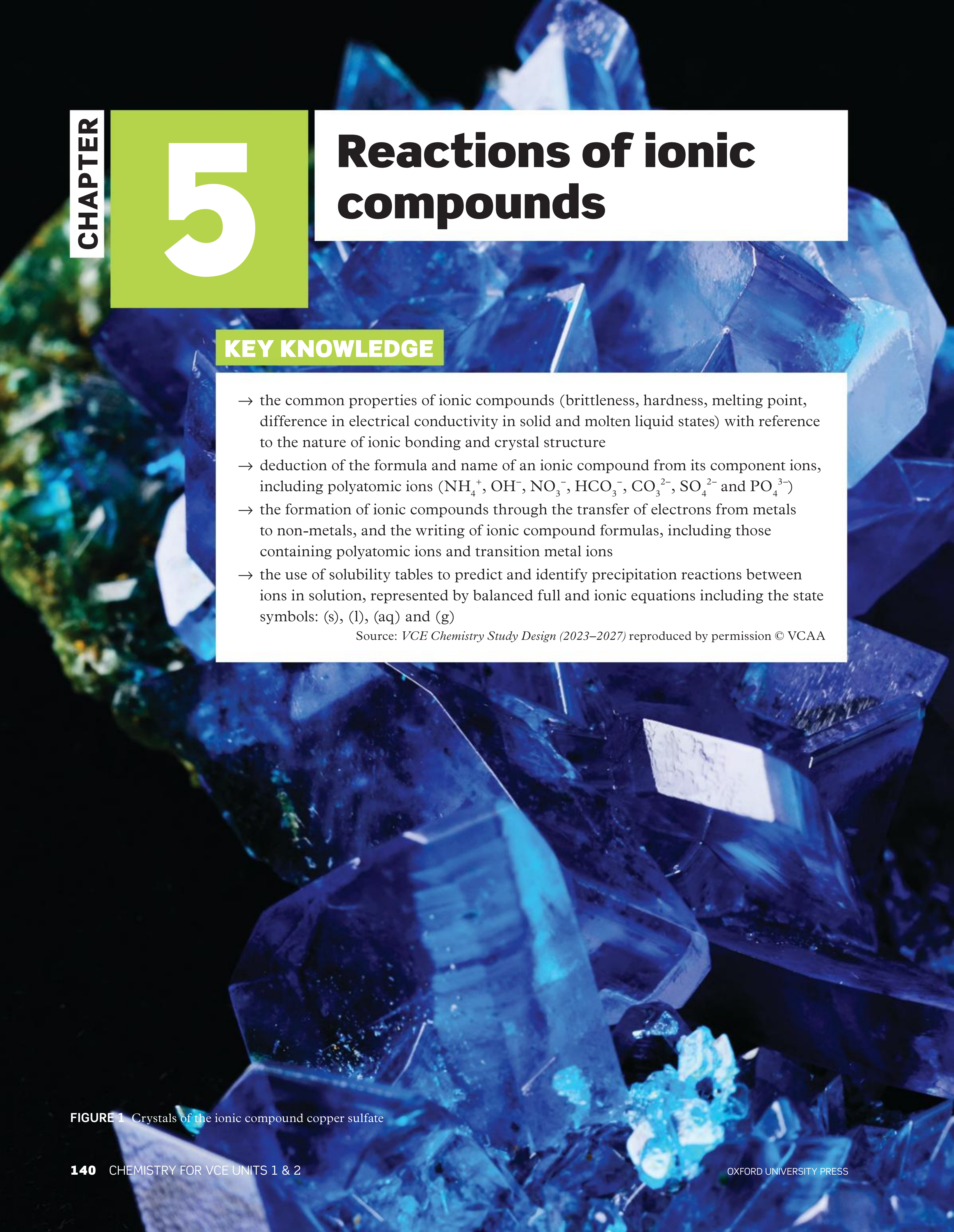
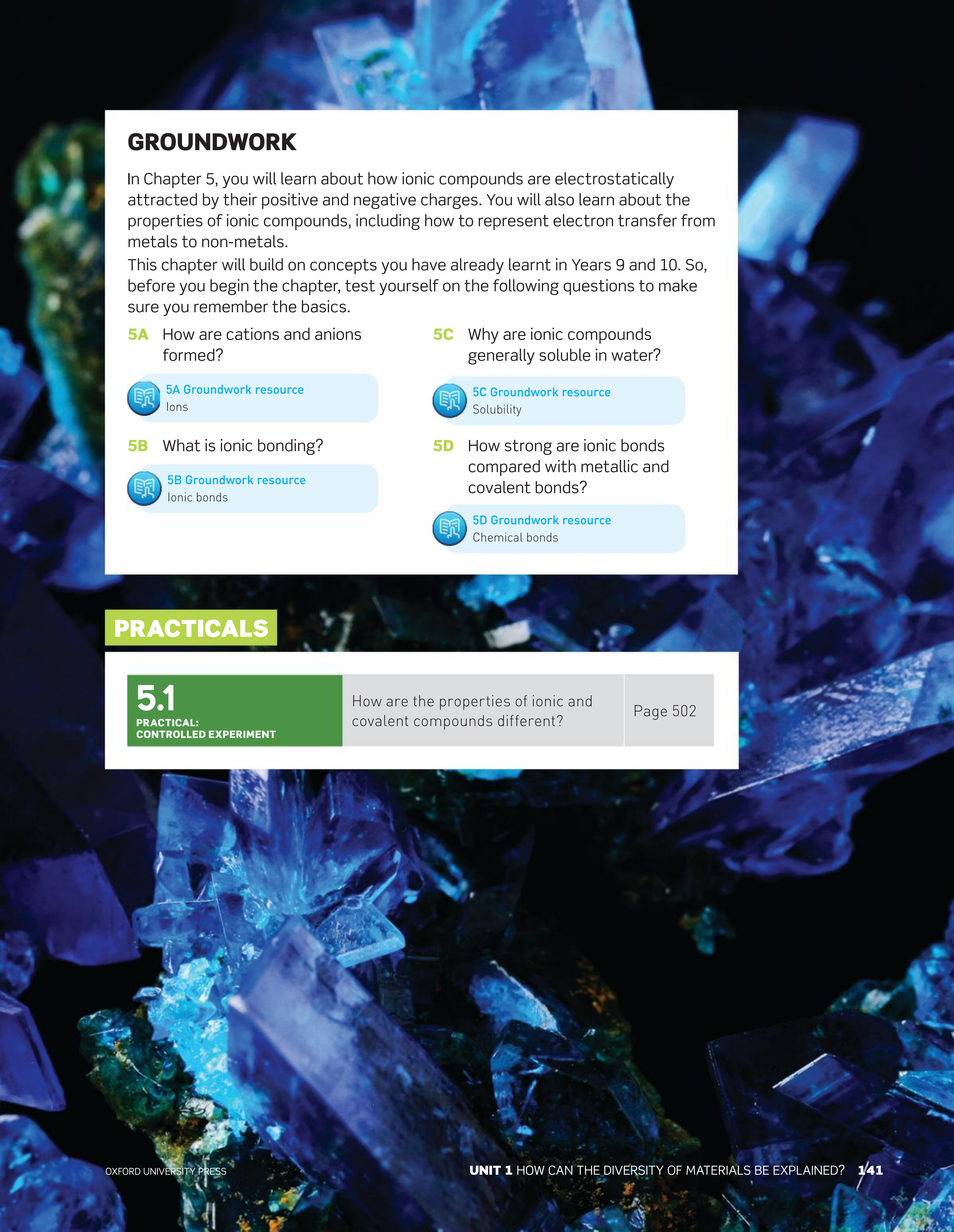


FIGURE 1 Crystals of the ionic compound copper sulfate



GROUNDWORK

In Chapter 5, you will learn about how ionic compounds are electrostatically attracted by their positive and negative charges. You will also learn about the properties of ionic compounds, including how to represent electron transfer from metals to non-metals.

This chapter will build on concepts you have already learnt in Years 9 and 10. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

5A How are cations and anions formed?

5A Groundwork resource
Ions

5B What is ionic bonding?

5B Groundwork resource
Ionic bonds

5C Why are ionic compounds generally soluble in water?

5C Groundwork resource
Solubility

5D How strong are ionic bonds compared with metallic and covalent bonds?

5D Groundwork resource
Chemical bonds

PRACTICALS

5.1 PRACTICAL: CONTROLLED EXPERIMENT	How are the properties of ionic and covalent compounds different?	Page 502

5.1 Properties of ionic compounds

KEY IDEAS

In this topic, you will learn that:

- ionic bonds are formed between positive cations and negative anions, held together by strong electrostatic forces, in a repeating, ordered, three-dimensional crystal lattice
- common properties of ionic compounds include hardness, brittleness, high melting and boiling points, being soluble in water and being able to conduct electricity in a liquid state.

Ionic bonding

Because of their electron configuration, many metal atoms are prone to losing electrons and forming positive cations. By contrast, many non-metals have nearly complete valence shells. These non-metals are more likely to accept electrons and form negative anions. An ionic bond is a chemical bond that forms as a result of the electrostatic attraction between a negative cation and a positive anion.

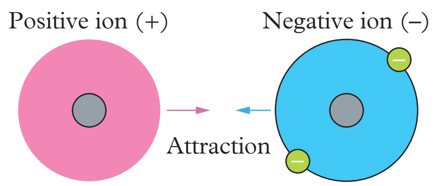


FIGURE 1 Ionic bonds form between positive and negative ions.

ionic compound
a chemical compound that is held in place by ionic bonds

Common table salt is an example of an **ionic compound**, a compound that is held together by ionic bonds. Table salt (sodium chloride or NaCl) forms through the ionic bonding of sodium cations and chloride anions.

A sodium atom has one valence electron; by losing this one electron it becomes a 1+ cation with a full valence shell. A chlorine atom has seven valence electrons in its outer shell; by accepting one electron, it becomes a 1– anion with a full valence shell. The transfer of sodium's one valence electron to chlorine results in the stable ionic compound sodium chloride.

FIGURE 2 Table salt (NaCl) is an example of an ionic compound.

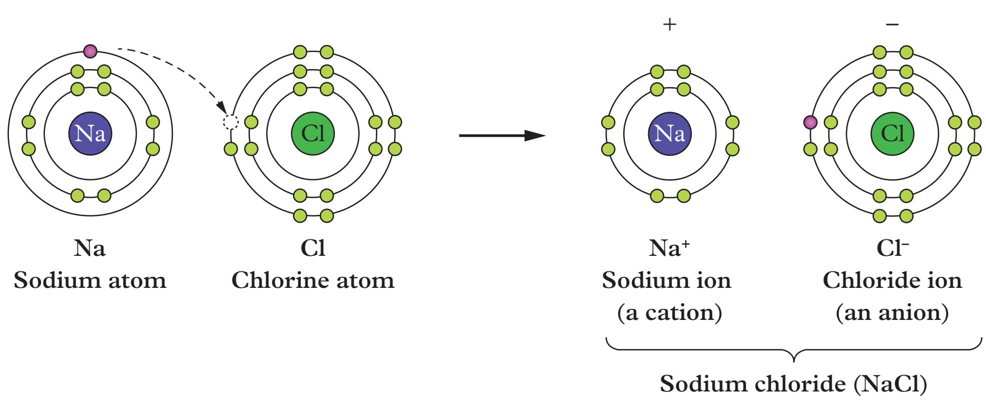


FIGURE 3 Sodium chloride forms when a sodium atom transfers one electron to a chlorine atom.

5.1
PRACTICAL:
CONTROLLED EXPERIMENT

How are the properties of ionic and covalent compounds different?
Go to page **502**.

Ionic crystal lattices

Ionic compounds form **crystal lattices** – ordered, three-dimensional arrangements of positive cations and negative anions. Ionic crystal lattices are held in place by the strong electrostatic forces (ionic bonds) that exist between the oppositely charged anions and cations.

Ions in a crystal lattice are arranged in a regular repeating pattern; they are packed together so that a positive cation is only surrounded by negative anions and a negative anion is only surrounded by positive cations. There are many ways that ionic lattices can be arranged. You can see few different organisations of ionic crystal lattices in Figure 5.

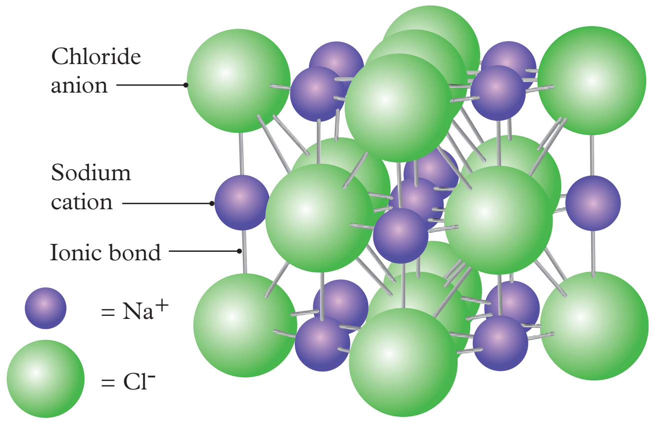


FIGURE 4 The crystal lattice arrangement of sodium chloride (NaCl), an ionic compound

Common properties of ionic compounds

Common properties among ionic compounds can be explained by the nature of ionic bonds and the arrangement of ionic crystal lattices.

Melting point

Ionic compounds usually have high melting and boiling points. This is because a large amount of energy is needed to separate and break the electrostatic attraction (ionic bonds) that exist between cations and anions in the ionic compound.

Hardness

Ionic compounds are relatively hard. This is because crystal lattice surfaces are not easily scratched because of the strong ionic bonds holding the ions together.

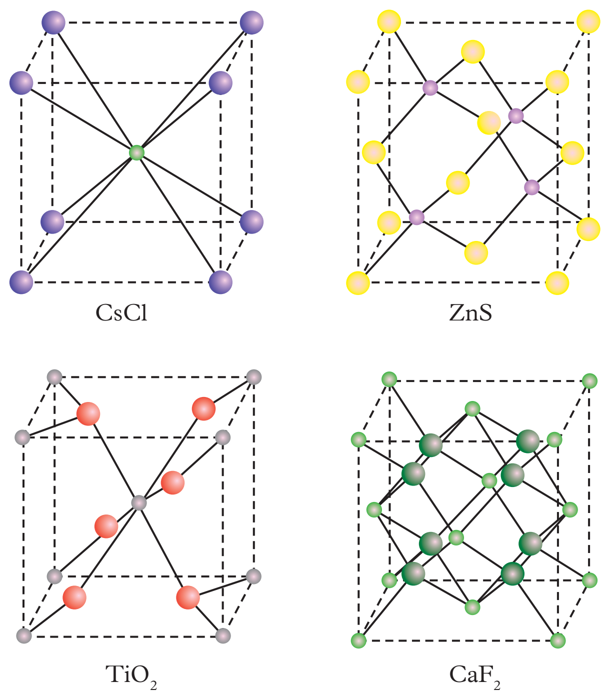


FIGURE 5 Different lattice arrangement of ionic compounds

Brittleness

Despite being relatively hard, ionic compounds are usually brittle. When struck with force, the crystal lattice structure is disrupted, which causes ions of the same charge to come close together. The repulsion between same-charged ions breaks or shatters the crystal (Figure 6).

crystal lattice
a three-dimensional arrangement of atoms or ions that repeats to make up a compound

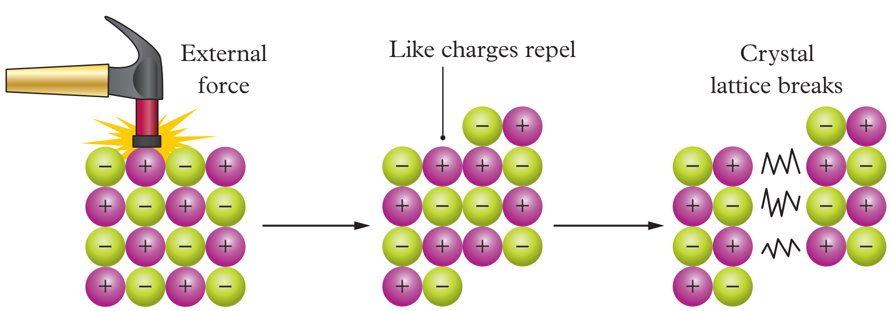


FIGURE 6 Solid ionic compounds are brittle.

Lustre

Ionic compounds are not lustrous. This is because electrons are not free to move in the lattice because they are held by strong electrostatic attraction. Because electrons in the lattice are not free moving, they do not reflect light.

electrolyte
an electrically conductive solution that contains free moving charged ions

Solubility in water

Most ionic compounds will dissolve in water to form free mobile ions. This is because water molecules can move between ions and break the ionic bonds, disrupting the rigid crystal structure. When an ionic compound dissolves in water, the solution is called an **electrolyte**.

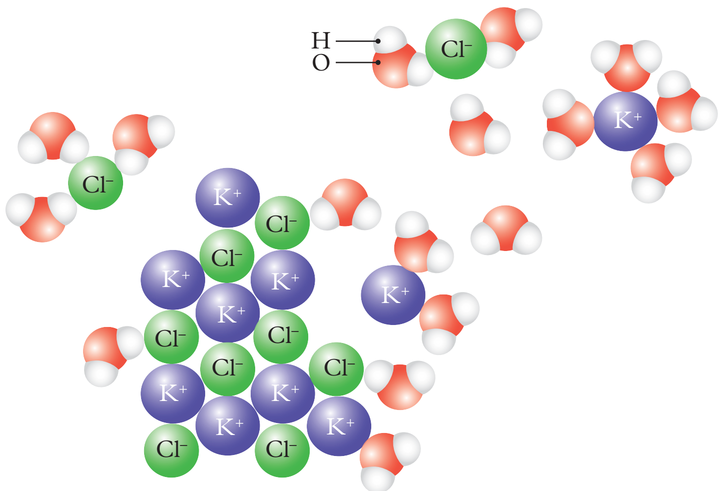


FIGURE 7 Solid potassium chloride will dissolve in water.

Electrical conductivity

Ionic compounds do not conduct electricity in the solid form. This is because ions within the lattice are unable able to move and carry an electrical charge. However, ionic compounds do conduct electricity when molten or in aqueous solution (dissolved in water).

When molten, oppositely charged ions in an ionic compound can move by sliding past one another and can conduct electricity. When dissolved in water, ions separate from the lattice and can move freely to conduct an electric current.

Study tip

Remember that ionic compounds do not conduct electricity in the solid state, only when molten or dissolved in water (aqueous solution).

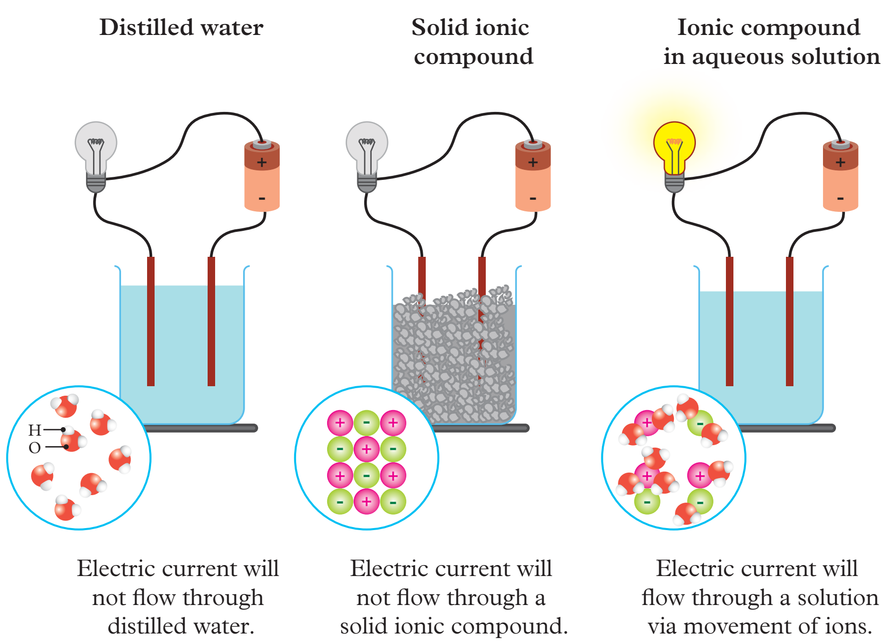


FIGURE 8 Ionic compounds conduct electricity in aqueous solution but not as a solid.

5.1 SKILL DRILL

Distinguishing between scientific and non-scientific ideas about ionic compounds

Key science skill: Construct evidence-based arguments and draw conclusions.

FIGURE 9 Students studying together

Three students are researching the properties of ionic compounds for an assignment. Each student has suggested what sort of information they should include in their presentation.

- Student 1 suggests that they should include information on how ionic compounds are lustrous because they once saw a large salt crystal that appeared to shine brightly in sunlight.
- Student 2 suggests that they should include information on how ionic compounds have high melting points because data from several peer-reviewed journal articles they have read supports this concept.
- Student 3 suggests that they should include information on how all ionic compounds can conduct electricity because in a classroom experiment, they were able to keep a circuit connected through use of an ionic compound in aqueous solution.

Practise your skills

1 Identify and explain whether each student's suggestion is based on opinion, anecdote or evidence.

2 Identify and explain whether each suggestion presents valid scientific ideas or non-scientific ideas.

Need help constructing evidence-based arguments and drawing conclusions? See Topic 1.9 (page 28).

5.1 CHECK YOUR LEARNING

Describe and explain

1 Describe how an ionic bond is formed.

2 Explain, using examples, how elements in different groups of the periodic table gain or lose specific numbers of electrons to form their respective ions.

3 Describe the structure of an ionic compound.

Apply, analyse and compare

4 For each of the ionic compounds shown in Figure 5, identify which element forms the cation and which forms the anion.

5 With reference to the structure, compare the electrical conductivity of an ionic compound in solid and liquid states.

6 Compare the appearance and malleability of ionic compounds and metal elements.

Design and discuss

7 Draw a labelled diagram that shows why ionic compounds are brittle and can break when struck by a force.

5.2 Formation of ionic compounds

KEY IDEAS

In this topic, you will learn that:

- ions are formed by the transfer of an atom's valence electrons
- ionic bonding can be represented by electron transfer diagrams.

Electron transfer diagrams

Ionic bonds often form between metals and non-metals. The ionic bond involves a transfer of valence electrons from the (lower electronegative) metal atoms to the (higher electronegative) non-metal atoms.

electron transfer diagram
a visual representation that shows the movement of electrons between atoms and/or molecules

Electron transfer can be represented visually by drawing an **electron transfer diagram**. Electron transfer diagrams feature electron shell diagrams of the atoms involved and show the movement of electrons between atoms using arrows.

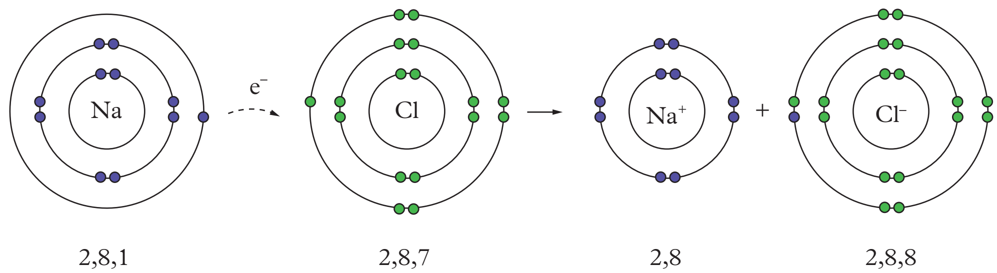


FIGURE 1 An electron transfer diagram showing the formation of NaCl

Drawing the formation of ions

If an atom has one valence electron, it will lose this one valence electron to gain a complete outer shell and become more stable. For example, a sodium atom has an electron configuration of 2,8,1. Sodium will lose its one valence electron to become more stable. It now has one less electron than it has protons and forms a cation with a 1+ charge, Na^+ (Figure 2).

If an atom has 2–4 valence electrons, to become more stable it will lose its valence electrons. For example, a magnesium atom has an electron configuration of 2,8,2. Magnesium will lose two valence shell electrons to become more stable. It now has two less electrons than protons and so forms a cation with a 2+ charge, Mg^{2+} (Figure 3).

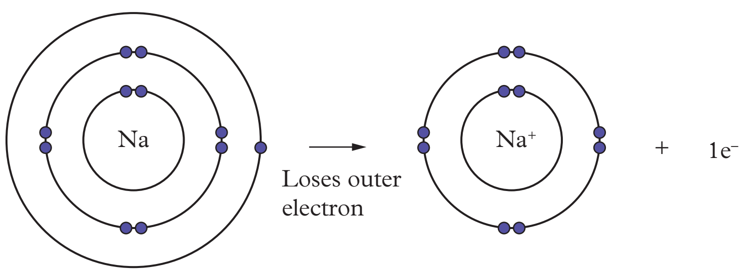


FIGURE 2 The formation of a sodium cation

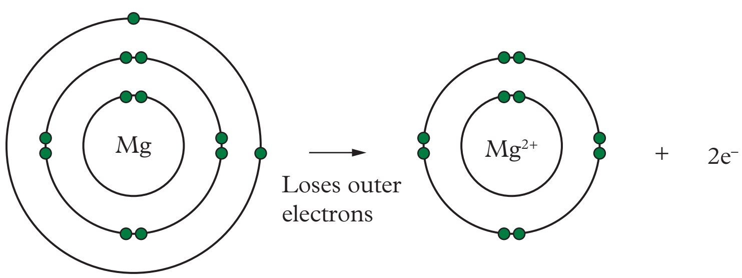


FIGURE 3 The formation of a magnesium cation

If an atom has an incomplete valence shell it can become more stable by accepting electrons and completing its valence shell. For example, a chlorine atom has an electronic configuration of 2,8,7. To become more stable, the atom accepts one electron. It now has one more electron than protons and forms an anion with a 1– charge, Cl^- (Figure 4).

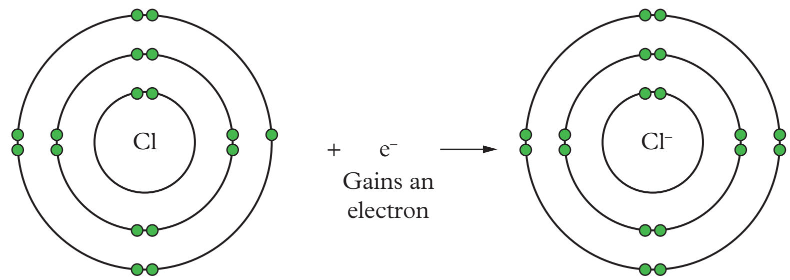


FIGURE 4 The formation of a chloride anion

An oxygen atom with electron configuration 2,6 will gain two electrons to form an anion with a 2– charge, O^{2-} (Figure 5). Note that all anions are formed from non-metal atoms, either as a single element (e.g. Cl^-) or in groups (e.g. OH^-).

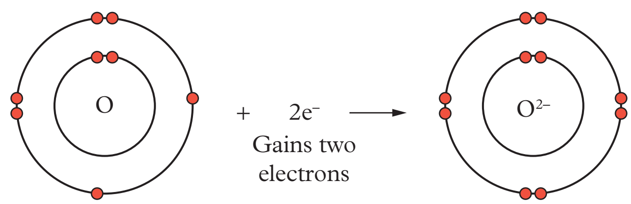


FIGURE 5 The formation of an oxide anion

Drawing the formation of ionic bonds

To represent the formation of an ionic bond between metal and non-metal elements, we draw arrows to show the movement of the electrons from the respective atoms and their charges formed. This is modelled in Figure 6 where a magnesium atom loses two electrons to become a stable magnesium ion with a 2+ charge. The two electrons removed from magnesium are accepted by an oxygen atom. The oxygen is now an oxide ion with a complete valence shell and 2– charge. The electrostatic attraction between the Mg^{2+} cation and O^{2-} anion forms an ionic bond, causing the formation of the ionic compound magnesium oxide.

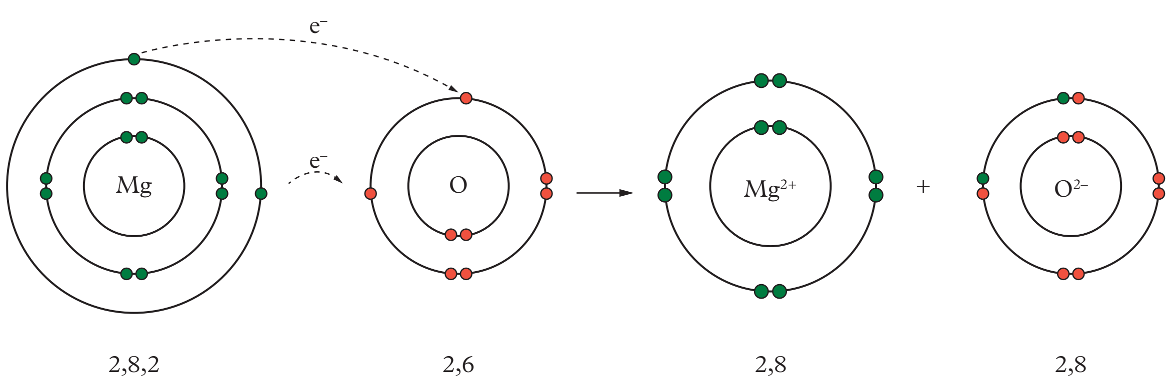


FIGURE 6 Representing ionic bond formation in magnesium oxide by an electron transfer diagram

Study tip

Remember that, when drawing electron transfer diagrams, you must have the correct ratio of atoms in the ionic compounds and balance their respective charges, to form a neutral compound.

The ratio of cations to anions in an ionic compound is not always one-to-one.

For example, the ionic compound magnesium chloride ($MgCl_2$) is formed from the attraction between one Mg^{2+} cation and two Cl^- anions. The combined 2– charge from the chloride anions balances the 2+ charge from the magnesium cation (Figure 7).

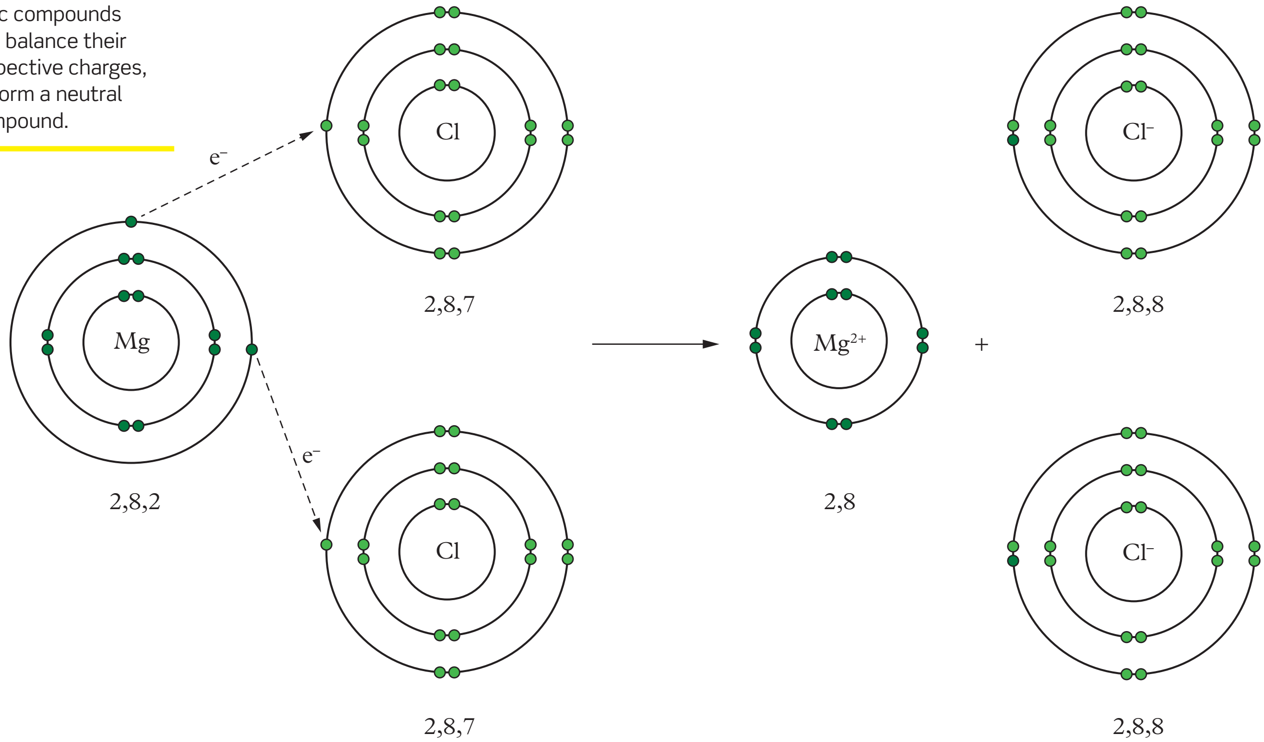


FIGURE 7 Representing ionic bond formation in magnesium chloride by an electron transfer diagram

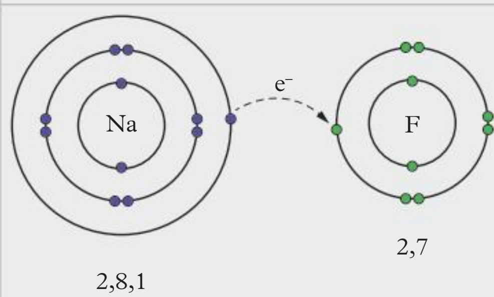
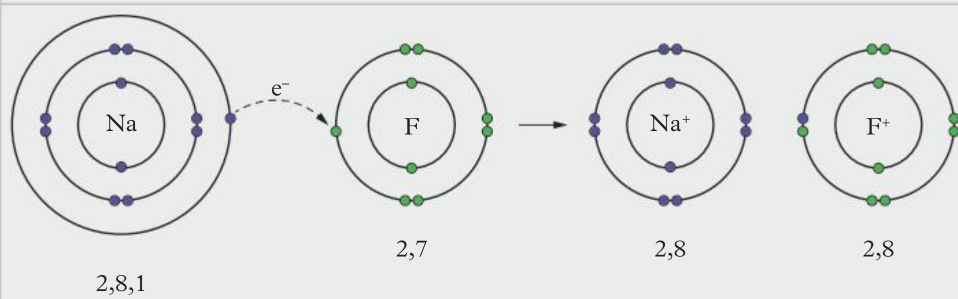
FIGURE 8 The ionic compound magnesium chloride ($MgCl_2$) is often in bath salts.

5.2 WORKED EXAMPLE

DRAWING ELECTRON TRANSFER DIAGRAMS

Draw an electron transfer diagram to represent the ionic bonding of sodium and fluorine to form sodium fluoride.

Solution

Think	Do
Step 1: Write the electron configurations of sodium and fluorine.	Sodium = 2,8,1 Fluorine = 2,7
Step 2: Determine how many electrons need to be lost or gained to make each atom stable.	Sodium needs to lose one valence electron to become stable. Fluorine needs to gain one valence electron to become stable.
Step 3: Draw the electron shell diagrams of both atoms and use arrows to represent the movement of the electron from the sodium atom to the fluorine atom.	
Step 4: Draw the ions that have formed after the electron transfer and check your work to make sure that the ionic compound formed has equal numbers of positive and negative charges.	

5.2 CHECK YOUR LEARNING

Describe and explain

1 Describe the steps involved when drawing electron transfer diagrams to represent ionic bonds.

2 Explain the importance of representing balanced charges when drawing electron transfer diagrams.

Apply, analyse and compare

3 How many valence electrons would the following elements lose or gain to obtain stability or a full outer shell of electrons?

- **a** Lead
- **b** Selenium
- **c** Barium
- **d** Antimony

4 Using examples, compare the size of the cation charges formed when the transition metals form ionic compounds with oxygen.

Design and discuss

5 Draw electron transfer diagrams to represent the ionic bonding between:

- **a** lithium and chlorine to form lithium chloride
- **b** aluminium and oxygen to form aluminium oxide
- **c** sodium and nitrogen to form sodium nitride.

5.3 Formulas and naming of ionic compounds

KEY IDEAS

In this topic, you will learn that:

+ ionic compounds can be formed from polyatomic ions
+ the formula of an ionic compound is the ratio of cations (positive ions) to anions (negative ions) when the overall charge of the compound is equal to zero
+ there are rules for writing and naming ionic compounds.

Formation of ions

An ionic compound has an equal number of positive and negative charges. This makes ionic compounds electrically neutral. The charge of each ion is determined by the positions of the elements in the periodic table and how many electrons they need to gain or lose to achieve a full outer shell. For example:

- group 1 elements lose a single electron to form a cation with a 1+ charge
- group 2 elements lose two electrons to form a 2+ cation
- aluminium, a group 13 element, loses three electrons to form a 3+ cation
- group 17 elements gain a single electron to form an anion with a 1– charge
- group 16 elements gain two electrons to form anions with a 2– charge
- group 15 elements gain three electrons to form anions with a 3– charge.

monatomic
made from one type of atom

polyatomic
made from more than one type of atom

Most cations are formed from **monatomic** (one type of atom) metal atoms. However, there are also **polyatomic** (more than one type of atom) cations, such as the ammonium ion, NH_4^+. Anions can also be either monatomic non-metal atoms such as Cl^- or polyatomic ions such as OH^-. Table 1 shows the valency of common monatomic and polyatomic ions. This table can be used to determine the formulas and charges of cations and anions that form ionic compounds.

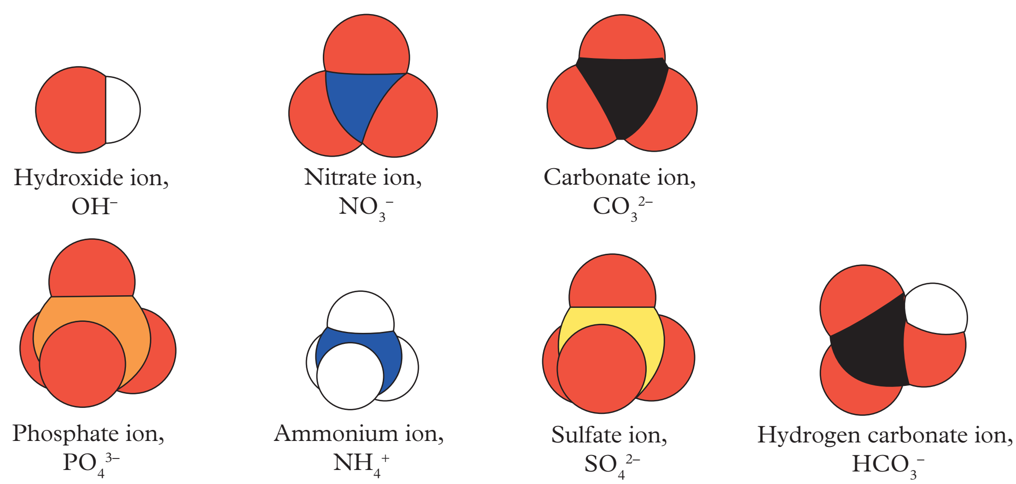


FIGURE 1 Some common polyatomic ions

TABLE 1 Valencies of common ions

Cations			
1+	2+	3+	4+
Lithium, Li^+	Magnesium, Mg^{2+}	Aluminium, Al^{3+}	Tin(IV), Sn^{4+}
Sodium, Na^+	Calcium, Ca^{2+}	Chromium(III), Cr^{3+}	Lead(IV), Pb^{4+}
Potassium, K^+	Barium, Ba^{2+}	Iron(III), Fe^{3+}	Chromium, Cr^{4+}
Silver, Ag^+	Zinc, Zn^{2+}		
Copper(I), Cu^+	Copper(II), Cu^{2+}		
Ammonium, NH_4^+	Mercury(II), Hg^{2+}		
	Iron(II), Fe^{2+}		
	Nickel(II), Ni^{2+}		
	Tin(II), Sn^{2+}		
	Lead(II), Pb^{2+}		
Anions			
1–	2–	3–	4–
Hydride, H^-	Oxide, O^{2-}	Nitride, N^{3-}	Carbide, C^{4-}
Hydroxide, OH^-	Sulfide, S^{2-}	Phosphate, PO_4^{3-}	Silicide, Si^{4-}
Hydrogen sulfide, HS^-	Sulfite, SO_3^{2-}	Phosphide, P^{3-}	
Hydrogen sulfate, HSO_4^-	Sulfate, SO_4^{2-}		
Hydrogen carbonate, HCO_3^-	Carbonate, CO_3^{2-}		
Nitrite, NO_2^-	Hydrogen phosphate, HPO_4^{2-}		
Nitrate, NO_3^-	Chromate, CrO_4^{2-}		
Acetate, CH_3COO^-	Dichromate, $Cr_2O_7^{2-}$		
Fluoride, F^-			
Chloride, Cl^-			
Bromide, Br^-			
Iodide, I^-			

Writing ionic formulas

empirical formula
the simplest whole-number ratio of atoms in a compound

When writing the **empirical formula** of an ionic compound, the cation is always written before the anion and the net charge of the compound formed must be equal to zero.

The steps for writing the formulas of ionic compounds are shown below for using aluminium oxide as an example.

1 Place the cation first with its correct charge. For example:

$$Al^{3+}$$

2 Place the anion second with its correct charge (from Table 1). For example:

$$Al^{3+}\ O^{2-}$$

3 Combine the positive and negative ions so that the number of positive charges equals the number of negative charges.

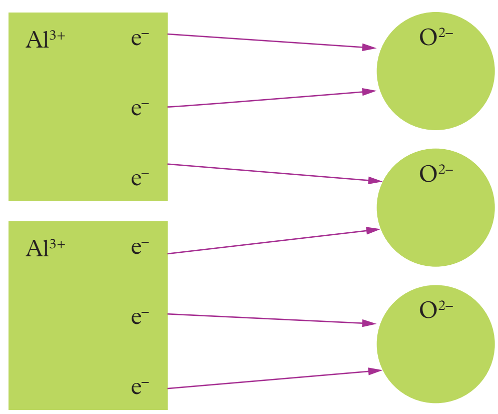


FIGURE 2 The ratio of cations to anions in an ionic compound must ensure that positive and negative charges from ions balance out.

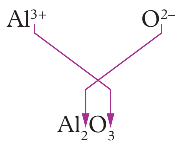


FIGURE 3 The swap method to determine ratio of cations to anions

To quickly determine the ratio of cations to anions, take the numerical value of the cation's superscript and move it to the bottom right-hand side of the anion symbol as a subscript. Then take the numerical value of the anion's superscript and make it the subscript of the cation. This is known as the swap method (Figure 3).

4 Write the number of each ion needed to achieve a neutral charge is as a subscript after the ion. For polyatomic ions, place the polyatomic ion in brackets and then put the subscript number after the brackets. The final formula is a neutral compound, so there are no charges shown in an ionic compound formula: Al_2O_3.

FIGURE 4 The ionic formula for amethyst (silicon dioxide) is SiO_2.

Naming ionic formulas

Most ionic compounds have two-word names. The first word in the name relates to the cation and the second word in the name relates to the anion.

Naming the cations

To name the cations of elements in groups 1, 2 and 13, the cation is simply the name of the element. For example, K is in group 1 so the cation K^+ is called the 'potassium ion'.

The cations of some elements, such as the transition metals (groups 3–12), can have more than one charge, so check if more than one charge is possible for the element. If more than one charge is possible, specify the charge by placing a Roman numeral representing the charge after the metal in brackets.

For example:

- iron(II) chloride contains the Fe^{2+} ion and the formula is $FeCl_2$
- iron(III) chloride contains the Fe^{3+} ion and the formula is $FeCl_3$.

Naming the anions

If an ionic compound contains only two elements (a metal and a non-metal) then:

- both names appear in the compound name
- the name of the metal always comes first, followed by the name of the non-metal whose ending changes to -ide; for example, metal oxide, metal phosphide, metal chloride, metal sulfide, metal nitride, metal bromide, metal iodide.

If the compound contains two elements (a metal and a non-metal) and oxygen, then:

- both names appear in the compound name
- the name of the metal always comes first, followed by the name of the non-metal, whose ending changes to -ate or -ite; for example, metal sulfate/sulfite, metal nitrate/nitrite, metal iodate, metal phosphate, metal chlorate, metal bromate, metal carbonate.

Naming hydrated ionic compounds

A number of ionic compounds can 'trap' water in their ionic lattice. These compounds are called **hydrates.** When heated, a hydrate will release water. For naming these compounds or writing their formula we need to include the number of water molecules.

For example, the formula for copper(II) sulfate pentahydrate is written as $CuSO_4 \cdot 5H_2O$. This means that there are five molecules of water for every formula unit of $CuSO_4$. The five water molecules are given the prefix 'penta' followed by hydrate (see Table 2). The hydrate dot or '•' between the $CuSO_4$ and $5H_2O$ indicates that water molecules are loosely bonded.

hydrate
an ionic compound that has water molecules surrounding the charged ions

TABLE 2 Prefixes of hydrated ionic compounds

Prefix	Number	Prefix	Number
mono	1	hexa	6
di	2	hepta	7
tri	3	octa	8
tetra	4	nona	9
penta	5	decca	10

Study tip

The best way to name an ionic compound is to look at the formula, identify the cation and the anion, then write their names together.

5.3 CHALLENGE

Naming ionic formulas

Use Tables 1 and 2 to write the correct names of the following ionic formulas.

1 $Cr_2(SO_4)_3$
2 $Ca_3(PO_4)_2$
3 $FeSO_4 \cdot 8H_2O$

5.3 CHECK YOUR LEARNING

Describe and explain

1 Explain the importance of using the simplest whole-number ratio of atoms and charges when writing ionic compound formulas.
2 Describe how the valency of a transition metal cation is included in the name of an ionic compound.

Apply, analyse and compare

3 Write the ionic formula for:
   a chromium(VI) oxide
   b sodium hydrogen sulfate
   c aluminium iodide.
4 Name the following ionic compounds.
   a Ca_2C
   b $(NH_4)_2SO_4$
   c Fe_2S_3
5 Contrast polyatomic and monatomic ions.

Design and discuss

6 Outline the steps required to write the empirical formula of an ionic compound, using the ionic bonding of aluminium and nitrate as an example.

5.4 Precipitation reactions

KEY IDEAS

In this topic, you will learn that:

- a precipitate is an insoluble solid that forms from chemical reactions in aqueous solutions
- a solubility table can be used to predict and identify whether a precipitate will form from chemical reactions in aqueous solutions
- precipitation reactions can be represented by balanced full and ionic chemical equations with state symbols (s), (l), (g) and (aq) included.

dissociate
break apart into smaller atoms, ions or molecules

net ionic equation
a chemical equation in which electrolytes in aqueous solution are expressed as dissociated ions

precipitation reaction
a reaction between two soluble ionic substances that forms an insoluble product

precipitate
the solid product formed from a precipitation reaction

When ionic compounds are in aqueous solution, ions **dissociate,** or break apart from the lattice structure to move freely in solution. For example, the dissociation of the ionic compound NaCl in water can be represented by the following **net ionic equation**:

$$NaCl(s) \rightarrow NaCl(aq) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

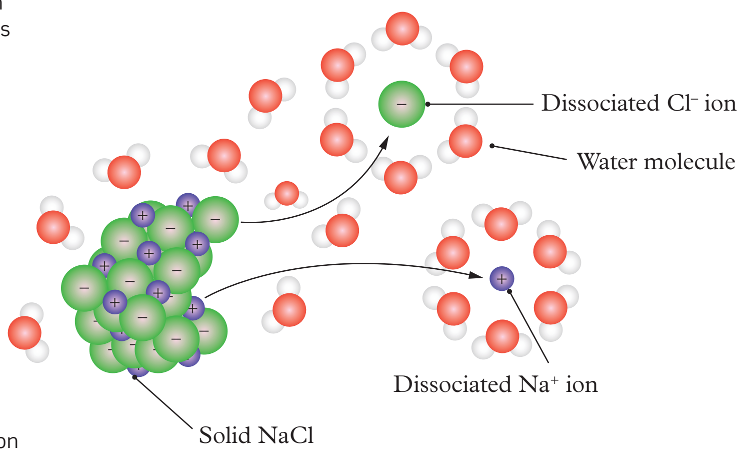


FIGURE 1 Water molecules break up the NaCl lattice, resulting in Na^{+} and Cl^{-} ions that move freely in solution.

A **precipitation reaction** is a chemical reaction that occurs when ions in aqueous solution combine to form an insoluble solid (Figure 2). This insoluble solid is called a **precipitate**. A solubility table can be used to predict and identify whether a precipitate will form from chemical reactions in aqueous solutions.

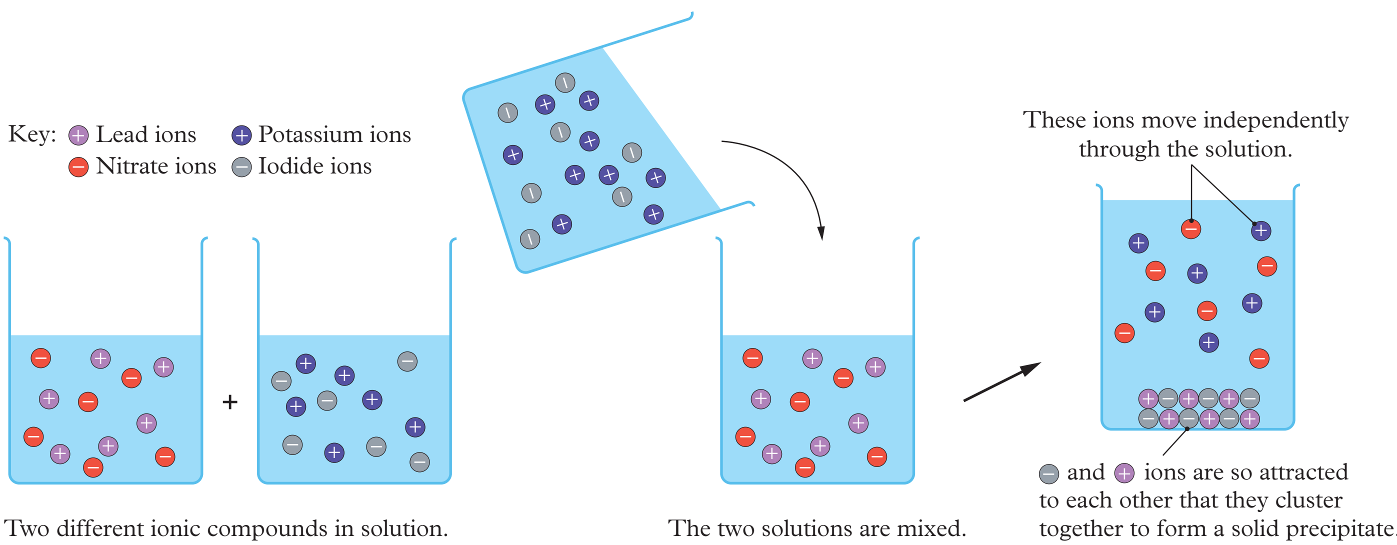


FIGURE 2 In a precipitation reaction, ions in aqueous solution combine to form an insoluble solid called a precipitate.

TABLE 1 Relative ion solubilities in water

	Soluble	Exceptions
1	Nitrates (NO_3^-)	None
2	Group 1 ions: Li^+, Na^+, K^+	None
3	Ammonium (NH_4^+) salts	None
4	Ethanoates (CH_3COO^-)	None
5	Most halide salts: chlorides (Cl^-), bromides (Br^-) and iodides (I^-)	Ag^+, Pb^{2+} and Hg^{2+} salts
6	Most sulfate (SO_4^{2-}) salts	Ag^{2+}, Ba^{2+}, Ca^{2+}, Pb^{2+}, Hg^{2+} and Sr^{2+} salts
	Insoluble	**Exceptions**
7	Most hydroxides (OH^-)	Li^+, Na^+, K^+, Ca^{2+}, Ba^{2+} and NH_4^+ salts
8	Most sulfides (S^{2-}), carbonates (CO_3^{2-}), phosphates (PO_4^{3-}) and chromates (CrO_4^{2-})	Li^+, Na^+, K^+ and NH_4^+ salts

FIGURE 3 From left to right: precipitation reactions forming copper(II) hydroxide, iron(III) hydroxide and iron(II) hydroxide

Predicting the products of aqueous reactions by using a solubility table

Consider the reaction of aqueous sodium chloride (NaCl) with aqueous silver nitrate ($AgNO_3$). Both ionic compounds dissociate in water to form their respective anions and cations, as shown in the following net ionic equations:

$$NaCl(s) \rightarrow NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$$
$$AgNO_3(s) \rightarrow AgNO_3(aq) \rightarrow Ag^+(aq) + NO_3^-(aq)$$

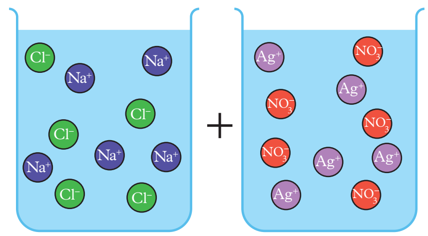


FIGURE 4 Dissociated ions of Na^+, Ag^+, Cl^- and NO_3^- in aqueous solutions

According to the solubility table, $Na^+(aq)$ and $NO_3^-(aq)$ ions are always soluble, so these ions will continue to move freely in solution. These ions are often referred to as **spectator ions** because they do not participate in the formation of the precipitate. Spectator ions can be easily identified because they do not change state. The ionic equation only includes the ions involved in the reaction and does not include spectator ions; it can be represented by the following equation:

$$Na^+(aq) + NO_3^-(aq) \rightarrow NaNO_3(aq)$$

spectator ion
an ion that does not participate in a reaction and has the same state and oxidation number as both a reactant and a product

According to the solubility table, the $Ag^+(aq)$ ions and the $Cl^-(aq)$ will react in solution to form an insoluble precipitate. In this example, the $Cl^-(aq)$ ions are not soluble because of the presence of $Ag^+(aq)$ ions. This can be represented by the following ionic equation in which the symbol '↓' (or down arrow) represents the insoluble solid formed during the precipitation reaction:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)\downarrow$$

The full, balanced chemical reaction of this precipitation reaction can be found by adding the products of the ionic equations:

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)\downarrow$$

The products formed are the soluble salt $NaNO_3(aq)$ and the insoluble precipitate AgCl(s).

Study tip

Use the mnemonic **s**mall–**m**edium–**l**arge to remember the exceptions to Rule 5 in the solubility table. All halide salts are soluble except for **s**ilver, **m**ercury and **l**ead.

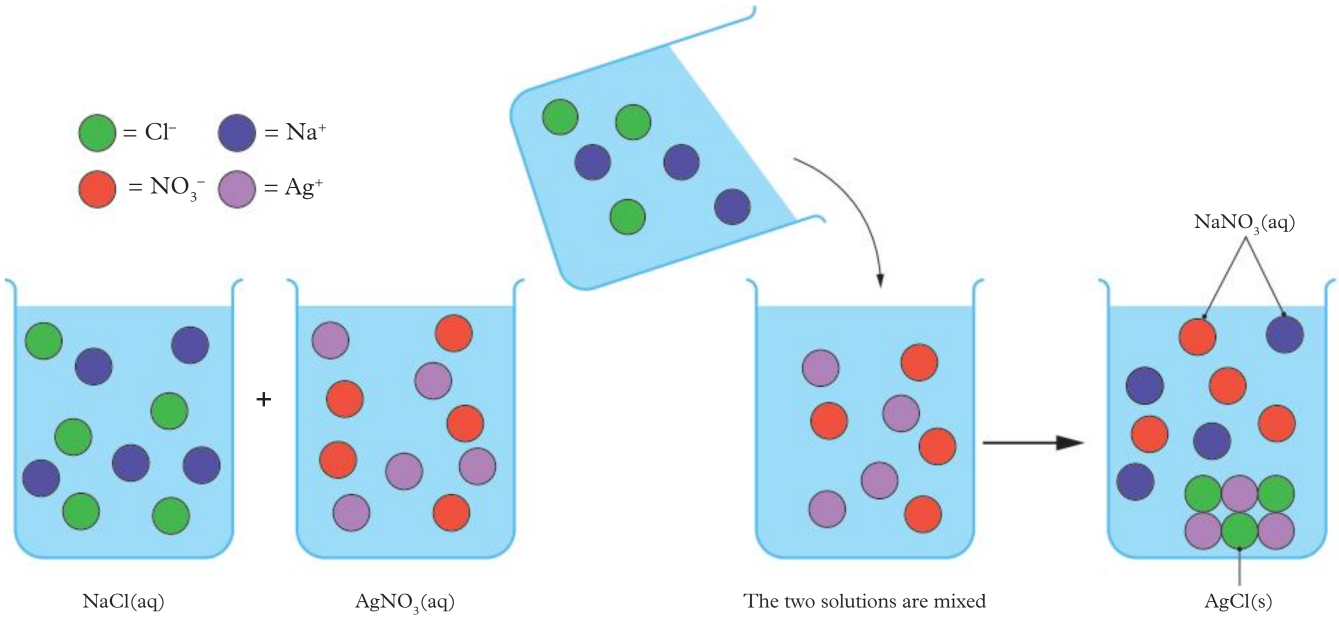


FIGURE 5 Representing the reaction products of NaCl(aq) and $AgNO_3$(aq)

double displacement reaction
a chemical reaction that occurs when two reactants exchange cations or anions to form two new products

Double displacement reactions

A **double displacement reaction** is a chemical reaction in which both sets of ions swap or exchange their partners in solution to form two new products (Figure 6). The reaction of aqueous NaCl with aqueous $AgNO_3$ is an example of a double displacement reaction. In fact, every precipitation reaction is considered a double displacement reaction. However, not all double displacement reactions form precipitates and only when they do are they considered precipitation reactions.

Study tip

Remember that a net ionic equation is a chemical equation in which only the ions undergoing chemical changes during the reaction are represented.

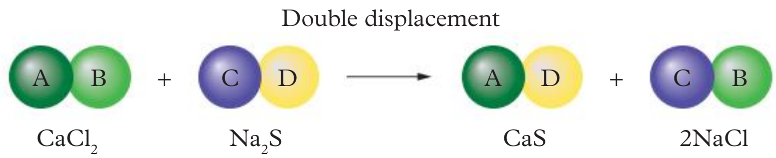


FIGURE 6 In a double displacement reaction, both sets of ions swap partners to form two new products.

Apply your understanding of precipitation reactions by exploring precipitates in nature in Real-world chemistry 5.4.

5.4 Real-world chemistry
Find me in your obook pro

5.4 WORKED EXAMPLE

PREDICTING PRODUCTS OF A REACTION IN AQUEOUS SOLUTION

Consider the reaction between aqueous solutions of barium chloride and sodium sulfate. Using Table 1 (Relative ion solubilities in water) (page 155) and Topic 5.3, Table 1 (Valencies of common ions) (page 151):

a predict the reaction products

b construct the net ionic equation for the reaction, including states

c construct the balanced chemical equation for the reaction, including states.

Solution:

Think	Do
Step 1: Use Table 1 (Topic 5.3) to write the correct ionic formula for barium chloride and sodium sulfate. Then write the ions that form them.	$BaCl_2$ and Na_2SO_4 $BaCl_2(aq)$ consists of $Ba^{2+}(aq)$ and $Cl^-(aq)$ ions. Na_2SO_4 consists of $Na^+(aq)$ and $SO4^{2-}(aq)$.
Step 2: Use the solubility rules in Table 1 (page 155) to determine the solubility of the ions present in aqueous solution.	According to the solubility table, the ions $Cl^-(aq)$ and $Na^+(aq)$ are always soluble with no exceptions. However, the $SO_4^{2-}(aq)$ ion is not soluble in the presence of $Ba^{2+}(aq)$, so a precipitate of $BaSO_4(s)$ is formed.
Step 3: Write the predicted products of the reaction as a sentence.	**a** The predicted products of this reaction are NaCl(aq) and $BaSO_4(s)$.
Step 4: Write down the ions that undergo a chemical change during this reaction.	$Ba^{2+}(aq)$ and $SO_4^{2-}(aq)$
Step 5: Write the chemical equation for the formation of the barium sulfate precipitate, from the ions in step 1.	$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)\downarrow$
Step 6: Check if the ionic equation in step 2 is balanced and that states are provided.	The charges and product-to-reactant ratio are balanced. **b** The net ionic equation for this reaction is $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)\downarrow$
Step 7: Use the valency table to write the correct ionic formulas for the reactants and products and include their states.	Reactants = $BaCl_2$ (aq) and Na_2SO_4 (aq) Products = NaCl (aq) and $BaSO_4$ (s)
Step 8: Write the unbalanced chemical equation for the reaction.	$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow NaCl(aq) + BaSO_4(s)\downarrow$
Step 9: Balance the chemical equation to ensure equal ratios of atoms on both sides of the equation. In this case a '2' needs to be placed in front of the product NaCl(aq).	**c** The balanced chemical equation for this reaction is: $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow 2NaCl(aq) + BaSO_4(s)\downarrow$

5.4 CHECK YOUR LEARNING

Describe and explain

1 What is a precipitate and how is it formed?

2 Define 'spectator ion' in a precipitation reaction.

3 Explain the term 'dissociation', using a balanced chemical equation.

Apply, analyse and compare

4 Describe the difference between writing a net ionic equation and writing a balanced equation for a complete reaction in an aqueous solution.

5 Use the solubility table (Table 1) to predict the products of the following reactions and construct their balanced chemical equations (including states):

- **a** magnesium iodide with sodium carbonate
- **b** mercury nitrate with sodium sulfide
- **c** sodium hydroxide with copper sulfate.

Design and discuss

6 Design a mnemonic or memory device to help you remember the rules of the solubility table.

CHAPTER 5 Review

Chapter summary

5.1
- An ionic bond is formed between positive cations and negative anions, held together by strong electrostatic forces, in a repeating, ordered, three-dimensional crystal lattice.
- Ionic compounds are hard, brittle, have high melting and boiling points, are soluble in water and conduct electricity in a liquid state.

5.2
- Ions are formed by transfer of an atom's valence electrons.
- This transfer is represented by drawing an electron transfer diagram.

5.3
- Ionic compounds can be formed from ions that have more than one atom.
- The formula of an ionic compound is the ratio of cations (positive ions) to anions (negative ions) so that the overall charge of the compound is equal to zero.
- There are rules for writing and naming ionic compounds.

5.4
- A precipitate is an insoluble solid that forms from chemical reactions in aqueous solutions.
- Use of a solubility table will predict and identify whether a precipitate will form from chemical reactions in aqueous solutions.
- Precipitation reactions can be represented by balanced full and ionic chemical equations with state symbols (s), (l), (g) and (aq) included.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• describe the common properties of ionic compounds, including brittleness, hardness, melting point, difference in electrical conductivity in solid and molten liquid states, with reference to the nature of ionic bonding and crystal structure	☐	☐	☐	Go back to Topic 5.1
• deduce the formula and name of an ionic compound from its component ions, including polyatomic ions (NH_4^+, OH^-, NO_3^-, HCO_3^-, CO_3^{2-}, SO_4^{2-} and PO_4^{3-})	☐	☐	☐	Go back to Topic 5.2
• describe how ionic compounds form and write ionic compound formulas, including those containing polyatomic ions and transition metal ions	☐	☐	☐	Go back to Topic 5.3
• use solubility tables to predict and identify precipitation reactions between ions in solution, represented by balanced full and ionic equations including the state symbols: (s), (l), (aq) and (g)	☐	☐	☐	Go back to Topic 5.4

Revision questions

Multiple choice

1 Element Y forms an iodide with the formula YI_2. It follows that element Y would form:

A an oxide with the formula Y_2O_3.

B a nitrate with the formula YNO_3.

C a sulfate with the formula Y_2SO_4.

D a hydroxide with the formula $Y(OH)_2$.

2 Potassium chloride does not conduct electricity in the solid state, but it does when molten. The reason the change in conductivity occurs with the change in physical state is that the process of melting:

A releases the sea of electrons held between the potassium cations.

B turns potassium chloride into a liquid, and all liquids conduct electricity.

C breaks the strong ionic bonds existing between the potassium and chloride ions.

D ionises potassium and chlorine atoms, forming potassium cations and chloride ions.

3 An ionic solid may form when:

A copper reacts with tin.

B carbon reacts with oxygen.

C chlorine reacts with oxygen.

D magnesium reacts with oxygen.

4 When a fluoride ion forms from a fluorine atom, which of the following does not change?

A The charge on the particle

B The number of inner shell electrons

C The number of electrons in the valence shell

D The diameter of the particle

5 An element, X, that would combine with aluminium to form an ionic compound, Al_2X_3, is most likely to occur in:

A group 16.

B group 11.

C group 2.

D group 1.

6 Which of the following equations represents a precipitation reaction?

A $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

B $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

C $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

D $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

7 Which of these formulas represents a compound in which the charges of the cations and anions are balanced (meaning the compound would have no overall charge)?

A Al_2O_3

B Ca_2O

C Li_3S

D H_3Cl_2

8 What is the name of the compound with formula K_2O?

A Potassium dioxide

B Potassium oxygen

C Potassium oxide

D Potassium oxalate

9 The chemical formulas of three anions are thiosulfate ion $S_2O_3^{2-}$, permanganate ion MnO_4^- and phosphate ion PO_4^{3-}. Which one of the following lists of formulas is correct?

A LiS_2O_3, $Zn(MnO_4)_2$, $Al_3(PO_4)_2$

B $Li_2S_2O_3$, $Zn(MnO4)_2$, $AlPO_4$

C LiS_2O_3, Zn_2MnO_4, $Al_2(PO_4)_3$

D $Li_2S_2O_3$, $ZnMnO_4$, $Al_3(PO4)_3$

10 Ionic substances are hard but brittle because the forces of attraction between particles are:

A strong and cannot adapt to change.

B strong and can adapt to change.

C weak and cannot adapt to change.

D weak and can adapt to change.

Short answer

Describe and explain

11 Describe why the formation of a K^{2+} cation and a Cl^{2-} anion is unlikely.

12 The arrangement of potassium ions and fluoride ions in solid potassium fluoride is represented in Figure 1. The arrangement extends in three dimensions to represent a crystal.

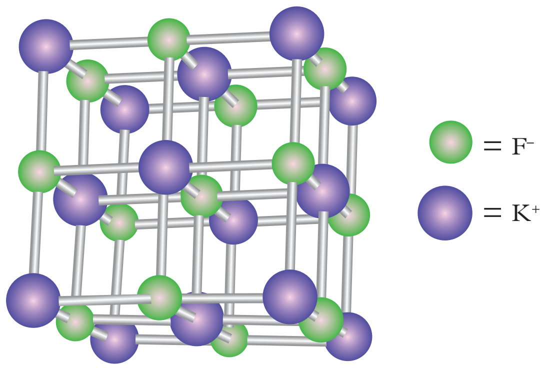


FIGURE 1 The arrangement of potassium and fluoride ions in potassium fluoride

a State the empirical formula of potassium fluoride.

b Would it be possible for magnesium fluoride to have the same particle arrangement as potassium fluoride? Explain your answer.

c Explain why molten potassium fluoride is a good conductor of electricity, whereas solid potassium fluoride is not.

13 Explain why all precipitation reactions are considered double displacement reactions but not all double displacement reactions are considered precipitation reactions.

Apply, analyse and compare

14 Write ionic formulas for the following pairs of ions.

a Chromium(III) and phosphide

b Iron(III) and oxide

c Lead(II) and nitrate

d Ammonium and carbide

15 Calcium perchlorate and gold chloride are ionic compounds with formulas of $Ca(ClO_4)_2$ and $AuCl_3$ respectively. What is the formula of gold perchlorate?

16 Predict what will happen when a colourless solution of lead(II) nitrate is added to a test tube containing another colourless solution of potassium iodide. Explain your answer by writing a balanced chemical equation (including states).

17 A freshly cut piece of potassium reacts violently in a gas jar containing chlorine gas (Cl_2). White smoke consisting of potassium chloride is formed and eventually deposits as a white solid in the gas jar.

a Write a balanced equation for the reaction between potassium and chlorine gas.

b Explain the changes, in detail, that occur in both the potassium and chlorine atoms during the reaction. Use electron transfer diagrams to illustrate your explanation.

18 Write the ionic formulas of the following compounds.

a Nickel(II) sulfate

b Ammonium chloride

c Aluminium hydrogen phosphate

d Iron(III) nitrate

e Silver sulfide

f Calcium carbide

19 Which of the substances in the table below is most likely to be an ionic compound? Explain your choice.

Substance	Melting temperature (°C)	Conducts at 700°C?	Conducts at 800°C?
A	750	Yes	Yes
B	730	No	Yes
C	1600	No	No

20 Name the following ionic compounds.

a NH_4Br

b $Pb(SO_4)_2$

c $Fe(NO_3)_3$

d Cu_3P

e $Al_2(SO_3)_3$

f K_2CO_3

21 Use the solubility table (Table 1, page 155) to predict the products of the following reactions and construct their net ionic and balanced chemical equations (including states).

a Silver nitrate and potassium chromate

b Aluminium sulfate and calcium hydroxide

c Lead(II) nitrate and sodium chloride

d Barium nitrate plus lithium sulfate

22 Compare the amount of energy that would be required (higher or lower) to melt the following pairs of ionic compounds, giving reasons for your choice.

a Potassium iodide and magnesium oxide

b Potassium iodide and lithium fluoride

c Barium sulfide or magnesium oxide

d Sodium chloride or sodium iodide

Design and discuss

23 Use electron transfer diagrams to show the formation of magnesium sulfide from magnesium and sulfur atoms.

24 Use the solubility table (Table 1, page 155) to construct balanced equations and predict the products of the following reactants in aqueous solution.

a $NaOH(aq)$ and $CaCl_2(aq)$

b $CuBr_2(aq)$ and $(NH_4)_2CO_3(aq)$

c $K_2SO_4(aq)$ and $Fe(NO_3)_3(aq)$

25 When iron is left exposed to the atmosphere, a red-brown, flaky compound of iron(III) oxide is formed.

a Write the correct formula for iron(III) oxide.

b Show how this chemical reaction occurs using electron transfer diagrams.

26 Design an ionic bonding game you could play with your classmates to practise naming ionic formulas or writing empirical formulas.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER

6 Separation and identification of the components of mixtures

KEY KNOWLEDGE

- → polar and non-polar character with reference to the solubility of polar solutes dissolving in polar solvents and non-polar solutes dissolving in non-polar solvents
- → experimental application of chromatography as a technique to determine the composition and purity of different types of substances, including calculation of R_f values

FIGURE 1 Chromatography separates the components of mixtures.

GROUNDWORK

In Chapter 6, you will learn that the polarity of substances influences their solubility and how this property can be applied to separate and identify components in a mixture.

This chapter will build on concepts you have already learnt in Chapter 3. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

6A How can you determine the polarity of a molecule?

6A Groundwork resource
Polarity

6B What are the three types of intermolecular forces?

6B Groundwork resource
Intermolecular forces

6C How do molecular properties affect the formation of intermolecular forces?

6C Groundwork resource
Molecular properties

PRACTICALS

6.1 PRACTICAL: SIMULATION	How do the sample components behave in chromatography?	Page 504
6.2 PRACTICAL: CONTROLLED EXPERIMENT	How can chromatography be used to separate food dyes in Skittles?	pro

6.1 Polar and non-polar character in chromatography

KEY IDEAS

In this topic, you will learn that:

+ like dissolves like – polar substances will dissolve in polar solvents, non-polar substances will dissolve in non-polar solvents
+ chromatography separates components in a mixture according to their affinity for the mobile and stationary phases
+ components of a mixture are separated according to the strength of their intermolecular forces.

Study tip

The *like dissolves like* rule will help you to remember that substances will dissolve in solvents that have similar properties, particularly polarity.

FIGURE 1 Sugar is a polar substance that is soluble in water.

Separation and identification of components in a mixture have many important applications, such as purifying synthetic medicines and analysing water quality. One important property that can be used to separate mixtures is polar or non-polar character. You learnt about this in Chapter 3.

The 'like dissolves like' rule

The interactions between polar and non-polar substances can be described using **solubility**. Solubility refers to the ability of a substance to dissolve in a liquid, which depends on the relative properties of the **solute** and **solvent** molecules.

solubility
the ability of a substance to dissolve

solute
the substance being dissolved in the solvent

solvent
the substance that another is being dissolved in, e.g. water

For a substance to be soluble in another substance, the intermolecular forces between the solvent and the solute must be stronger than the solvent–solvent forces and the solute–solute forces that already exist.

In Chapter 3, you learnt about intermolecular forces between molecules. In order of strongest to weakest, they are:

- hydrogen bonds
- dipole–dipole attractions
- dispersion forces.

The type of intermolecular forces present depends on the polarity of the molecules. For example, oil (a non-polar substance) is not soluble in water (a polar substance), so it forms two separate layers. But ethanol (a polar substance) is soluble in water. *Like dissolves like!*

Study tip

For a solute to be soluble, the solvent–solute interactions must be stronger than the solvent–solvent and solute–solute interactions.

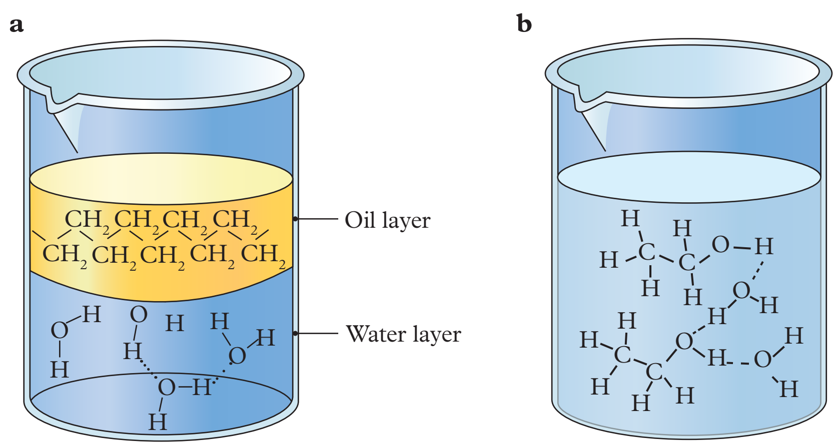


FIGURE 2 Like dissolves like. **a** Non-polar oil and polar water do not mix. **b** However, polar ethanol is soluble in water.

6.1
PRACTICAL: SIMULATION
How do the sample components behave in chromatography? Go to page **504**.

In this example, there are limited intermolecular forces between oil and water. The oil–oil (solute–solute) forces and water–water (solvent–solvent) forces are stronger than the oil–water (solute–solvent) forces (Figure 2a).

But ethanol and water interact strongly through hydrogen bonding. The ethanol–water (solute–solvent) forces can overcome ethanol–ethanol (solute–solute) and water–water (solvent–solvent) forces (Figure 2b).

In general, water can **dissolve** other polar substances very well because it can form strong hydrogen bonds or dipole–dipole attractions with them.

Principles of chromatography

Chromatography is an analytical technique that gives both qualitative and quantitative data. It uses molecular properties, such as polarity and intermolecular forces, to separate components of a mixture. You can use chromatography to identify components in a mixture and determine how much of each component is present.

Mobile and stationary phases

Chromatography uses two **phases** to separate a mixture:

- **mobile phase** – a solvent that moves over or through the stationary phase (e.g. methanol, water, acetonitrile)
- **stationary phase** – a substance that stays still (either a solid with a high surface area or a liquid coated onto a solid support).

The mobile phase moves through a column and carries a sample (solute) over the stationary phase.

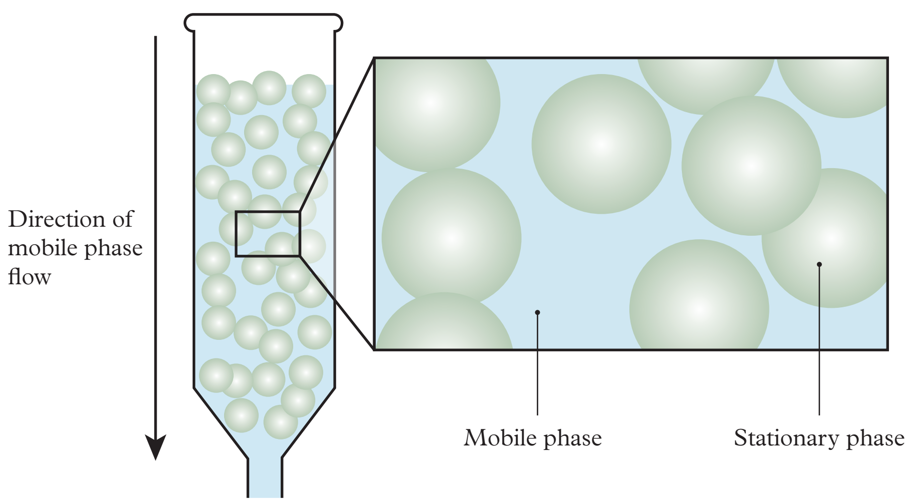


FIGURE 3 The mobile phase and stationary phases in column chromatography. You will learn about these in Topic 6.2.

Depending on their intermolecular forces, components of the sample can be more attracted to, or have a greater **affinity** for, the mobile phase or the stationary phase. This makes the components move at different rates and separates them.

dissolve
when solute–solute bonds break and the solute molecules mix evenly with the solvent molecules

chromatography
an analytical technique that separates the components of a sample mixture based on the properties of the molecules

phase
(in chromatography) a form of matter that has uniform chemical and physical properties; can be a pure substance or a mixture

mobile phase
the solvent phase that flows, moving the components of a sample over the stationary phase

stationary phase
the solid phase to which the components of a sample are adsorbed; can sometimes be a liquid coated onto a solid support

Study tip

Mobile phase **m**oves, **st**ationary phase **st**ays still.

Study tip

Make sure you understand the principles of chromatography – you will need this for Units 3 and 4.

affinity
the strength of the interaction of a substance within a sample with the mobile or stationary phase

Affinity depends on two types of interactions (Figure 4):

- **adsorption**: components of the sample adsorb, or 'stick', onto the stationary phase from the mobile phase if they are attracted to, or have an affinity for, the stationary phase
- **desorption**: components of the sample desorb off, or 'unstick' from, the stationary phase and return to the mobile phase if they are attracted to, or have an affinity for, the mobile phase.

Study tip

Adsorption is when a substance sticks to a surface, usually in a single layer. **Ab**sorption is when a substance penetrates a solid or liquid.

adsorption
the attraction of a substance within a sample to the stationary phase; how the substance 'sticks' to the stationary phase

desorption
the release of a substance within a sample from the stationary phase into the mobile phase; how the substance 'unsticks' from the stationary phase

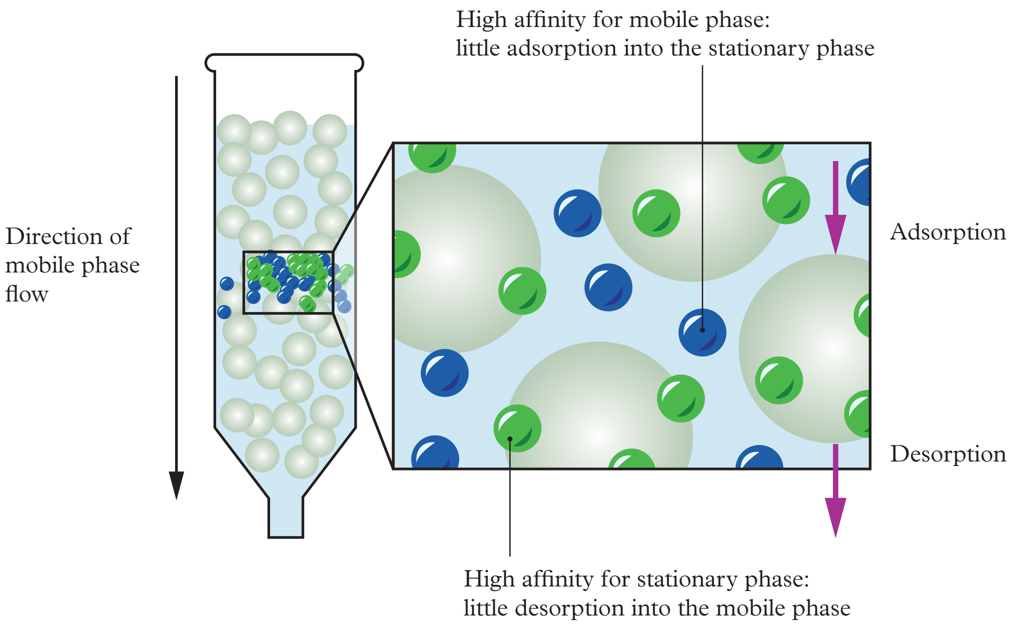


FIGURE 4 The components in a sample can be separated depending on their affinity for the mobile and stationary phases.

The effect of different solute affinities for the mobile and stationary phases can be seen in Figure 5.

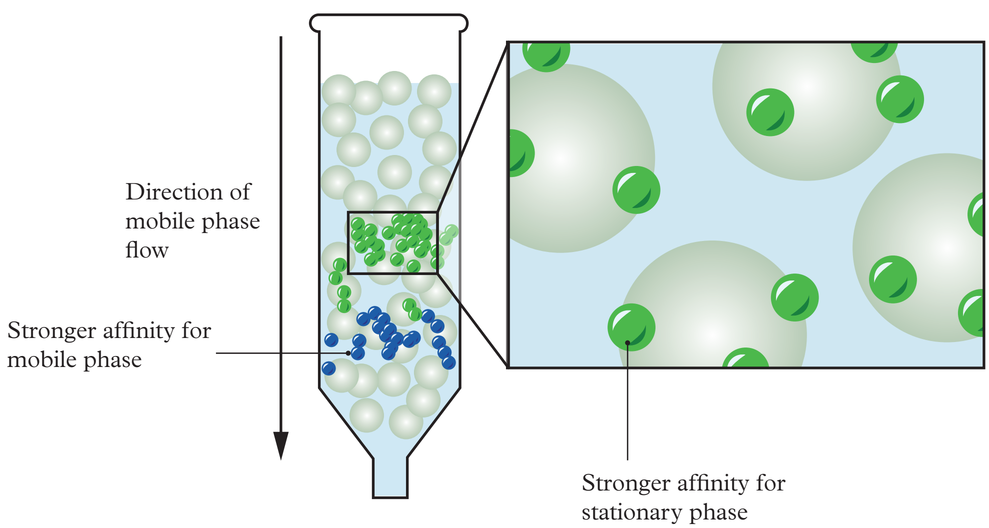


FIGURE 5 Solutes separate out according to their different affinities for the mobile and stationary phases.

If a molecule has a stronger affinity for the mobile phase (like the blue solute in Figure 4), it will desorb more readily from the stationary phase. It will travel further with the mobile phase and leave the column first.

If a molecule has a stronger affinity for the stationary phase (like the green solute in Figure 4), it will adsorb more readily to the stationary phase. It will take longer to travel through the mobile phase and leave the column later.

Properties of the solute, mobile phase and stationary phase, such as polarity, affect adsorption and desorption, and are therefore extremely important in the separation of molecules.

Separating out solutes with polar character

Substances will have a higher affinity for a mobile or stationary phase if they have the same or similar properties. But, in order to separate mixtures, the mobile and stationary phases need to have different (and sometimes opposite) properties so that solute molecules in the mixture will interact and have different affinities for the stationary and mobile phases.

For example, a non-polar stationary phase can be used with a polar mobile phase to separate solutes with different polarities from a sample mixture. The non-polar solutes will have high affinity for the non-polar stationary phase. Meanwhile, the polar solutes will desorb into, and form dipole–dipole attractions and/or hydrogen bonds with, the polar mobile phase. This will allow the solutes to move through the mobile phase at different rates.

Study tip

Polar solutes will have a:

- high affinity for polar mobile and stationary phases
- low affinity for non-polar mobile and stationary phases.

Non-polar solutes will have a:

- high affinity for non-polar mobile and stationary phases
- low affinity for polar mobile and stationary phases.

Dipole–dipole attractions

Polar solutes within a mixture contain permanent dipoles. They will interact with a polar mobile phase or stationary phase, which also contains permanent dipoles. You can review how dipole–dipole attractions form in Topic 3.4.

If a non-polar stationary phase and polar mobile phase are used, a polar solute will have a high affinity for the mobile phase because of dipole–dipole attractions. It will also have low affinity for and desorb more readily from the stationary phase. You would expect the polar solutes to leave the column first.

Hydrogen bonding

You might also remember from Chapter 3 that hydrogen bonding is a stronger type of dipole–dipole attraction that occurs when a molecule has a hydrogen atom covalently bonded to a fluorine, oxygen or nitrogen atom.

Water is a molecule well known for its ability to undergo hydrogen bonding. It is a polar molecule that is commonly used as a solvent in mobile phases. When water is used as a mobile phase, polar solutes in the sample will interact by hydrogen bonding with the water molecules. This will give them a higher affinity for the mobile phase and they will desorb more readily from the stationary phase. You would expect the polar solutes to leave the column first.

FIGURE 6 Water is a common solvent used as a mobile phase.

Separating out solutes with non-polar character

Non-polar components (solutes) within a sample mixture will dissolve into or have stronger intermolecular interactions with a non-polar solvent (the mobile phase). They will also have stronger intermolecular interactions with a non-polar stationary phase.

Study tip

Dispersion forces are the weakest intermolecular force, followed by dipole–dipole attractions. Hydrogen bonding is the strongest.

Dispersion forces

In Chapter 3, you also learnt that weak dispersion forces exist between all molecules. This is due to electrons moving constantly and randomly throughout a covalent molecule. At any given time, there will be more electrons on one side of the molecule than the other, creating a temporary dipole or dipole moment.

Dispersion forces are the main intermolecular force that allow non-polar solutes to dissolve in or have affinity for non-polar substances. These solutes cannot dissolve in polar solvents because the dipole–dipole attractions and hydrogen bonds between solvent molecules are too strong to overcome.

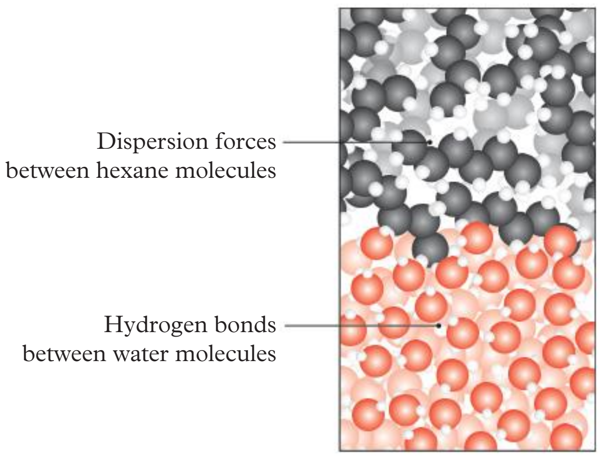


FIGURE 7 The lack of interaction between polar water molecules and non-polar hexane molecules.

Figure 7 shows a mixture of water and hexane. The water molecules interact by hydrogen bonding. These interactions are so strong that the hexane molecules, which interact with other hexane molecules by dispersion forces, cannot mix with them. For this reason, hexane forms a layer rather than dissolving in water.

If a non-polar stationary phase and polar mobile phase are used, a non-polar solute will have low affinity for the mobile phase. This is because the dipole–dipole attractions or hydrogen bonds are too strong to overcome. Instead, the solute would have a stronger affinity for and adsorb more readily to the non-polar stationary phase.

Separation based on carbon chain length

In some cases, we cannot separate mixtures on the basis of polar or non-polar character. One example is a mixture of methanol, ethanol, propan-1-ol and butan-1-ol (Figure 8). All of the molecules have the same polar –OH group on the terminal (or end) carbon. Instead, the mixture is separated by the size of the carbon chain.

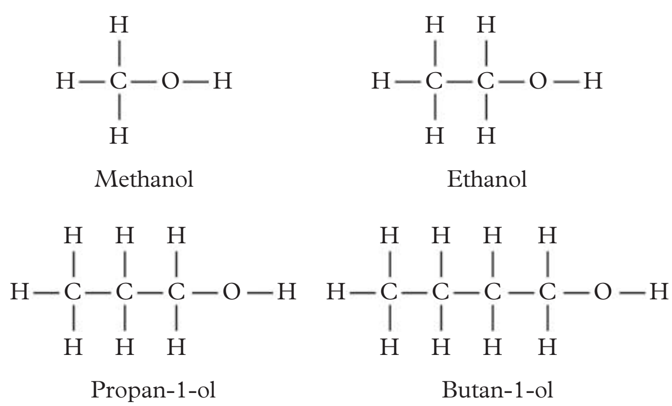


FIGURE 8 Methanol, ethanol, propan-1-ol and butan-1-ol all have a terminal polar –OH group and therefore cannot be separated by polarity.

The longer the carbon chain, the greater the dispersion forces, because more electrons can form a temporary dipole more frequently. This means that there will be greater affinity for a non-polar stationary phase.

The types of forces between solute–solvent molecules are summarised in Table 1. This will be useful to help you predict the separation of solutes, as shown in Worked example 6.1. Real-world chemistry 6.1 also shows how chromatography can be applied to the separation of amino acids.

6.1 Worked example
Find me in your obook pro

6.1 Worked example
Video tutorial

6.1 Real-world chemistry
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TABLE 1 Solubility, intermolecular interactions and affinity for mobile and stationary phases of polar and non-polar solutes

Solubility	Polar solute	Non-polar solute
Soluble in polar solvents?	Yes	No
Soluble in non-polar solvents?	No	Yes
Types of solute–solvent intermolecular forces:		
• Dipole–dipole attractions	✓	
• Hydrogen bonding	✓	
• Dispersion forces	✓	✓
Affinity for:		
• Polar mobile or stationary phase	High	Low
• Non-polar mobile or stationary phase	Low	High

6.1 SKILL DRILL

Evaluating data and designing a method to investigate solubility

Key science skill: Analyse and evaluate data and investigation methods

To determine whether some molecules are polar, a chemist dissolves each of them in a test tube of water. The chemist hypothesises that if the solute is polar, it will dissolve in the water, and if it is non-polar it will form layers.

The chemist tests the following substances and obtains the following results.

Solute	Result
Sodium chloride	Dissolves in water
Potassium nitrate	Dissolves in water
Methanol	Dissolves in water
Glucose	Dissolves in water
Ethanol	Dissolves in water
Octane	Forms layers with water

Practise your skills

1 Identify an issue with the chemist's hypothesis.
2 Explain why the data collected is qualitative and not quantitative.
3 Design a method to determine which molecule has the strongest intermolecular forces. Hint: What property could you study that is an indicator of intermolecular strength?

Need help analysing and evaluating data? See Topic 1.8 (page 24).

6.1 CHALLENGE

Choosing mobile and stationary phases

A chemist wishes to separate a mixture of sodium chloride, butanol (containing C–C and C–H bonds and an –OH group) and hexane (containing only C–C and C–H bonds). What properties would you recommend the mobile phase and stationary phase have, in order to separate all components of the sample? Justify your answer by discussing the intermolecular forces involved.

6.1 CHECK YOUR LEARNING

Describe and explain

1 Define the following terms and explain why they are essential for the separation of mixtures.

- **a** Mobile phase
- **b** Stationary phase
- **c** Affinity
- **d** Desorption
- **e** Intermolecular forces and the 'like dissolves like' rule

2 Are polar components of a mixture more attracted to a polar or non-polar mobile phase? Explain, using an example.

Apply, analyse and compare

3 A chemist is separating a mixture containing esters and alcohols (see Figure 9). Determine which substance will have a stronger affinity for a polar mobile phase.

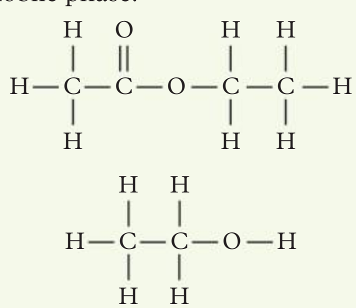


FIGURE 9 Molecular structures of ethyl ethanoate (an ester) and ethanol (an alcohol)

4 Compare the intermolecular forces (from Chapter 3) that could be used to separate a mixture of substances.

5 A mobile phase primarily consists of water. Evaluate why water is chosen as a mobile phase.

6 A chemist needs to separate a sample of dyes that have different polarities.

Component	Polarity
A	Slightly polar
B	Polar
C	Least polar

- **a** Predict which component would have the strongest affinity for a polar solvent or mobile phase. Explain your response.
- **b** Predict which component would have the strongest affinity for a non-polar stationary phase. Explain your response.
- **c** Predict which component would adsorb most readily to a polar stationary phase. Explain your response.
- **d** Predict which component would desorb most readily from a polar stationary phase. Explain your response.

Design and discuss

7 A mobile phase used in an analysis contains only dispersion forces. Discuss the intermolecular forces that the stationary phase should have.

8 Why do like forces interact? Discuss how the 'like dissolves like' rule works.

6.2 Chromatography

KEY IDEAS

In this topic, you will learn that:

+ substances can be separated by paper, thin-layer and column chromatography
+ calculating R_f values helps to identify substances within a mixture
+ separation can be optimised by changing the mobile and stationary phases.

paper chromatography
an analytical technique for separating and identifying mixtures; the stationary phase is a thin strip of absorbent paper

origin
the line applied to a chromatogram to mark the point where the sample or standard is placed

capillary action
the movement of a liquid through a narrow space (e.g. the cellulose network in absorbent paper) without any help, usually against gravity

solvent front
the point the mobile phase reaches on a chromatogram before the analysis is terminated

chromatogram
the pattern of bands, spots or peaks formed on the chromatography paper or TLC plate, demonstrating the separation of a mixture

The word 'chromatography' is derived from the Greek words *chroma* (colour) and *graphein* (to write). Early chromatography was used to separate photosynthetic plant pigments into four coloured components: orange, yellow, blue and green. Not long after this was achieved, it was discovered that chromatography could be applied to most other mixtures of chemicals, regardless of whether they were coloured or not.

In this topic, you will learn about three types of chromatography:

- paper chromatography
- thin-layer chromatography (TLC)
- column chromatography.

Paper chromatography

Paper chromatography in an analytical technique used to separate and identify components of mixtures. It is also useful to detect contaminants or impurities.

It uses absorbent paper made from cellulose, which contains many polar –OH groups. However, the fibres form a network that has few intermolecular forces with other molecules. This makes it non-polar compared with polar solvents such as water, which are used as a mobile phase. Any component of the sample that has whole or partial charges, such as ionic or polar substances, has a higher affinity for the water mobile phase. Any component of the sample that only forms dispersion forces has a higher affinity for the stationary phase.

The steps of paper chromatography are summarised in Figure 1.

1. A thin strip of absorbent paper is cut to fit inside a container that holds the mobile phase. A pencil line is ruled across the bottom of the paper to mark where the sample will be placed – this is called the **origin**. Pen or ink should not be used to mark the origin because inks will separate and contaminate the sample.
2. The sample is placed on the paper and standards of known identity can also be added to help identify the substances in the sample. In Figure 1 (on the next page), the standards are marked A, B and C.
3. The mobile phase is added to the container so that it sits below the origin line.
4. The paper is placed into the container and left to absorb the mobile phase. The sample travels up the piece of paper by **capillary action**, adsorbing and desorbing to separate the mixture into its components.
5. When the solvent is near the top of the paper (the **solvent front**), the paper is removed from the solvent. The final pattern of separated components is called a **chromatogram**.

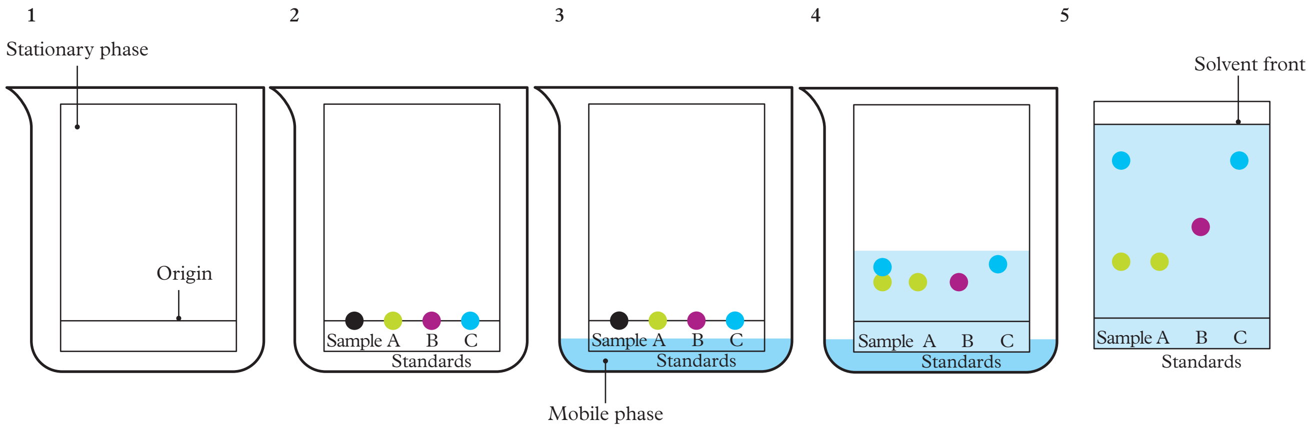


FIGURE 1 The mobile phase moves up a piece of filter paper (stationary phase) to separate a sample of ink from a black marker pen. Components in the sample can be identified by comparing them to known standards A–C.

Components that have a strong affinity for the stationary phase move slowly (green dots). They do not move far from the origin. Components that have a higher affinity for the mobile phase move faster (blue dots). They have moved further and are located closer to the solvent front.

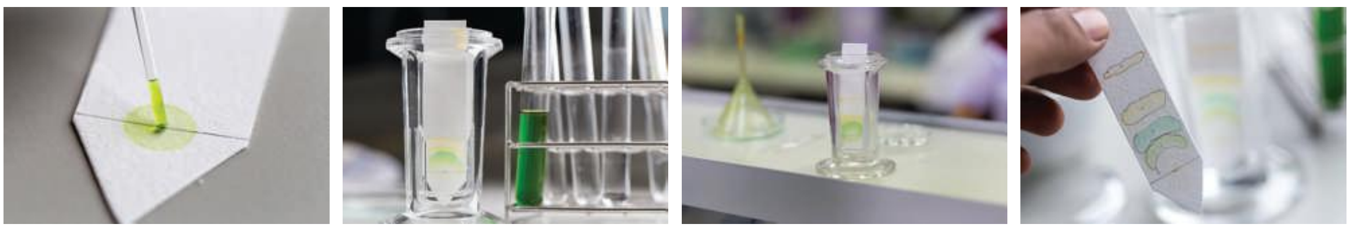


FIGURE 2 Paper chromatography can be used to separate the components of a photosynthetic pigment (leaf stain). The pigment is placed on the origin, and then the paper strip is placed in a container where it is in contact with the mobile phase. The mobile phase is drawn up the paper by capillary action.

Retardation factor (R_f) calculations

retardation factor (R_f) the ratio of the distance travelled by a component of a sample, from the origin, in relation to origin, in relation to the distance travelled by the mobile phase

Once separation has been achieved, the resulting chromatogram is analysed to identify the substances in the sample and to determine the purity of the sample. Although we can simply look at a chromatogram to judge whether two substances are identical, it is more precise to calculate the **retardation factor (R_f)** of a substance. The term 'retardation' refers to the slowing down of something. Therefore, in chromatography, the retardation factor of a substance is how much the stationary phase is slowing down the movement of the substance. Specifically, it is the ratio of the distance moved by a substance to the distance moved by the solvent, or mobile phase:

$$R_f = \frac{\text{distance moved by the solute from the origin}}{\text{distance moved by the solvent from the origin}}$$

This can be simplified to:

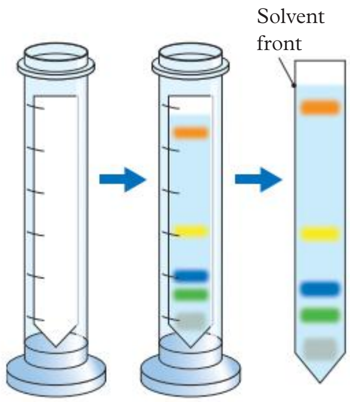
$$R_f = \frac{\text{distance of sample dot}}{\text{distance of mobile phase}}$$

Each molecule has a unique R_f depending on the properties of the mobile and stationary phases. Increasing the polarity of the mobile phase increases its affinity for polar components within the sample. This makes the substances move further up the paper.

Study tip

The distance moved by the solvent and the solute should be measured from the origin line, not the bottom of the stationary phase.

All R_f values must be expressed as a decimal (never a fraction) and cannot be greater than or equal to one. An R_f of 1 indicates that the substance has not separated from the mobile phase. In this case, the analysis is not valid and that component cannot be identified. The analysis must be run again under different conditions of mobile or stationary phase. The R_f values for carotene, xanthophyll, chlorophyll a and chlorophyll b, which have been separated from leaf stain, are calculated in Figure 3.



Pigment	**R_f value**
Carotene	0.95
Xanthophyll	0.48
Chlorophyll a	0.30
Chlorophyll b	0.18
Leaf stain	

FIGURE 3 Separation of a leaf stain into four pigment components. The R_f values are calculated from the chromatogram.

Using chromatography to determine purity

A chromatogram can be used to work out what substances are present in a mixture, because each substance will eventually produce its own spot on the paper.

A pure substance contains only one kind of particle, all of which have the same physical properties. So, a pure substance should only produce one spot on a chromatogram.

If a substance contains impurities, these components will be separated according to their physical properties. Each substance, or impurity, will be identified by its own spot on the chromatogram. An example analysis is shown in Worked example 6.2.

6.2 WORKED EXAMPLE

CALCULATING R_f

A food dye was analysed against three standard dyes A–C to determine which dyes it contained. The resulting chromatogram is shown in Figure 4.

Use the data to determine:

a the R_f values of the dye standards

b the R_f values of all dyes within the sample

c which dye is more strongly attracted to the stationary phase

d which dye is more strongly attracted to the mobile phase

e which dyes are in the sample

f if the analysis of this data is valid.

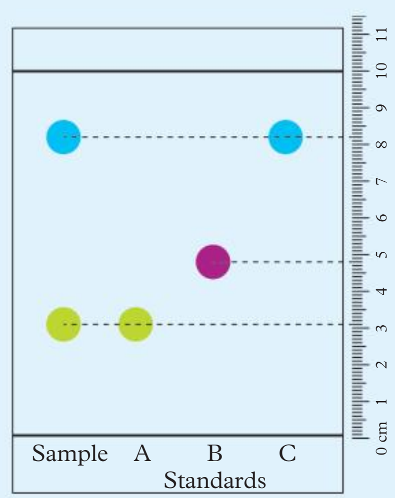


FIGURE 4 The chromatogram from analysing a food dye against three standard dyes

Solution

Think	Do
Step 1: Use the ruler on the side of the chromatogram to read the distance travelled by the mobile phase and each standard spot.	Mobile phase = 10 mm A = 3.1 mm B = 4.8 mm C = 8.2 m
Step 2: Use the formula to calculate R_f. $R_f = \frac{\text{distance of sample dot}}{\text{distance of mobile phase}}$	**a** For A: $R_f = \frac{3.1}{10} = 0.31$ For B: $R_f = \frac{4.8}{10} = 0.48$ For C: $R_f = \frac{8.2}{10} = 0.82$
Step 3: Use the ruler on the side of the chromatogram to read the distance travelled by the sample spots.	3.1 mm and 8.2 mm
Step 4: Use the formula to calculate R_f.	**b** $R_f = \frac{3.1}{10} = 0.31$ $R_f = \frac{8.2}{10} = 0.82$
Step 5: To determine the dye that is more strongly attracted to the stationary phase, identify the standard dye with the lowest R_f (closest to the origin).	**c** Dye A is most strongly attracted to the stationary phase.
Step 6: To determine the dye that is more strongly attracted to the mobile phase, identify the standard dye with the highest R_f (closest to the solvent front).	**d** Dye C is most strongly attracted to the mobile phase.
Step 7: To determine the dyes that are in the sample, match the R_f values with those of the standards.	**e** The R_f values of 0.31 and 0.82 correspond to dyes A and C.
Step 8: To determine if the analysis is valid, check that the calculated R_f vales are *not* greater than or equal to 1.	**f** All R_f values are less than 1, so all of the substances have separated from the mixture and the data is valid.

thin-layer chromatography (TLC)
an analytical technique for separating and identifying mixtures; the stationary phase is typically a thin layer of silica gel, aluminium oxide or cellulose supported on a piece of glass or plastic

Thin-layer chromatography

Thin-layer chromatography (TLC) works by the same principles as paper chromatography, but one major change makes the process more efficient (both faster and able to detect much smaller amounts of substance).

In TLC, the stationary phase consists of a sheet of glass or plastic that is thinly coated in aluminium oxide (most polar), silica gel or cellulose (least polar). This coating is spread on the surface of the plastic or glass; hence, the name 'thin-layer' chromatography.

The mobile phase for TLC is usually a mixture of organic solvents that are less polar than the stationary phase, including dichloromethane, ethyl acetate, hexane and other hydrocarbons.

Study tip

Never leave R_f as a fraction. This is because it is analysed on a scale of 0–1, with 0 being no affinity for the mobile phase and 1 being no affinity for the stationary phase.

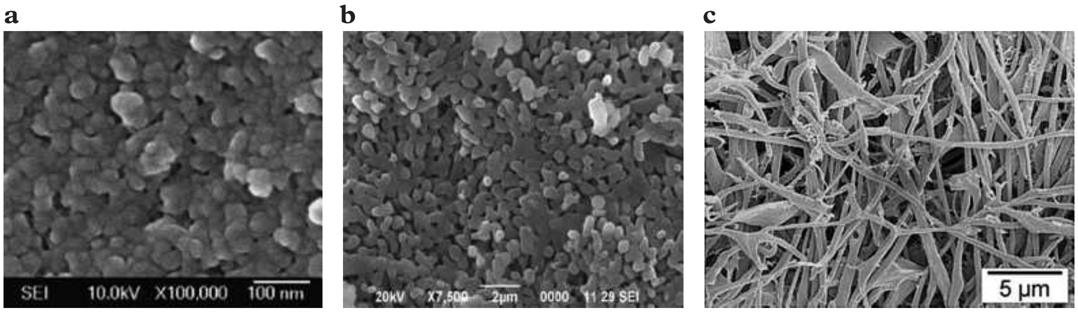


FIGURE 5 Electron micrographs of **a** silica gel, **b** aluminium oxide and **c** cellulose, which are used to coat glass and plastic for TLC

Paper chromatography and TLC are not limited to being used with coloured compounds. **Fluorescent** TLC plates can be used under ultraviolet (UV) light to see the components of a sample that would not otherwise be visible. These are shown as brighter areas on the plate where the sample emits fluorescence (Figure 6).

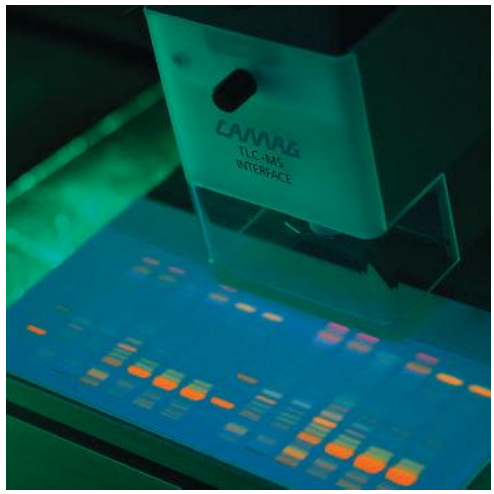


FIGURE 6 TLC performed on fluorescent plates

fluorescent
emitting visible light when exposed to radiation

Two-dimensional paper or thin-layer chromatography

When components interact similarly with the mobile or stationary phases, a single paper chromatography or TLC run may not separate all components sufficiently. Rather than repeating the analysis with new mobile and stationary phase conditions, chemists rotate the chromatogram by 90°, so that the sample is on the bottom of the sheet. They rule a new origin line and run the analysis again, using a different mobile phase with different properties (Figure 7). The chromatogram is analysed by using R_f calculations for analysis 1 and for analysis 2. This is called **two-dimensional paper or thin-layer chromatography**.

two-dimensional paper or thin-layer chromatography
an analytical technique used to separate components; a first paper or thin-layer chromatography analysis is completed, and the sheet is rotated 90° and a second analysis is run using a different mobile phase

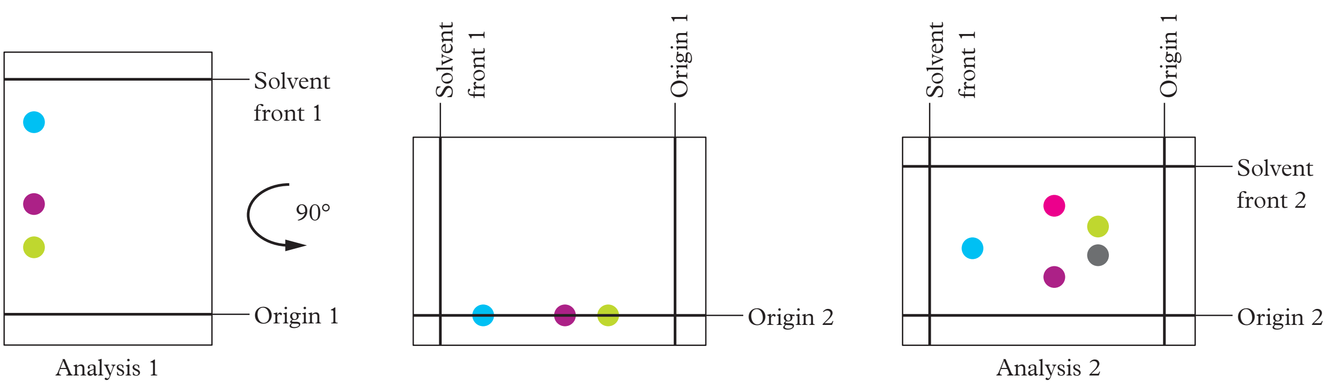


FIGURE 7 In two-dimensional paper chromatography, the chromatogram is rotated 90° after the first analysis. This gives a new origin line. The second analysis separates the green and purple dots further into their individual components.

Column chromatography

column chromatography a technique used to separate and purify individual components from mixtures of compounds

separation science the science of separating a mixture on the basis of the different properties of its components

column the tube-like structure that contains the stationary phase in column chromatography, through which the mobile phase and sample flow

elute come out of the bottom of a chromatography column

retention time (R_t) the time that a component is retained by a chromatography column

Column chromatography is a more advanced form of **separation science** than paper and thin-layer chromatography. It uses the same basic principles, but the stationary phase is contained within a tube that is open at both ends.

The simplest form of column chromatography uses a glass **column** (Figure 8), which contains a stationary phase of aluminium oxide (alumina) or silica. This resembles finely ground sand. The small size of the particles provides more surface area for adsorption to occur.

The stationary phase must also be packed tightly to minimise pore spaces (gaps). Otherwise, the sample will move too fast and will not interact with the stationary phase. This means that less separation will occur.

FIGURE 8 Column chromatography in a laboratory

The steps of column chromatography are summarised in Figure 9.

1 The mobile phase is poured into the top of the column and allowed to run slowly down through the stationary phase until it reaches the tap at the bottom.

2 Once the column is soaked in the mobile phase, the sample is added to the top of the column and the tap at the bottom is opened. This allows the mobile phase to run through the column with the sample.

3 The sample separates into its various components depending on their affinities for the mobile or stationary phases. A higher affinity for the mobile phase means that a component spends longer interacting with it. As the mobile phase flows through the column, the component moves with it and is collected earlier. More mobile phase is constantly added into the column.

4 The various components of the sample are collected in separate beakers as they **elute** (come out of) the column. Once the components are separated, they are identified and quantified by various techniques.

The time that it takes a sample to move through the column, from loading to detection, is called the **retention time (R_t)**. It indicates the amount of interaction that has occurred between each component of the sample and the stationary phase.

Study tip

Do not confuse R_t with R_f. In column chromatography, a low R_t means the substance has travelled through the column quickly and has a low affinity for the stationary phase. In paper or thin-layer chromatography, a low R_f means the substance has travelled up the stationary phase slowly and has a high affinity for the stationary phase.

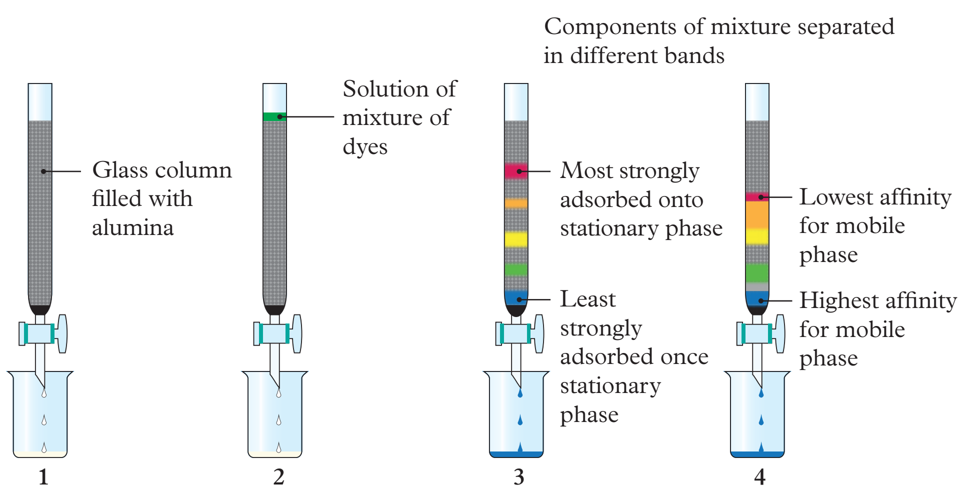


FIGURE 9 Separation of the components of a dye sample by column chromatography

The three chromatography techniques you learnt about are summarised in Table 1.

TABLE 1 A comparison of paper, thin-layer and column chromatography

Feature	Paper	Thin-layer	Column
Stationary phase	Paper (cellulose derivative)	Silica gel, aluminium oxide or cellulose coated onto a sheet of glass or plastic	Aluminium oxide or silica particles packed into a column
Mobile phase	Liquid solvent	Liquid solvent	Liquid solvent
Type of data collected	Qualitative	Qualitative	Qualitative and quantitative

Challenge yourself by trying to solve the separation problem in Challenge 6.2A.

6.2 REAL-WORLD CHEMISTRY

The horsemeat scandal

One of the biggest food scandals across Europe happened in 2013. Routine food testing in British and Irish supermarkets revealed horse DNA in beef burgers supplied by the ABP Food Group. A form of column chromatography, called **high-performance liquid chromatography (HPLC)**, combined with **mass spectrometry**, was used to analyse DNA in the burgers. You will learn about mass spectrometry in Chapter 7 and HPLC in Units 3 and 4. The results were then compared with regular beef products.

Scientists found that 37% of the burgers tested contained horse DNA and 85% contained pig DNA. Some burgers or other meat products contained 100% horse meat, whereas other tests revealed it to be present in only small percentages. Although these meats are not harmful and no health issues were raised, it created major problems of trust for consumers. The scandal caused nearly 10 million beef patties to be recalled and removed from supermarket shelves.

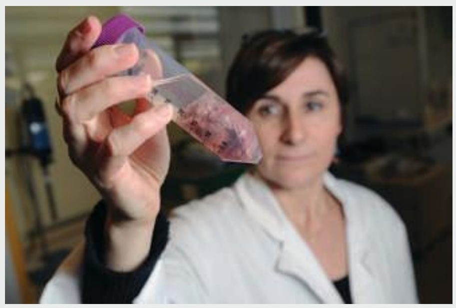


FIGURE 10 HPLC showed that some beef burgers in the UK contained horse and pig meat.

Apply your understanding

1 Explain why it was important to compare the burger DNA with beef DNA.
2 Identify one way that the scientists could increase the reproducibility of these tests.
3 Identify one way that the scientists could increase the accuracy of these tests. Hint: Think about where the meat has come from.

6.2A Challenge
Find me in your obook pro

Study tip

If R_f = 1.0, then no separation from the mobile phase has occurred. A typical exam question may ask how to separate a component in a further analysis. Look at the properties of the stationary phase and determine what you could do to make the substance more attracted to it.

high-performance liquid chromatography (HPLC)
a form of column chromatography in which the mobile phase is pumped through a column at a controlled flow rate; the particles of the stationary phase are much smaller than in column chromatography and are densely packed to make separation more sensitive and efficient

mass spectrometry
an analytical technique in which a sample is bombarded with electrons to form charged fragments; the fragments are analysed and put together like a puzzle to identify the substances present

6.2B CHALLENGE

Improving a separation

A chemist runs a mixture of dyes through a chromatography column to separate them. The chemist knows the following information about the sample and separation.

- In the sample, all dyes have a concentration of 10 ppm.
- The mobile phase is highly polar.
- The stationary phase is highly non-polar.

The samples collected as they elute from the column are shown in Table 2.

TABLE 2 Samples in a mixture of dyes

Component	Retention time (min)	Colour
A	0.5	Red
B	1.5	Green
C	3.2	Orange

The chemist determines that the green component (B) is a mixture of two dyes.

1 Evaluate the validity of the data.

2 Discuss how the chemist could separate the mixture in component B. Be specific about the properties of the mobile and stationary phases.

3 Assuming that the green component is made of a more polar blue dye and less polar yellow dye, identify the dye you would expect to have a lower retention time in the separation from your answer to Question 2.

6.2 CHECK YOUR LEARNING

Describe and explain

1 Using Figure 11, describe the steps involved in calculating retardation factor.

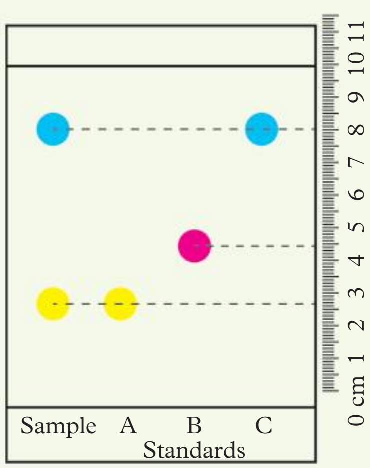


FIGURE 11 A chromatogram from analysing a food dye against three standard dyes

2 Explain how the components of a mixture are identified by using paper or thin-layer chromatography.

Apply, analyse and compare

The chromatogram in Figure 12 relates to Questions 3 and 4.

3 Analyse the chromatogram in Figure 12.

a Calculate the R_f value for every component.

b Identify the component with a higher affinity for the mobile phase. Justify your response.

c Identify the component with a higher affinity for the stationary phase. Justify your response.

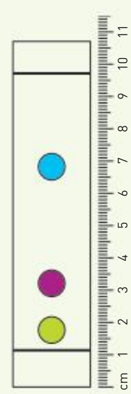


FIGURE 12 A chromatogram

4 If the chromatogram in Figure 12 was obtained by using a polar stationary phase, describe the types of intermolecular bonding that may be present between each component and the mobile and stationary phases.

Design and discuss

5 The R_f values of 10 food dyes are shown in Table 3. These R_f values are specific for TLC with a silica plate and a 1% ethanol mobile phase.

TABLE 3 The R_f values of some food dyes in 1% ethanol

Dye	R_f
Brilliant blue FCF	0.13
Indigotine	0.15
Fast green FCF	0.38
Erythrosine	0.40
Quinoline yellow	0.50
Carmoisine	0.60
Tartrazine	0.65
Green S	0.72
Patent blue V	0.88
Ponceau 4R	0.91

Three foods A–C were tested to determine which food dyes were present. The resulting chromatogram is shown in Figure 13.

a Calculate the R_f values of the components of each sample and determine which food colourings were present in the foods.

b There are two components of these samples that look the same. Evaluate whether these components are the same and explain your answer.

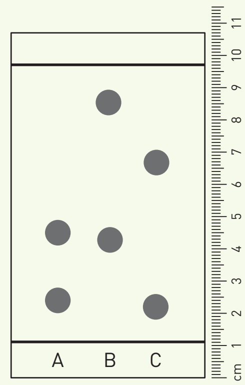


FIGURE 13 A chromatogram of food dyes A–C

6 A note left at a crime scene was written in black pen. Two black pens found at the scene were analysed by forensic scientists to determine which pen had written the note. To ensure valid results, five teams were given one of the two pens to test. Each team developed a method that involved using paper chromatography with a water mobile phase. The data in Table 4 was obtained.

a Use the data to calculate the R_f values of components 1–3.

b Determine which of the teams analysed the same pens.

c Determine which teams analysed the pen that was responsible for the note at the crime scene.

d Discuss whether component 1 is the same substance for both pens.

e Evaluate whether it is possible to identify the components of the sample on the basis of this analysis.

TABLE 4 Results from paper chromatography analysis of pen types

Team	Distance (cm) travelled by:			
	Solvent front	Component 1	Component 2	Component 3
1	11	6.16	8.69	2.31
2	15	4.80	8.85	13.05
3	9	5.04	7.11	1.89
4	21	11.76	16.59	4.41
5	18	5.76	10.62	15.66
Crime scene note	25	8.00	14.75	21.75

CHAPTER

6 Review

Chapter summary

6.1

- *Like dissolves like!* Polar solutes will dissolve in and have affinity for polar solvents. Non-polar solutes will dissolve in and have affinity for non-polar solvents.
- Chromatography separates mixtures by using a mobile phase and a stationary phase.
- Polar character determines the intermolecular bonds (dipole–dipole attractions, hydrogen bonding, dispersion forces) that can form between the sample components, mobile phase and stationary phase.
- Components adsorb to the stationary phase and desorb into the mobile phase. Depending on their affinity for the mobile and stationary phases, components will move faster or slower through the mobile phase and separate from a mixture.

6.2

- Paper, thin-layer and column chromatography are used to separate and identify components and impurities in mixtures.
- The chromatograms obtained from paper and thin-layer chromatography can be used to calculate a retardation factor (R_f). Both techniques are qualitative.
- Column chromatography uses a tube-like structure that is packed with a stationary phase. The time taken for components to move through the column, or their retention times (R_t), is used to identify them (qualitative). The amount of the components can also be measured (quantitative).

Key formulas

Retardation factor (R_f)	$R_f = \frac{\text{distance of sample dot}}{\text{distance of mobile phase}}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Sort of’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• explain how the polarity of molecules affects their solubility in polar and non-polar solvents	☐	☐	☐	Go back to Topic 6.1
• describe the intermolecular forces involved in separation science, including dispersion forces, dipole–dipole attractions and hydrogen bonding	☐	☐	☐	Go back to Topic 6.1
• describe the basic principles of chromatography, including adsorption, desorption, and affinity for the mobile and stationary phases	☐	☐	☐	Go back to Topic 6.1
• describe how the composition and purity of different types of substances can be determined by different types of chromatography, including paper chromatography, thin-layer chromatography and column chromatography	☐	☐	☐	Go back to Topic 6.2
• calculate R_f values	☐	☐	☐	Go back to Topic 6.2

Revision questions

Multiple choice

1 Which of the following is incorrect? Water:

A has a dipole.

B interacts with other water molecules through hydrogen bonding.

C dissolves some ionic substances.

D does not form dispersion forces.

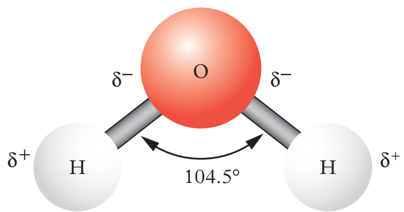


FIGURE 1 The water molecule

2 A polar solute will have high affinity for a:

A polar mobile phase but not a polar stationary phase.

B polar mobile phase and a polar stationary phase.

C non-polar mobile phase and a polar stationary phase.

D polar mobile phase and a non-polar mobile phase.

3 The order of intermolecular forces from strongest to weakest is:

A dispersion forces, dipole–dipole attractions, hydrogen bonding.

B hydrogen bonding, dispersion forces, dipole–dipole attractions.

C dipole–dipole attractions, hydrogen bonding, dispersion forces.

D hydrogen bonding, dipole–dipole attractions, dispersion forces.

4 Column chromatography was performed to identify the pigments within a plant sample. The results are shown in Figure 2. What conclusions could be drawn about the blue pigment?

A It has a high R_t and a high affinity for the stationary phase.

B It has a low R_t and a high affinity for the stationary phase.

C It has a high R_t and a high affinity for the mobile phase.

D It has a low R_t and a high affinity for the mobile phase.

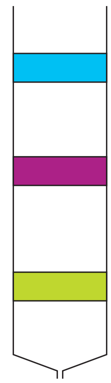


FIGURE 2 Separation of pigments in a plant sample

5 Which of the following would result in the highest R_f value?

	Mobile phase	Stationary phase	Substance
A	polar	non-polar	polar
B	non-polar	polar	polar
C	polar	non-polar	non-polar
D	non-polar	polar	polar

6 R_f can be calculated by:

A adding the distance of the solvent front to the distance moved by the solute.

B determining the difference between the distance of the solvent front and the distance moved by the solute.

C dividing the distance moved by the solute by the distance of the solvent front.

D determining the distance moved by the solute.

7 Which chromatographic techniques are qualitative as well as quantitative?

A Paper chromatography, TLC and column chromatography

B Paper chromatography and TLC

C Column chromatography only

D None of the above

8 In a mixture of solutes, the following R_f values were calculated based on a polar mobile phase and a non-polar stationary phase. Use the results to choose the correct answer below.

Solute	R_f
A	0.50
B	0.72
C	0.92

A A is the least polar and desorbs the most from the stationary phase.

B C is the most polar and has the strongest affinity for the mobile phase.

C A is the most polar and desorbs the least from the stationary phase.

D C is the least polar and has the strongest affinity for the mobile phase.

9 In a TLC separation to determine the amino acids in a seaweed extract, a student determines that two amino acids have the same R_f. To fully separate these molecules the best option would be to:

A turn the separation sideways and change the properties of the mobile phase for two-dimensional chromatography.

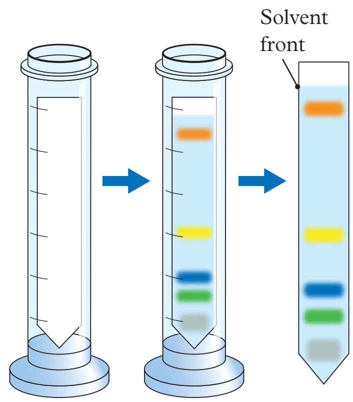
B run the separation again with a more polar mobile phase.

C run the separation again with a more polar stationary phase.

D run the separation again with paper chromatography.

FIGURE 3 Seaweed extract

Figure 4 relates to Questions 10–12. It shows the separation of a leaf stain into its individual pigments. The R_f value is calculated from the chromatogram.



Pigment	R_f value
Carotene	0.95
Xanthophyll	0.48
Chlorophyll a	0.30
Chlorophyll b	0.18
Leaf stain	

FIGURE 4 Separation of a leaf stain into its individual pigments

10 Which solute has the greatest adsorption to the stationary phase?

A Green

B Blue

C Yellow

D Orange

11 Which solute has the strongest affinity for the mobile phase?

A Green

B Blue

C Yellow

D Orange

12 Assuming that the mobile phase is polar, which solute has the strongest intermolecular forces?

A Green

B Blue

C Yellow

D Orange

Short answer

Describe and explain

13 Which type of intermolecular bonding is the strongest? Explain why.

14 Explain how non-polar molecules are able to interact if they have no overall charges.

15 Explain how substances separate from a mixture during chromatography. Use the terms 'mobile phase', 'stationary phase', 'affinity', 'adsorb' and 'desorb' in your answer.

16 Describe how chromatography can be used to identify impurities in a substance.

17 Why is it necessary to prepare a set of standards when performing qualitative analysis in chromatography?

18 Explain why an R_f value can never be 0 or 1 in a valid separation.

19 Explain how components can be separated if they overlap in TLC or paper chromatography.

20 Explain how the components can be separated if they elute from a chromatography column at the same time.

Apply, analyse and compare

21 A separation was conducted to determine the pigments in a sample of food colouring. The chromatogram included the standards of food dye that the manufacturer claims is in the food product. Figure 5 shows the chromatogram of the pigment and standard dyes A–C.

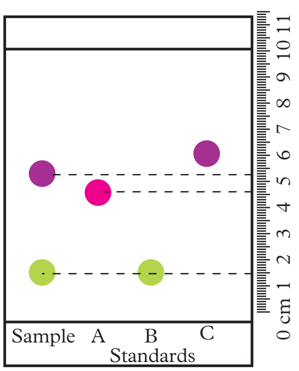


FIGURE 5 Separation of pigments in food and comparison to standard dyes

a Calculate the R_f values for the solutes in the sample.

b Identify the standard dye with the highest affinity for the polar mobile phase.

c Identify the standard dye that desorbs the least from the stationary phase.

d Determine which standard dyes are in the food sample.

e Evaluate the manufacturer's claims regarding the dyes in the food.

22 Figure 6 shows a separation conducted to determine the types of ink in a sample of marker pen. The stationary phase is paper and the mobile phase is 10% ethanol in water.

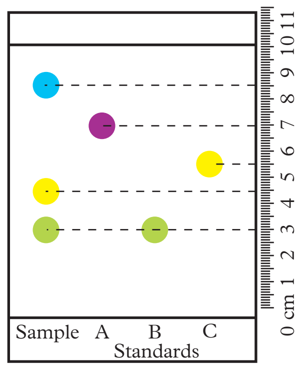


FIGURE 6 Separation of dyes in a marker pen and comparison to standard dyes

a Calculate the R_f values for the solutes in the sample.

b Identify the polarity of the mobile and stationary phases. Justify your reasoning.

c Identify the standard dye with the strongest solute–solvent intermolecular interactions.

d Identify the standard dye that adsorbs the strongest to the stationary phase.

e Determine which standard dyes are in the marker.

f Describe how the separation can be improved to ensure that the identity of all dyes is determined.

23 The following R_f values were calculated from a sample that has been separated using paper chromatography.

Component	R_f	Properties
A	0.83	Most polar
B	0.74	
C	0.54	
D	0.32	

a Calculate the distance that each solute moved from the origin, assuming that the solvent front travelled 17 cm.

b Identify the polarity of the mobile and stationary phases. Justify your reasoning.

c Identify the component with the highest affinity for the mobile phase.

d Identify the component with the strongest adsorption to the stationary phase.

24 Figure 7 shows the results of a separation conducted by column chromatography. The column is filled with fine beads that are coated in a polar stationary phase. The mobile phase is non-polar.

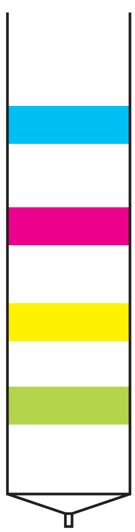


FIGURE 7 Separation of the solutes in a mixture by column chromatography

a Identify what a researcher must do with each solute as it elutes from the column.

b Identify whether the separation itself is qualitative, quantitative or neither. Justify your answer.

c Identify the solute with the greatest affinity for the mobile phase. Justify your answer.

d Identify the solute with the greatest affinity for the stationary phase. Justify your answer.

e Identify the solute with the highest retention time. Explain why this is the case by referring to the solute's properties.

f Identify the solute with the greatest desorption from the stationary phase.

Design and discuss

25 A column chromatography analysis was performed to determine the R_t of a polar substance, using a polar mobile phase and a non-polar stationary phase. Discuss the effects of decreasing the polarity of the mobile phase on the R_t of the substance.
Draw a separation, similar to that in Figure 7, which demonstrates this effect.

26 Using a non-polar stationary phase, a scientist performs a TLC analysis on amino acids labelled with fluorescent dyes.
The scientist believes that the amino acids present in the sample are those shown in Figure 8.

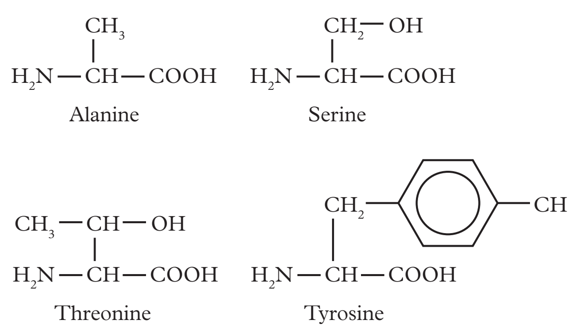


FIGURE 8 Amino acids analysed in a TLC analysis

a Write a hypothesis about the order of the amino acids, from highest R_f to the lowest. Explain your answer.

b Identify the reason why the scientist added fluorescent dyes to the amino acids.

c The scientist finds that two amino acids are overlapping on the chromatogram. Predict which two may not be separating and discuss your choice.

d Design a method to separate the two overlapping components of the mixture and discuss the results that you predict will be obtained.

27 Design a method to analyse a mixture of three alcohols. Each alcohol has only one hydroxyl group and no more than 10 carbon atoms.

In your answer, you must outline or predict:

a the properties of the mobile phase

b the properties of the stationary phase

c the order in which the alcohols will separate

d which alcohol has a higher affinity for the mobile and stationary phases

e how the alcohols can be identified in the separation.

FIGURE 9 A beaker containing three alcohols

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 1

AREA OF STUDY 1

Checkpoint

Multiple choice

Question 1

Some properties of metals are:

I ductility
II electrical conductivity
III magnetism
IV malleability.

The structure of metals is described as a 'lattice of cations in a sea of electrons'.

Which of the properties I–IV can be explained by this simple model?

A I and II only
B I, II and III only
C I, II and IV only
D II, III and IV only

Question 2

Gold (Au) can form the ionic compounds AuCl and $AuCl_3$. The correct name for $AuCl_3$ is:

A gold trichloride.
B gold chloride.
C gold(I) chloride.
D gold(III) chloride.

Question 3

The electronic configuration of a particle is $1s^22s^22p^63s^23p^63d^8$. This particle is most likely to be:

A an atom of iron in its ground state.
B an atom of cobalt in an excited state.
C a double-charged nickel cation.
D a double-charged chromium anion.

Question 4

Some properties of the elements that change across a period in the periodic table are:

I electronegativity
II atomic radius
III metallic character.

Moving across period 3 from sodium to chlorine, which of these properties increase?

A I only
B I and II only
C II and III only
D I, II and III

Question 5

A substance melts at 2300°C and does not conduct electricity under any conditions. The substance is most likely to be:

A Au
B SiO_2
C CH_4
D N_2

Question 6

Which of the following would result in the lowest R_f value?

	Mobile phase	Stationary phase	Substance
A	Polar	Non-polar	Polar
B	Non-polar	Polar	Polar
C	Polar	Polar	Non-polar
D	Non-polar	Polar	Non-polar

Question 7

A mixture of three solutes is separated using paper chromatography. It has a polar mobile phase and a non-polar stationary phase. The R_f values are calculated and are shown in the table.

Solute	R_f
A	0.96
B	0.72
C	0.43

Based on the results, identify the correct statement.

A A is the most polar and is attracted to the stationary phase.

B C is the least polar and is attracted to the stationary phase.

C A is the least polar and is attracted to the mobile phase.

D C is the most polar and is attracted to the mobile phase.

Question 8

The number of electrons and the mass numbers of a series of fictitious particles are shown in the table.

Particle	Number of electrons	Mass number
U	12	24
V^-	10	19
W^+	10	23
X	10	20
Y^{2+}	10	24
Z	10	22

Which two particles are isotopes?

A X and Z

B U and Y

C W and Y

D V and X

Question 9

Which of the following statements describes the bonds in a molecule of CCl_4 and the overall polarity of the whole molecule?

A Polar bonds and polar molecule

B Non-polar bonds and polar molecule

C Polar bonds and non-polar molecule

D Non-polar bonds and non-polar molecule

Question 10

Which of the following is *not* a property of an ionic solid?

A Conducts electricity when molten

B Has a crystalline structure

C Has a high melting temperature

D Is malleable

Short answer

Question 1 (6 marks)

The following table lists properties of substances.

A	Conducts electricity
B	Is very soluble in water
C	Has covalent bonding
D	Has a lattice network structure
E	Has hydrogen bonding
F	Has dispersion forces

a List the letters that correspond to any properties that are most applicable to diamond. 2 marks

b List the letters that correspond to any properties that are applicable to an ammonia molecule but not a methane molecule. 2 marks

c List the letters that correspond to any properties that are most applicable to the gases hydrogen and carbon dioxide. 2 marks

Question 2 (5 marks)

A freshly cut piece of potassium reacts violently in a gas jar containing chlorine gas. A white smoke consisting of potassium chloride is formed and eventually deposits as a white solid in the gas jar.

a Write a balanced equation for the reaction between potassium and chlorine gas. 2 marks

b Explain the changes, in detail, that occur in both the potassium and chlorine atoms during the reaction. Use electron dot diagrams to illustrate your explanation. 3 marks

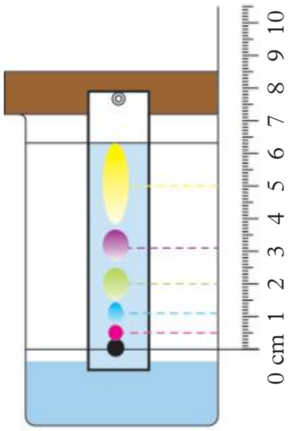
Question 3 (6 marks)

Draw a structural formula (showing all lone/non-bonding pairs on the central atom) for each of the following molecules and state whether it is polar or non-polar and the type of intermolecular bonding.

Molecule	Structural formula	Polar or non-polar	Type of intermolecular bonding
CCl_4			
H_2S			
NH_3			

Question 4 (4 marks)

The following paper chromatogram was produced from the ink of a black pen. The stationary phase was paper and the mobile phase was 15% v/v ethanol in water.



a Identify the polarity of the mobile phase. 1 mark

b Calculate the R_f of the colour with the lowest affinity for the mobile phase. 1 mark

c Identify the colour with the highest affinity for the mobile phase and explain your choice, with reference to intermolecular forces. 2 marks

Question 5 (4 marks)

Write balanced full and ionic equations to represent:

a dilute sulfuric acid being added to magnesium ribbon and a colourless, odourless gas evolving 2 marks

b the formation of cloudiness when carbon dioxide is bubbled through limewater $Ca(OH)_2$ solution. 2 marks

Question 6 (6 marks)

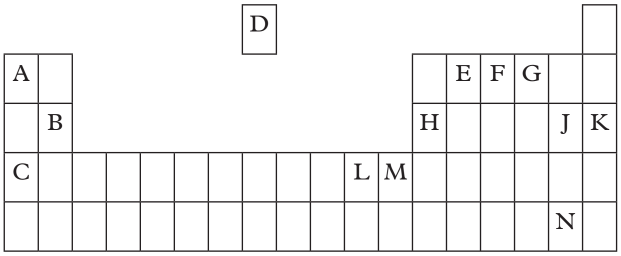
Copy and complete the following table.

Name of compound	Formula of compound
Ammonium acetate	
Copper(II) hydrogen carbonate	
Aluminium sulfite	
	MgH_2
	PbO_2
	$Fe_2(HPO_4)_3$

6 marks

Question 7 (9 marks)

The diagram represents the modern periodic table with letters of the alphabet replacing the symbols of elements.



a Identify the letter that represents:

i a non-metal with four electrons in its outer shell

ii an element in group 13

iii a transition element with nine electrons in the 3d orbital

iv the element with the highest electronegativity

v an element in period 3 with a total of 12 electrons.

5 marks

b Atoms of two of these elements can react to form a compound. For the following examples, write the chemical formula (using the letters) and state the type of bonding.

i F and C

ii E and D

2 marks

c What type of bonding would exist if B and H were mixed in a 3:2 ratio? 1 mark

d State the electron shell and subshell configurations of the anion formed from G.

1 mark

TOTAL MARKS

/50 marks

CHAPTER 7

Quantifying atoms and compounds

KEY KNOWLEDGE

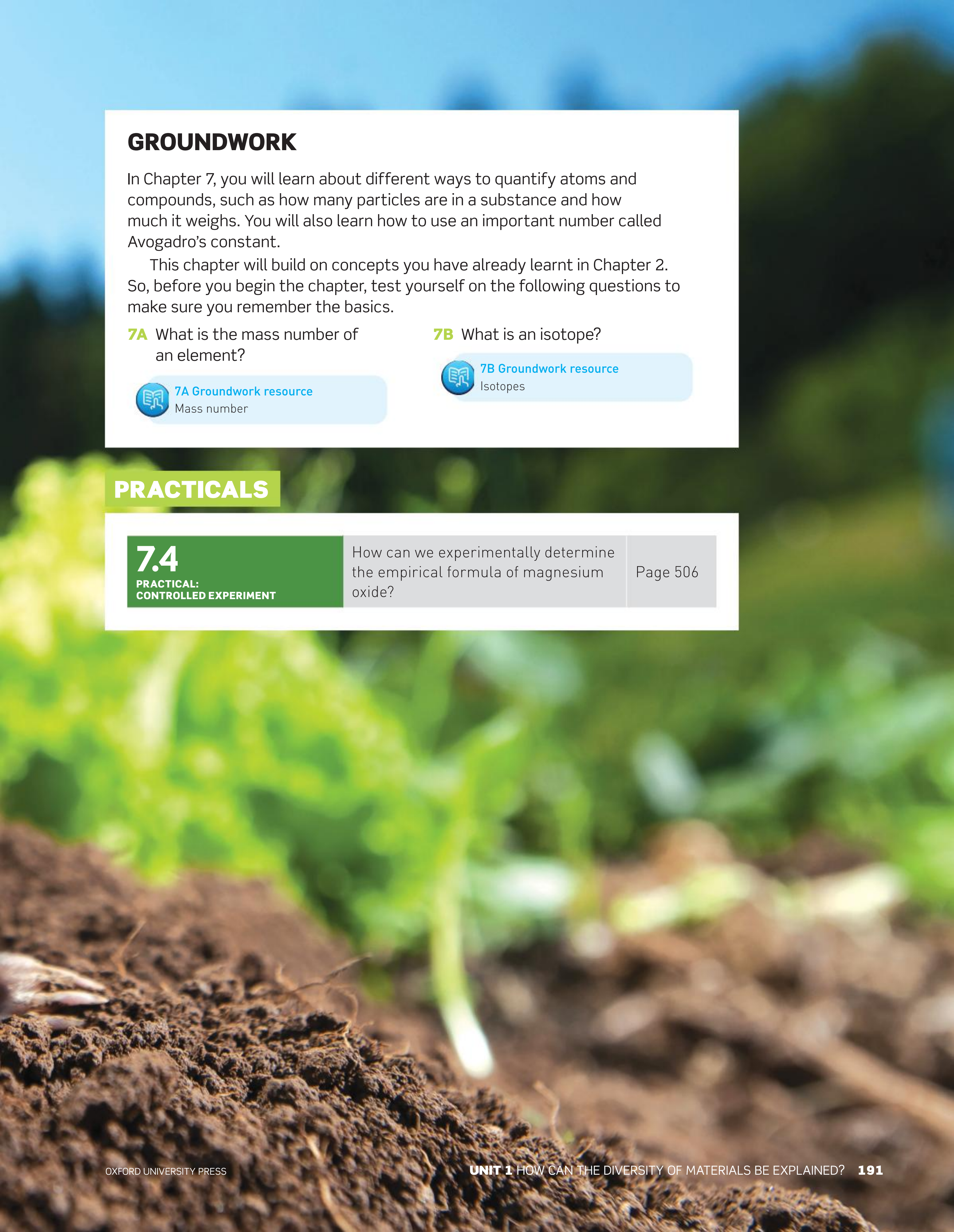
→ the relative isotopic masses of isotopes of elements and their values on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly

→ determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)

→ Avogadro's constant as the number 6.02×10^{23} indicating the number of atoms or molecules in a mole of any substance; determination of the amount, in moles, of atoms (or molecules) in a pure sample of known mass

→ determination of the molar mass of compounds, the percentage composition by mass of covalent compounds and the empirical and molecular formula of a compound from its percentage composition by mass

FIGURE 1 A mole is a type of burrowing animal, as well as a unit used in Chemistry to measure the number of atoms, molecules, ions or electrons.



GROUNDWORK

In Chapter 7, you will learn about different ways to quantify atoms and compounds, such as how many particles are in a substance and how much it weighs. You will also learn how to use an important number called Avogadro's constant.

This chapter will build on concepts you have already learnt in Chapter 2. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

7A What is the mass number of an element?

7A Groundwork resource
Mass number

7B What is an isotope?

7B Groundwork resource
Isotopes

PRACTICALS

7.4 PRACTICAL: CONTROLLED EXPERIMENT	How can we experimentally determine the empirical formula of magnesium oxide?	Page 506

7.1 Relative isotopic mass

KEY IDEAS

In this topic, you will learn that:

- the mass number of an atom (its number of protons + neutrons) is only an approximation of an atom's mass
- the relative isotopic mass (RIM) of an atom is a more accurate way to measure an atom's mass
- relative isotopic mass (RIM) is the mass of an atom compared with the mass of a carbon-12 atom, which has a RIM of 12 exactly.

In Chapter 2, we learnt that the mass number of an atom is calculated by adding the total number of protons and neutrons in the nucleus. For example, a sodium atom with 11 protons and 12 neutrons in its nucleus has a mass number of 23. (The mass of an electron in an atom is considered negligible because it is so small.)

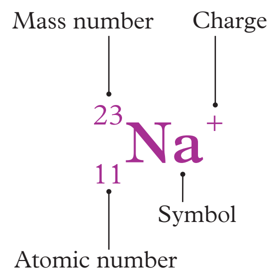


FIGURE 1 Nuclide notation includes a mass number, which is a count of the total number of nucleons (protons + neutrons) in the nucleus.

But there is a more accurate way to find the mass of an atom, called **relative isotopic mass (RIM)**.

relative isotopic mass (RIM) the ratio of the average mass of one atom of an isotope to one-twelfth of the mass of an atom of carbon-12; RIM has no units and has a value equal to the mass of the atom in Da or u

Why use relative isotopic mass?

Mass numbers are only a rough calculation of an atom's mass. Consider the carbon-14 and nitrogen-14 atoms (Figure 2). They both have mass numbers of 14 because they each have 14 nucleons (protons and neutrons). But their actual masses, or relative isotopic masses are different. Each isotope has its own relative isotopic mass.

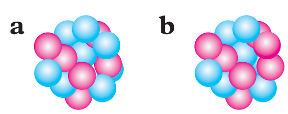


FIGURE 2 The nucleons of **a** carbon-14 and **b** nitrogen-14

This is because:

- protons and neutrons have very slightly different masses (neutrons have approximately 0.14% more mass than protons)
- electrons also have mass, which is not included in a mass number
- the binding energy of the nucleus (and other factors) make an isotope's relative isotopic mass slightly different from its mass number.

Once these factors are considered, it is possible for atoms with exactly the same mass numbers to have slightly different relative isotopic masses. See Table 1 for some examples.

Study tip

Remember, isotopes are atoms of an element that have the same number of protons, but a different number of neutrons. You can revise this in Chapter 2.

TABLE 1 Atoms with the same mass numbers can have different relative isotopic masses because mass numbers are only approximations of mass.

Atom name	Nuclide notation	Mass number	RIM	Mass in grams
Carbon-14	${}^{14}_{6}C$	14	14.003 241	$2.325\ 293 \times 10^{-23}$
Nitrogen-14	${}^{14}_{7}N$	14	14.003 074	$2.325\ 265 \times 10^{-23}$
Palladium-108	${}^{108}_{46}Pd$	108	107.903 89	$1.791\ 786 \times 10^{-22}$
Cadmium-108	${}^{108}_{48}Cd$	108	107.904 18	$1.791\ 791 \times 10^{-22}$

Therefore, the mass number of an isotope is a whole-number approximation of its relative isotopic mass.

Daltons (Da) and atomic mass units (u)

Relative isotopic mass is measured on a standard scale using **carbon-12**. The relative isotopic mass of the carbon-12 isotope is assigned a value of exactly 12.

carbon-12 (${}^{12}C$)
an isotope of carbon with 6 protons, 6 neutrons and 6 electrons

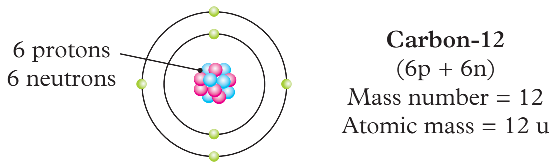


FIGURE 3 The relative isotopic mass of the carbon-12 isotope is assigned a value of exactly 12.

Carbon-12 was chosen as the standard for RIM because it:

- is more stable than other carbon isotopes, such as carbon-13 and carbon-14
- can be cheaply and widely obtained
- is safer and has low toxicity to (most) living organisms
- is abundant and found easily in nature.

Before carbon-12 was adopted as a standard, other isotopes such as oxygen-16 and hydrogen-1 were used, but these were abandoned in 1961 in favour of carbon-12.

So, relative isotopic mass (RIM) can be described as the mass of an isotope atom, relative to the mass of an atom of carbon-12 (exactly 12). It can also be described as:

$$\text{RIM} = \frac{\text{mass of an isotope atom}}{\frac{1}{12}\text{mass of a carbon-12 atom}}$$

FIGURE 4 Activated charcoal (carbon) is safe to ingest and is used for treating some forms of poisoning.

dalton (Da)
a unit of mass where one atom of carbon-12 has an mass of exactly 12 Da

atomic mass unit (u)
a unit of mass where one atom of carbon-12 has a mass of exactly 12 u

Study tip

Relative isotopic mass (RIM) has no units, but its numerical value is the same as the mass of the isotope in daltons (Da) or atomic mass units (u).

The mass of atoms can also be measured in **daltons (Da)** and **atomic mass units (u)**. The values of these three measurements are all identical: the only difference between them is the units. See Table 2 for some examples.

TABLE 2 Relative isotopic masses (RIM) of some isotopes of hydrogen and carbon compared with the mass of each atom in daltons and atomic mass units. Relative isotopic mass equals the mass measured in daltons or atomic mass units

Isotope	Nuclide notation	RIM (no units)	Mass in daltons (Da)	Mass in atomic mass units (u)
Hydrogen-1	$^{1}_{1}H$	1.007 825	1.007 825	1.007 825
Deuterium	$^{2}_{1}H$	2.014 102	2.014 102	2.014 102
Tritium	$^{3}_{1}H$	3.016 049	3.016 049	3.016 049
Carbon-12	$^{12}_{6}C$	12 exactly	12 exactly	12 exactly
Carbon-13	$^{13}_{6}C$	13.003 354	13.003 354	13.003 354
Carbon-14	$^{14}_{6}C$	14.003 241	14.003 241	14.003 241

7.1 SKILL DRILL

Comparing relative isotopic masses

Key science skill: Analyse and evaluate data and investigation methods

Practise your skills

Use the data in Table 2 to answer these questions.

1 List the number of protons, neutrons and electrons in a:
 a deuterium atom
 b carbon-12 atom.
2 How many times greater is the relative isotopic mass of carbon-14 than the relative isotopic mass of carbon-12? Round your answer to five significant figures.
3 Why is the relative isotopic mass of a carbon-12 atom *not* exactly six times greater than the relative isotopic mass of a deuterium atom?

FIGURE 5 RIM is relative to the mass of carbon-12.

4 How many times greater is the relative isotopic mass of tritium than the relative isotopic mass of hydrogen-1? Round your answer to five significant figures.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

7.1 CHECK YOUR LEARNING

Describe and explain

1 Define 'relative isotopic mass'.
2 What are the units of relative isotopic mass?

Apply, analyse and compare

3 Describe the relationship between mass number and relative isotopic mass.
4 What is the difference between relative isotopic mass and an isotope's mass measured in daltons?

Design and discuss

5 Explain why carbon-12 was chosen as a standard by which to measure relative isotopic mass.

7.2 Relative atomic mass

KEY IDEAS

In this topic, you will learn that:

+ mass spectrometers can measure the relative isotopic masses of each isotope in an element sample and the relative abundance of each isotope
+ relative atomic mass (RAM) can be calculated from data from a mass spectrometer presented in graph or table form.

Most elements have a number of isotopes. It's not always possible, or practical, to know how much of each isotope is in an element sample. So, chemists use **relative atomic mass (RAM)**. Relative atomic mass is a weighted average (mean) of isotopic mass.

Weighted mean is different from mean because it takes the percentage abundance of each isotope into account. RAMs are listed in the periodic table in the VCE Chemistry databook.

relative atomic mass (RAM) the weighted mean of all the relative isotopic masses of an element

Mass spectrometry

The abundance of different isotopes can be determined by using a **mass spectrometer**.

The mass spectrometer separates isotopes in an element sample and presents the:

- number of isotopes in the element sample
- relative isotopic mass of each isotope
- relative abundance of each isotope.

Explore how mass spectrometry was used to determine if life exists on Mars in Real-world chemistry 7.2.

mass spectrometer an apparatus for measuring the masses and relative abundances of isotopes (and molecules and molecular fragments) by ionising them and determining their trajectories in electric and magnetic fields

Study tip

Be careful not to confuse relative isotopic mass (RIM) and relative atomic mass (RAM).

7.2 Real-world chemistry
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77	78	79
Ir	**Pt**	**Au**
192.2	195.1	197.0
Iridium	Platinum	Gold

FIGURE 1 Relative atomic masses are written under each element's symbol on the periodic table.

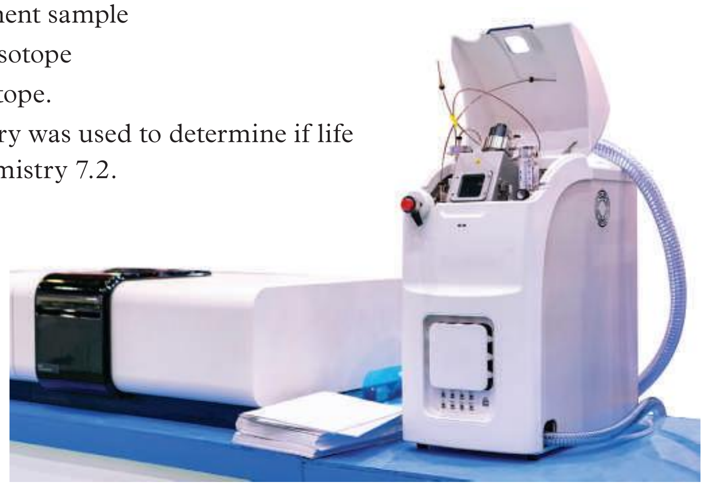

FIGURE 2 A mass spectrometer is a laboratory device for analysing the composition of a sample.

Interpreting a mass spectrum

A **mass spectrum** (plural 'spectra') is a column graph that shows the relative abundance of each isotope in a sample of an element. Relative abundance compares the abundance of each isotope with each other, so that the most abundant isotope has a value of 100%. Figure 3 (on the next page) shows the mass spectrum for a sample of selenium containing six different isotopes.

The *y*-axis shows the relative abundance of each isotope; the most common isotope is set at a value of 100%. Therefore, the sum of the relative abundances can total more than 100%, and we should convert them before using them in any calculations.

mass spectrum a column graph that shows the relative abundance of each isotope in a sample of an element

m/z
the mass-to-charge ratio; the *x*-axis on a mass spectrum; charge is +1 so *m/z* represents the relative isotopic mass of each isotope

The *x*-axis shows the mass-to-charge ratio (***m/z***) of each isotope. The charge in each instance is 1+ so the values on the *x*-axis are equal to relative isotopic mass.

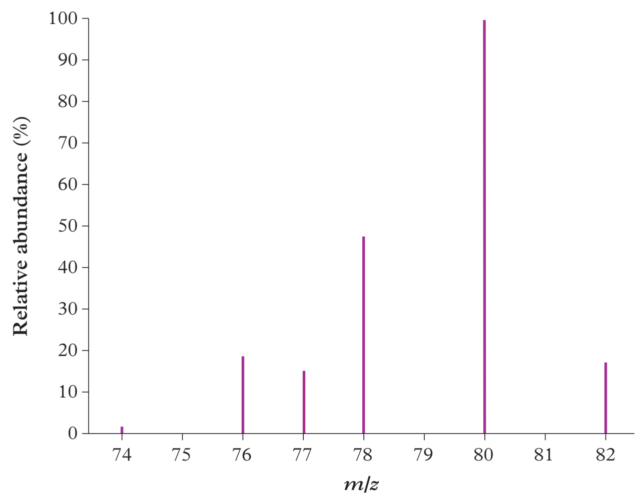

FIGURE 3 The mass spectrum of selenium, showing the relative abundance of its six naturally occurring isotopes

We can interpret a mass spectrum by looking at the:

- number of isotopes in the element sample = the number of visible peaks
- relative isotopic mass of each isotope as shown on the *x*-axis (but usually only whole numbers are visible)
- **relative abundance** of each isotope as shown on the *y*-axis.

relative abundance
the abundance of each isotope, where the abundance of the most abundant isotope is set to a value of 100%; the sum of all the relative abundances of all the isotopes will be greater than 100%

percentage abundance
the abundance of each isotope, where the sum of all the percentage abundances is exactly 100%

Calculating percentage abundance from mass spectra

It is useful to first convert the mass spectrum into a data table (see Table 1).

We then need to convert each relative abundance into **percentage abundance**. This is the relative abundance of each isotope compared to all the other isotopes for that element. Percentage abundance will add up to exactly 100%.

To convert the relative abundances on a mass spectrum (which add up to more than 100%) into percentage abundances (which add up to exactly 100%), we use the following formula:

$$\%\text{ abundance of isotope} = \frac{\text{relative abundance}}{\text{sum of all the relative abundances}} \times 100$$

For example, to find the percentage abundance of ^{74}Se, we divide the relative abundance of ^{74}Se, which is 1.79%, by the sum of all the relative abundances, which is 201.57%.

$$\begin{aligned}\%\text{ abundance}\,(^{74}\text{Se}) &= \frac{\text{relative abundance}\,(^{74}\text{Se})}{\text{sum of all the relative abundances}} \times 100 \\ &= \frac{1.79}{(1.79 + 18.89 + 15.38 + 47.91 + 100 + 17.60)} \times 100 \\ &= \frac{1.79}{201.57} \times 100 \\ &= 0.0089 \times 100 = 0.890\%\end{aligned}$$

By dividing each relative abundance by the sum of all the relative abundances (which in this example is 201.57%), we can calculate a set of percentage abundances. These values have been calculated and added to Table 1.

TABLE 1 Isotopes of selenium and their relative abundances obtained from the mass spectrum for selenium and the calculations for percentage abundances

Isotope (from x-axis)	Relative abundance (from y-axis) (%)	% abundance
^{74}Se	1.79	$\frac{1.79\%}{201.57\%} = 0.89$
^{76}Se	18.89	$\frac{18.89\%}{201.57\%} = 9.37$
^{77}Se	15.38	$\frac{15.38\%}{201.57\%} = 7.63$
^{78}Se	47.91	$\frac{47.91\%}{201.57\%} = 23.77$
^{80}Se	100.00	$\frac{100.00\%}{201.57\%} = 49.61$
^{82}Se	17.60	$\frac{17.60\%}{201.57\%} = 8.73$
Total	**201.57**	**100.00**

Practise interpreting a mass spectrum using Skill drill 7.2. If you're up for an extra challenge, try Challenge 7.2A.

Calculating relative atomic mass

The RAM of an element can be calculated from the data obtained from its mass spectrum by using the formula:

RAM = sum of each (RIM × its percentage abundance)

This can also be written as:

RAM = ∑(RIM × percentage abundance)

Calculating relative atomic mass from a mass spectrum

When calculating the RAM of an element from a mass spectrum, we are only able to read the RIM values on the x-axis of a mass spectrum to the nearest whole number. This is due to limitations of reading a printed graph.

For example, the RAM of selenium can be calculated using the data from Table 1:

$$\begin{aligned}\text{RAM(selenium)} &= \sum(\text{RIM} \times \text{percentage abundance})\\ &= (74 \times 0.89\%) + (76 \times 9.37\%) + (77 \times 7.63\%) + (78 \times 23.77\%)\\ &\quad + (80 \times 49.61\%) + (82 \times 8.73\%)\\ &= (0.66) + (7.12) + (5.88) + (18.54) + (39.69) + (7.16)\\ &= 79.04\end{aligned}$$

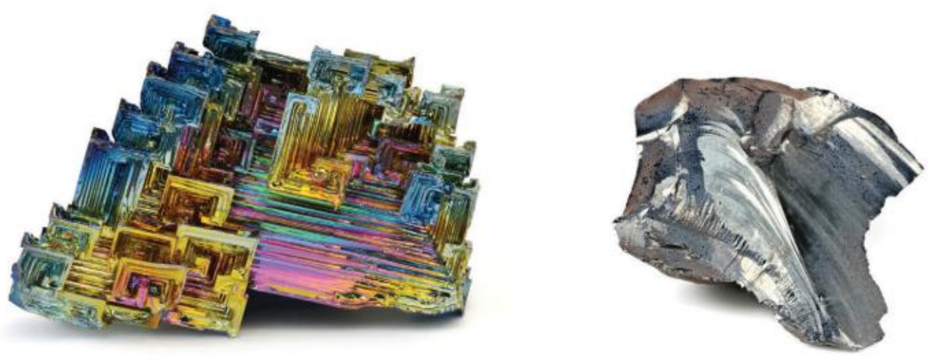

FIGURE 4 Elements (such as bismuth and selenium) can be analysed by mass spectrometry to determine their RAM.

Study tip

Percentage abundance (% abundance) and relative abundance are not the same! Percentage abundance values add up to 100%. Relative abundance values add up to more than 100%.

Study tip

When exact RIM values are not provided, then calculate RAM using estimated integer RIM equal to the mass number of each isotope.

7.2 Skill drill
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7.2A Challenge
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It is convenient to arrange the calculation in table format as shown in Table 2.

TABLE 2 Calculating the RAM of selenium

Isotope (from *x*-axis)	Assumed RIM	% abundance	RIM × % abundance
^{74}Se	74	× 0.89	= 0.66
^{76}Se	76	× 9.37	= 7.12
^{77}Se	77	× 7.63	= 5.88
^{78}Se	78	× 23.77	= 18.54
^{80}Se	80	× 49.61	= 39.69
^{82}Se	82	× 8.73	= 7.16
Total	–	**100.00**	**79.04**

Calculating relative atomic mass from mass spectrum data and RIM values

If exact RIM values are provided (not just whole-number estimates from the mass spectrum), then we can calculate a more accurate RAM for the element. For example, if we are provided with the RIM values for ^{54}Fe (53.94), ^{56}Fe (55.93), ^{57}Fe (56.94) and ^{58}Fe (57.93), then we can calculate RAM for iron accurately, as shown in Table 3.

TABLE 3 Calculating the RAM of iron using accurate RIM values as the sum of (RIM × % abundance) for each isotope

Isotope	RIM	% abundance	RIM × % abundance
^{54}Fe	53.94	5.85	3.15
^{56}Fe	55.93	91.75	51.32
^{57}Fe	56.94	2.12	1.21
^{58}Fe	57.93	0.28	0.16
Total	–	**100.00**	**55.85**

Make sure you double check your RAM

Always check your calculated RAM against the value in the periodic table. Your calculated RAM should match the value in the periodic table very closely. It may not match exactly because of rounding and because isotopes with very low abundance are sometimes omitted from chemistry questions.

Figure 6 shows that our calculated RAM for iron of 55.85 matches very closely with the value in the periodic table of 55.8. If our calculated RAM was not close to 55.8, we should go back and check the calculation for mistakes.

Another way to check your answer is by confirming that the RAM falls within the range of all the RIM values. For example, if we have isotopes of ^{54}Fe, ^{56}Fe, ^{57}Fe and ^{58}Fe, then the RAM must always be somewhere between 54 and 58. If our calculated RAM was not between 54 and 58, then we should go back and check the calculation for mistakes.

26
Fe
55.8
Iron

FIGURE 6 Iron has a RAM of 55.8.

FIGURE 5 Iron exists in four stable isotopes

7.2B CHALLENGE

Deducing percentage composition from RIM and RAM data

If an element consists of only two isotopes, it is possible to work backwards and calculate the percentage composition of each isotope from RIM and RAM data provided.

We can use the following formula to calculate the percentage abundance of the light isotope.

Let x = the abundance of the lighter isotope

Therefore $(1 - x)$ = the abundance of the heavier isotope

Chlorine has two isotopes: ^{35}Cl (light) and ^{37}Cl (heavy). You may use the mass numbers 35 and 37 as approximations of RIM for these two isotopes.

1 Find the relative atomic mass of chlorine in your VCE Chemistry data book.
2 Calculate the percentage abundance of ^{35}Cl in chlorine.
3 Use your answer to Question 2 to calculate the percentage abundance of ^{37}Cl.

7.2 CHECK YOUR LEARNING

Describe and explain

1 Describe what a mass spectrometer does.
2 Outline what information can be obtained from a mass spectrum.
3 A fictitious element has three isotopes with relative abundances of 42.1%, 100.0% and 65.8%. Use this data to calculate the percentage abundance of each isotope.

Apply, analyse and compare

4 State the difference between relative abundance and percentage abundance.
5 Calculate the relative atomic mass of titanium from the mass spectrum in Figure 7.
6 Calculate the relative atomic mass of tungsten using the mass spectrum data provided in the table.

Isotope	RIM	% abundance
^{182}W	181.95	26.53
^{183}W	182.95	14.33
^{184}W	183.95	30.68
^{186}W	185.95	28.46

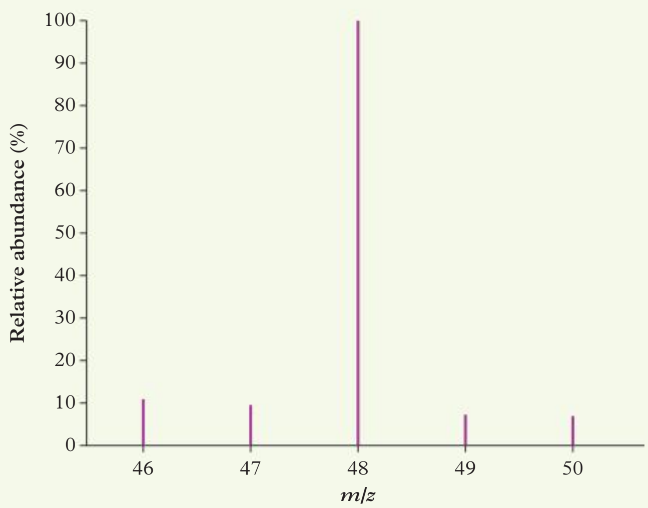

FIGURE 7 The mass spectrum of titanium

7.3 Avogadro's constant and molar mass

KEY IDEAS

In this topic, you will learn that:

- Avogadro's constant has the symbol N_A and a value of 6.02×10^{23} and is a convenient way to quantify the number of atoms or molecules in a pure sample of known mass
- particles (such as atoms, molecules, ions or electrons) can be quantified in moles, where one mole of substance contains 6.02×10^{23} particles
- the formula $n = \frac{m}{M}$ can calculate the amount of substance in moles of a pure substance, where m is mass in grams, and M is the molar mass
- the formula $N = n \times N_A$ can calculate the number of individual particles when the amount of substance in moles is known.

particles
atoms, molecules, ions or electrons

Chemistry involves enormous numbers of tiny **particles**. A paperclip weighing just one gram contains around 10^{22} (10 000 000 000 000 000 000 000) atoms. A grain of rice is much smaller but contains around 10^{20} (100 000 000 000 000 000 000) atoms. Even a tiny cell nucleus, visible only under a microscope, contains approximately 10^{13} atoms (10 000 000 000 000). Chemical samples are usually larger than all these quantities, so we need a very large unit of measurement to help count them.

Chemists count substances in moles

amount
the quantity of a substance measured in moles

mole
an amount of substance that contains 6.02×10^{23} particles

Avogadro's constant (N_A)
6.02×10^{23}

FIGURE 1 Avogadro's constant was named after Italian scientist Amedeo Avogadro.

Number words, such as single, trio and dozen, represent numerical quantities. Many of these number words tend to be used in specific contexts. For example, a ream is 500 sheets of paper; a century is 100 years, and a trio is three people (usually three musicians). A mole is a quantity that describes an **amount** of substance.

The **mole** is a unit that refers to the amount of substance that contains exactly 6.02×10^{23} particles (such as atoms, molecules, ions or electrons). This number is also known as **Avogadro's constant** and is given the symbol N_A.

One mole of a substance is generally a manageable quantity. For example, one mole of water is a large spoonful, one mole of iron is 56 paperclips, and one mole of air can fill two-and-a-half basketballs. Each of these one-mole samples contains 6.02×10^{23} particles. More examples are shown in Table 2.

TABLE 1 Some number words, including mole (which equals 6.02×10^{23} particles)

Number word	Value (and examples)
Single	1
Pair	2 (socks)
Trio	3 (people)
Dozen	12 (eggs)
Century	100 (years)
Ream	500 (sheets of paper)
Millennium	1000 (years)
Million	10^6
Billion	10^9
Trillion	10^{12}
Mole	6.02×10^{23} (particles such as an atom, molecule, ion or electron)
Googol	10^{100} (is more than the number of atoms in the Universe)
Googolplex	10^{googol}

TABLE 2 Examples of samples, their amount in moles (n) and the number of particles they contain

Number of moles (n)	Number of particles (N)	In everyday life
1.0 mole of helium	6.02×10^{23}	About 4 balloons
2.0 moles of sodium chloride	1.20×10^{24}	About 1 cup
10.0 moles of water	6.02×10^{24}	About 1 small cup

It is apparent from Table 2 that the number of particles in a sample is equal to the amount of substance in moles multiplied by 6.02×10^{23}.

Avogadro's constant is defined by carbon-12. One mole of substance contains 6.02×10^{23} particles, which is the same number of particles as there are atoms in exactly 12 grams of carbon-12.

Converting number of particles into moles

The relationship between number of particles (N) and amount of substance in moles (n) of a substance is written as:

$$n = \frac{N}{N_A}$$

where n is amount of substance in moles, N is number of particles and N_A is 6.02×10^{23}.

For example, 4.55×10^{24} water molecules can be measured in moles:

$$\begin{aligned} n(H_2O) &= \frac{N(H_2O)}{N_A} \\ &= \frac{4.55 \times 10^{24}}{6.02 \times 10^{23}} \\ &= 7.56 \text{ mol} \end{aligned}$$

FIGURE 2 Each of these beakers contains 1 mole of a substance. The mass of a mole of a substance can be found by using the molar mass of its elements.

Note that the unit moles is often abbreviated to 'mol'.

We can also rearrange the equation to find the number of particles in a sample if the amount of substance in moles is known.

For example, the number of particles in 0.0525 mol of carbon dioxide molecules is:

$$\begin{aligned} N(CO_2) &= n(CO_2) \times N_A \\ &= 0.0525 \times 6.02 \times 10^{23} \\ &= 3.16 \times 10^{22} \text{ particles} \end{aligned}$$

Practise using experimental data to determine Avogadro's constant in Challenge 7.3.

7.3 Challenge
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Ratios involving moles

One mole of hydrogen molecules (H_2) contains two moles of individual H atoms. The ratio of hydrogen molecules to hydrogen atoms is always 1:2. To find the number of hydrogen atoms in 7 moles of H_2, we need to multiply $7 \times 2 = 14$ moles of hydrogen atoms. (It's similar to how 7 pairs of socks contains 14 socks!)

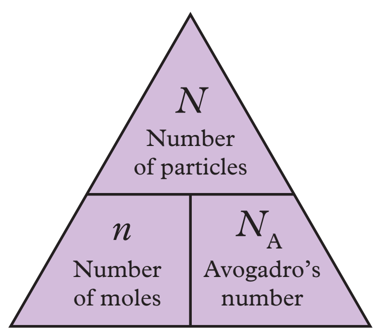


FIGURE 3 A useful way to remember the relationship between N, n and N_A

Study tip

The triangle in Figure 3 can also be useful for finding the right equation. To use it, simply cover up the value that you want to find. For example, if you want to find the number of particles, cover up N to leave $n \times N_A$.

TABLE 3 Finding the amount of each particle in moles within each compound

Sample	Contains the following particles
1.2 moles of H_2O	2.4 moles of H atoms, 1.2 moles of O atoms
4.0 moles of NaCl	4.0 moles of Na^+ ions, 4.0 moles of Cl^- ions
0.13 moles of $AlCl_3$	0.13 moles of Al^{3+} ions, 0.39 moles of Cl^- ions
100.0 moles of $Mg(NO_3)_2$	100.0 moles of Mg^{2+} ions, 200.0 moles of NO_3^- ions, which contain a total of: 200.0 moles of N atoms and 600.0 moles of O atoms

Calculating amounts of substance in moles using mass and molar mass

molar mass
the mass in grams of one mole of substance

The following equation represents the relationship between amount of substance in moles, mass and **molar mass** (M):

$$n = \frac{m}{M}$$

where n is the amount of substance in moles, m is the mass of the sample in grams and M is the molar mass of the substance in grams per mole.

The mass (m) of the sample can be obtained by weighing it on an electronic balance. The molar mass of an element can be obtained from the periodic table.

For example, the amount of substance in moles of 18.5 grams of aluminium can be calculated as follows:

$$\begin{aligned} n(\mathrm{Al}) &= \frac{m(\mathrm{Al})}{M(\mathrm{Al})} \\ &= \frac{18.5}{27.0} \\ &= 0.685 \text{ mol} \end{aligned}$$

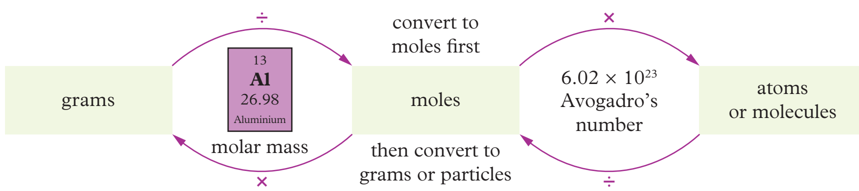


FIGURE 4 The pathway for converting between grams, moles and particles (i.e. atoms or molecules)

Calculating molar mass

The molar masses of elements are written in the periodic table as relative atomic masses (RAM). The molar mass of a compound can be calculated by adding together the RAM of each element in the compound. For example, the molar mass of water can be calculated as follows.

$$\begin{aligned} M(H_2O) &= (2 \times M(\mathrm{H})) + M(\mathrm{O}) \\ &= (2 \times 1.0) + 16.0 \\ &= 18.0 \end{aligned}$$

Study tip

A mole is an amount of substance that contains an Avogadro's constant number of particles. Avogadro's constant was called Avogadro's number until 1971, but both names are still in use today.

More examples of molar masses of compounds are shown in Table 4.

TABLE 4 Molar masses of some compounds

Compound	Molar mass (M, in g mol^{-1})
Ozone (O_3)	3 × 16.0 = 48.0
Carbon dioxide (CO_2)	12.0 + (2 × 16.0) = 44.0
Hydrazine (N_2H_4)	(2 × 14.0) + (4 × 1.0) = 32.0
Glucose ($C_6H_{12}O_6$)	(6 × 12.0) + (12 × 1.0) + (6 × 16.0) = 180.0

7.3 WORKED EXAMPLE

CALCULATING THE AMOUNT OF A SUBSTANCE FROM ITS MASS

Combining $n = \frac{m}{M}$ and $n = \frac{N}{N_A}$ allows you to calculate the number of particles in a sample.

Potassium carbonate is a white powder used in the production of pharmaceuticals and baked goods. A sample of potassium carbonate has mass 50.0 grams.

a Write the chemical formula for potassium carbonate.

b Calculate the molar mass of potassium carbonate.

c Calculate the amount in moles of the potassium carbonate in the sample.

d Calculate the number of potassium ions in the sample.

Solution

Think	Do
Step 1: Write the chemical formula, using the symbols found on the periodic table and correctly assigning ions.	**a** Potassium ions (K^+) and carbonate ions (CO_3^{2-}) combine in a 2:1 ratio to form K_2CO_3.
Step 2: Calculate the molar mass by finding the molar mass of each element and multiplying this by the number of atoms in the compound.	**b** $M(K_2CO_3) = (2 \times 39.1) + 12.0 + (3 \times 16.0)$ $= 138.2 \text{ g mol}^{-1}$
Step 3: Substitute the values into the molar equation. $n = \frac{m}{M}$	**c** $n(K_2CO_3) = \frac{m(K_2CO_3)}{M(K_2CO_3)}$ $= \frac{50.0}{138.2}$ $= 0.362 \text{ mol}$(3 sig fig)
Step 4: Substitute the values into the equation to convert to moles. $n = \frac{N}{N_A}$	**d** $N(K_2CO_3) = n(K_2CO_3) \times N_A$ $= 0.3618 \times 6.02 \times 10^{23}$ $= 2.18 \times 10^{23}$ (3 sig fig) $N(K^+) = 2 \times 2.178 \times 10^{23}$ $= 4.36 \times 10^{23}$ ions (3 sig fig)

7.3 CHECK YOUR LEARNING

Describe and explain

1 Explain what molar mass is.

2 Calculate the molar mass of:
   - **a** nitric oxide (NO)
   - **b** benzene (C_6H_6)
   - **c** phosphorus trichloride (PCl_3).

3 Calculate the amount of substance in moles in:
   - **a** 4.50 g of sodium
   - **b** 204 g of iron
   - **c** 0.105 g of neon.

4 Calculate the mass in grams of:
   - **a** 0.0880 mol of gold
   - **b** 1.55 mol of silver
   - **c** 0.900 mol of platinum.

Apply, analyse and compare

5 Calculate the amount in moles of oxygen atoms in:
   - **a** 1.0 mol ozone (O_3)
   - **b** 2.5 mol ethanoic acid ($C_2H_4O_2$)
   - **c** 12.5 mol sulfur dioxide (SO_2).

6 Calculate the number of oxygen atoms in each sample in Question 5.

7 Which sample contains more atoms: 10.0 g neon or 3.00 g water?

8 Which sample contains more ions: 120.0 g of sodium chloride or 50.0 g of aluminium fluoride?

7.4 Percentage composition and empirical formulas

KEY IDEAS

In this topic, you will learn that:

+ percentage composition of an element in a compound can be calculated by using

$$\% \text{ composition} = \frac{\text{mass due to one element}}{\text{mass due to whole compound}}$$

+ empirical formula is the simplest whole-number ratio of molecular formula.

So far, we've discussed how to quantify the number of moles or mass of a pure substance. But what happens when we want to analyse a compound that contains different elements?

When analysing a substance containing more than one element, we need to work out how much of each element is present. To do this, we can use:

- **percentage composition** – a percentage split that identifies the mass of each element
- empirical formula – a simplified molecular formula, which shows the ratio of atoms contributed by each element in a compound.

percentage composition
percentage by mass of each element in a compound

Percentage composition

Percentage composition tells us how much of the mass of a compound is due to a particular element. For example, in 18.0 grams of water, H_2O, 16.0 grams is due to the mass of oxygen atoms and 2.0 grams is due to the mass of hydrogen atoms. As percentages, we can say that water is 16.0 ÷ 18.0 = 88.9% oxygen and 2.0 ÷ 18.0 = 11.1% hydrogen.

Study tip

Percentage composition can be calculated by using experimentally determined masses or simply by using the chemical formula and the molar masses from the periodic table:
% composition = mass due to one element ÷ mass due to whole compound
= M(one element) ÷ M(whole compound)

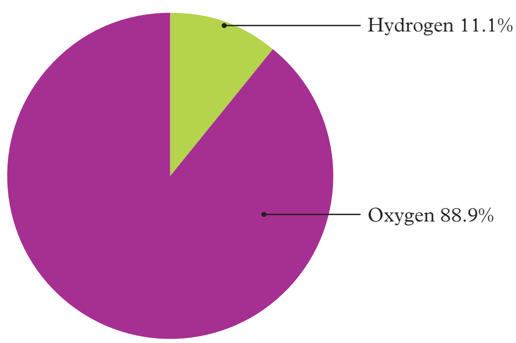


FIGURE 1 The percentage composition of water represented as a pie chart

Calculating percentage composition

To calculate percentage composition, use the following formula:

$$\% \text{ composition} = \frac{\text{mass due to one element}}{\text{mass due to whole compound}}$$

If you do not know the mass of the sample, assume there is one mole of the compound and use molar mass instead of mass to obtain the same result.

For example, the percentage composition of copper in copper sulfate can be calculated:

$$\%\left(\text{Cu in CuSO}_4\right) = \frac{\text{mass due to one element}}{\text{mass due to whole compound}}$$

7.4
PRACTICAL: CONTROLLED EXPERIMENT

How can we experimentally determine the empirical formula of magnesium oxide?
Go to page **506**.

We will assume there is one mole of substance:

$$\begin{aligned}\%\left(\text{Cu in CuSO}_4\right) &= \frac{M(\text{Cu})}{M\left(\text{CuSO}_4\right)}\\ &= \frac{63.5}{63.5 + 32.1 + (4 \times 16.0)}\\ &= \frac{63.5}{159.6}\\ &= 39.8\%\ \text{Cu}\end{aligned}$$

Empirical formula and molecular formula

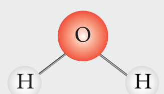
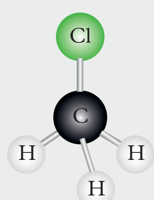
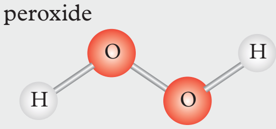
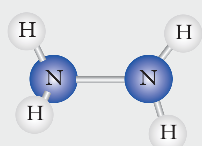
Molecular formulas represent the real number of atoms of each element in the compound. For example, lactic acid has the molecular formula $C_3H_6O_3$ because each molecule contains 3 carbon atoms, 6 hydrogen atoms and 3 oxygen atoms.

Empirical formula

Empirical formulas are the simplest whole-number ratios of molecular formulas. For example, the lactic acid molecular formula $C_3H_6O_3$ can be divided by three to give an empirical formula CH_2O. We divide by 3 because 3 is the **greatest common divisor (GCD)** of 3, 6 and 3. Some examples of molecular formulas and their empirical formulas are shown in Table 1.

greatest common divisor (GCD) the largest positive integer that divides into all the numbers in the set (e.g. 6 is the GCD of 12 and 18)

TABLE 1 Molecular formulas and their empirical formulas (simplest whole-number ratios of the molecular formulas)

Molecule	Molecular formula	Empirical formula
Water 	H_2O	H_2O
Chloromethane 	CH_3Cl	CH_3Cl
Hydrogen peroxide 	H_2O_2	HO
Hydrazine 	N_2H_4	NH_2
Glucose 	$C_6H_{12}O_6$	CH_2O

The empirical formula of a compound can be determined experimentally or 'empirically' (hence the name 'empirical formula'). The empirical formula does not necessarily represent the actual number of atoms in the molecule; it only represents the ratio of atoms of each element.

Determining empirical formula from experimental data

Empirical formulas can be determined from mass data using an $m|n|$ratio table. This converts the mass ratio into a mole ratio. See Worked example 7.4A.

7.4A WORKED EXAMPLE

DETERMINING EMPIRICAL FORMULA USING EXPERIMENTAL DATA

A compound contains 1.84 g sodium, 2.57 g sulfur and 1.92 g oxygen. Determine the empirical formula of the compound.

Solution

Think	Do
Step 1: Construct an $m\|n\|$ratio table and insert the values provided.	<table><tr><td></td><td>Na</td><td>S</td><td>O</td></tr><tr><td>m</td><td>1.84 g</td><td>2.57 g</td><td>1.92 g</td></tr><tr><td>n</td><td></td><td></td><td></td></tr><tr><td>Ratio</td><td></td><td></td><td></td></tr></table>
Step 2: Calculate the amount in moles of each element using $n = \frac{m}{M}$ and insert these values into the table.	<table><tr><td></td><td>Na</td><td>S</td><td>O</td></tr><tr><td>m</td><td>1.84 g</td><td>2.57 g</td><td>1.92 g</td></tr><tr><td>n</td><td>$= \frac{1.84}{23.0} = 0.0800$</td><td>$= \frac{2.57}{32.1} = 0.0800$</td><td>$= \frac{1.92}{16.0} = 0.120$</td></tr><tr><td>Ratio</td><td></td><td></td><td></td></tr></table>
Step 3: Calculate the molar ratio of each element using ratio $= \frac{n}{\text{smallest } n}$. In this, the smallest n is 0.0800 (the lowest value of n in the whole table). Round the ratio values to the nearest integer, half, third or quarter, i.e. to the nearest $\frac{1}{2}, \frac{1}{3}, \frac{2}{3}, \frac{1}{4}, \frac{3}{4}$ or an integer. Never round by more than 0.1 of the actual ratio value.	<table><tr><td></td><td>Na</td><td>S</td><td>O</td></tr><tr><td>m</td><td>1.84 g</td><td>2.57 g</td><td>1.92 g</td></tr><tr><td>n</td><td>$= \frac{1.84}{23.0} = 0.0800$</td><td>$= \frac{2.57}{32.1} = 0.0800$</td><td>$= \frac{1.92}{16.0} = 0.120$</td></tr><tr><td>Ratio</td><td>$= \frac{0.0800}{0.0800} = 1$</td><td>$= \frac{0.0800}{0.0800} = 1$</td><td>$= \frac{0.120}{0.0800} = 1\frac{1}{2}$</td></tr></table>
Step 4: If you have a fraction in one of the final ratios, multiply all ratios by the denominator.	Multiply the ratios by two: $1:1:1\frac{1}{2}$ becomes 2:2:3 The empirical formula of the compound is $Na_2S_2O_3$.

7.4B Worked example
Find me in your obook pro

7.4B Worked example
Video demonstration

Determining molecular formula from experimental data

If the molar mass (M) of the compound is provided, then we can calculate the molecular formula of the compound using the following steps:

1 Calculate the empirical formula.

2 Find the molar mass of the empirical formula.

3 Calculate molecular formula by multiplying the empirical formula by $\frac{M(\text{molecular formula})}{M(\text{empirical formula})}$.

Worked example 7.4B walks you through this calculation. If you're looking to try a more complex calculation, check out Challenge 7.4.

7.4 Challenge
Find me in your obook pro

7.4 CHECK YOUR LEARNING

Describe and explain

1 Calculate the percentage composition of each element in hydrogen chloride (HCl).

2 Calculate the percentage composition of each element in citric acid ($C_6H_8O_7$).

3 State the empirical formulas of:

a ethyne (C_2H_2)

b decane ($C_{10}H_{22}$)

c phosphorus pentoxide (P_4H_{10}).

Apply, analyse and compare

4 A compound consisting of carbon, hydrogen and oxygen contains 11.52 g carbon, 1.92 g hydrogen and 7.68 g oxygen.

a Calculate the empirical formula of the compound.

b If the compound has a molar mass of 98.0 g mol^{-1}, deduce the molecular formula of the compound.

5 A compound consisting of carbon, hydrogen and fluorine contains 36.47% carbon, 5.1% hydrogen and 58.2% fluorine.

a Calculate the empirical formula of the compound.

b If the compound has a molar mass of 66 g mol^{-1}, deduce the molecular formula of the compound.

6 33.83 g of a compound consists of 9.1 g of carbon and 0.752 g of hydrogen. The remaining mass is oxygen.

a Calculate the empirical formula of the compound.

b If the compound has a molar mass of 90.0 g mol^{-1}, deduce the molecular formula of the compound.

7 A 302.97 g sample of a compound consists of 24.97% carbon, 50.03% hydrogen and 25.00% oxygen.

a Calculate the empirical formula of the compound.

b If the compound has a molar mass of 180.0 g mol^{-1}, deduce the molecular formula of the compound.

c A student claims that the unknown compound is fructose. Evaluate their claim.

CHAPTER 7 Review

Chapter summary

7.1

- Relative atomic mass (RIM) is a more accurate way to measure an atom's mass.
- Relative isotopic mass (RIM) is the mass of an atom compared with the mass of a carbon-12 atom, which has a RIM of exactly 12.

7.2

- Mass spectrometers can measure relative isotopic masses of each isotope in an element sample and the relative abundance of each isotope.
- Relative atomic mass (RAM) can be calculated by using data from a mass spectrometer presented in graph or table form.

7.3

- Avogadro's constant has the symbol N_A and a value of 6.02×10^{23} and can be used to quantify the number of atoms or molecules in a pure sample of known mass.
- Particles (such as atoms, molecules, ions or electrons) can be quantified in moles, where one mole of substance contains 6.02×10^{23} particles.
- Molar mass is the mass (g) of one mole of a substance and can be calculated by rearranging the equation for amount in moles.

7.4

- Percentage composition refers to the ratio of elements present in a compound.
- Empirical formula is the simplest whole-number ratio of the molecular formula.

Key formulas

Percentage abundance	$\%\text{ abundance} = \dfrac{\text{relative abundance}}{\text{sum of all the relative abundances}}$
Relative atomic mass (RAM)	$\text{RAM} = \Sigma(\text{RIM} \times \text{percentage abundance})$
Amount in moles	$n = \dfrac{m}{M}$
Number of particles	$N = n \times N_A$
Percentage composition	$\%\text{ composition} = \dfrac{\text{mass due to one element}}{\text{mass due to whole compound}}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• identify the relative isotopic masses of isotopes of elements and their values on the scale in which the relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly	☐	☐	☐	Go back to Topic 7.1
• determine the relative atomic mass of an element using mass spectrometry data	☐	☐	☐	Go back to Topic 7.2
• use Avogadro's constant (6.02×10^{23}) to determine the amount, in moles, or atoms (or molecules) in a pure sample of known mass	☐	☐	☐	Go back to Topic 7.3
• determine the molar mass of compounds, the percentage composition by mass of covalent compounds, and the empirical and molecular formula of a compound from its percentage composition by mass	☐	☐	☐	Go back to Topic 7.4

Revision questions

Multiple choice

1 What is the relative atomic mass of chlorine?

A 17
B 35
C 35.5
D –1

2 How many protons, neutrons and electrons are in $^{56}Fe^{3+}$?

A 26 protons, 30 neutrons and 26 electrons
B 26 protons, 30 neutrons and 23 electrons
C 26 protons, 56 neutrons and 26 electrons
D 26 protons, 56 neutrons and 23 electrons

3 A fictitious element called victorium (symbol Vi) has two isotopes: ^{97}Vi and ^{99}Vi. A sample of victorium contains 69.0% ^{97}Vi (RIM = 96.97) and 31.0% ^{99}Vi (RIM = 98.96). Calculate the relative atomic mass of victorium.

A 97.6
B 97.59
C 98.3
D 98.34

4 An element has four isotopes with relative abundances of 30.2%, 67.8%, 100.0% and 12.4%. What is the percentage abundance of the most abundant isotope?

A 47.5%
B 67.8%
C 100.0%
D 210.4%

5 Rubidium has two stable isotopes: ^{85}Rb (RIM = 84.91) and ^{87}Rb (RIM = 86.91). The RAM of rubidium is 85.47. The percentage abundances of the two isotopes of rubidium are:

A 66% ^{85}Rb and 34% ^{87}Rb.
B 68% ^{85}Rb and 32% ^{87}Rb.
C 70% ^{85}Rb and 30% ^{87}Rb.
D 72% ^{85}Rb and 28% ^{87}Rb.

6 The amount of sodium ions in 3.0 mol of sodium phosphate (Na_3PO_4) is:

A 1.0 mol.
B 3.0 mol.
C 4.5 mol.
D 9.0 mol.

7 The number of particles in 5.8 mol of any substance is:

A 5.8

B 6.02×10^{23}

C 3.5×10^{24}

D 5.8×10^{25}

8 The mass of 0.00515 mol of $Ca(OH)_2$ is:

A 0.0294 g

B 196 mg

C 382 mg

D 7.41×10^{-6} kg

9 The total number of ions in 0.0500 mol copper(II) chloride is

A 1.50×10^{22}

B 3.01×10^{22}

C 6.02×10^{23}

D 9.03×10^{22}

FIGURE 1 Copper(II) chloride is a bright blue solid.

10 The amount of substance (in mol) in 65.0 g neon gas is:

A 3.11 mol

B 3.22 mol

C 3.33 mol

D 3.44 mol

Short answer

Describe and explain

11 Explain the difference between mass number and relative atomic mass.

12 Explain why the relative isotopic mass of magnesium-24 is *not* exactly double the relative isotopic mass of carbon-12.

13 Outline the three pieces of information that can be obtained from a mass spectrum.

14 Calculate the amount (in mol) of oxygen atoms in each of the following samples.

a 5.0 mol H_2O_2

b 12.0 mol N_2O_4

c 6.70×10^{-2} mol MgO

15 Calculate the molar mass (M) of each of the following substances.

a Glucose ($C_6H_{12}O_6$)

b Silver nitrate ($AgNO_3$)

c Potassium permanganate ($KMnO_4$)

16 Calculate the amount (in mol) of substance in each of the following samples.

a 20.0 g octane (C_8H_{18})

b 1.95 kg sodium carbonate (Na_2CO_3)

c 6.48 mg magnesium chloride ($MgCl_2$)

Apply, analyse and compare

17 Calculate the relative atomic mass of a fictitious element called chemistrium (symbol Ch), given that it consists of three isotopes: 4.2% ^{44}Ch, 85.8% ^{46}Ch and 10.0% ^{48}Ch. You can use the mass numbers as approximations of RIM.

18 Calculate the percentage abundance of each isotope of antimony, given that it consists of two isotopes: ^{121}Sb (RIM = 120.9) and ^{123}Sb (RIM = 122.9). The relative atomic mass of antimony is 121.8.

FIGURE 2 Antimony

19 A compound used in dietary supplements contains 20.5 g iron, 13.2 g carbon and 35.3 g oxygen.

a Calculate the empirical formula of the compound.

b Given that the molar mass of the compound is 375.6 g mol^{-1}, calculate the empirical formula of the compound.

20 A compound used in rechargeable batteries contains 7.3% lithium, 32.7% phosphorus and 60.1% fluorine by mass.

a Calculate the empirical formula of the compound.

b Given that the molar mass of the compound is 94.9 g mol^{-1}, calculate the empirical formula of the compound.

21 An 18.35 g sample of hydrated calcium sulfate is heated to constant mass. The mass of the anhydrous calcium sulfate is 14.51 g.

a Calculate the mass of the water removed during heating.

b Deduce the chemical formula of the hydrated calcium sulfate and write it as $CaSO_4.xH_2O$, where x is the degree of hydration.

Design and discuss

22 Rocky planets with a different surface gravity from Earth have different relative abundances of elements in their crust. Smaller planets with weaker surface gravity have higher concentrations of heavier isotopes of elements at the surface.

Imagine that in 2050, a mission to Mars has collected multiple samples of Martian soil and returned them safely to Earth. Design a method that describes how the relative atomic masses of elements could be calculated using these Martian soil samples.

FIGURE 3 An artist's impression of astronauts collecting soil samples from Mars in 2050

23 Is the number of neutrons in an atom always the same as the number of protons? Discuss.

24 Imagine you have purchased a density cube made from pure zinc, which measures exactly 2 cm on each side. How could you use this density cube to calculate an approximate value for the volume of a zinc atom?

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter

CHAPTER 8

Families of organic compounds

KEY KNOWLEDGE

→ the grouping of hydrocarbon compounds into families (alkanes, haloalkanes, alkenes, alcohols, carboxylic acids) based upon similarities in their physical and chemical properties, including general formulas and general uses based on their properties

→ representations of organic compounds (structural formulas, semi-structural formulas) and naming according to the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature (limited to non-cyclic compounds up to C8, and structural isomers up to C5)

→ plant-based biomass as an alternative renewable source of organic chemicals (for example, solvents, pharmaceuticals, adhesives, dyes and paints) traditionally derived from fossil fuels

→ materials and products used in everyday life that are made from organic compounds (for example, synthetic fabrics, foods, natural medicines, pesticides, cosmetics, organic solvents, car parts, artificial hearts), the benefits of those products for society, and the health and/or environmental hazards they pose.

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

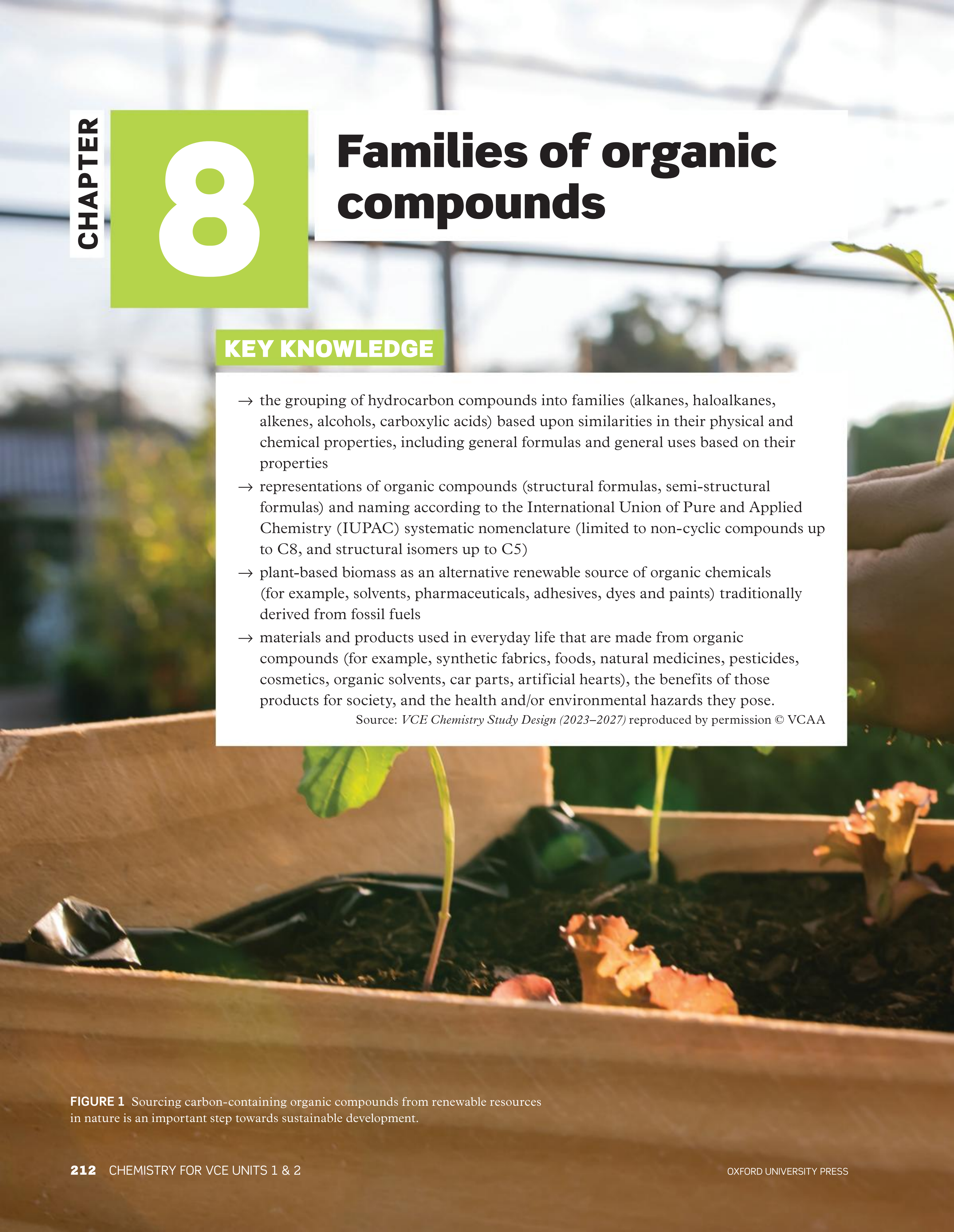


FIGURE 1 Sourcing carbon-containing organic compounds from renewable resources in nature is an important step towards sustainable development.

GROUNDWORK

In Chapter 8, you will learn about organic compounds, including what they are, how we name and represent them, where they can be sourced from, and the benefits and hazards of organic compounds in the products that we use every day.

This chapter will build on concepts you have already learnt in Chapters 1 and 3. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

8A How do the United Nations Sustainable Development Goals, the green chemistry principles, and the concept of a circular economy relate to each other?

8A Groundwork resource
Sustainability

8B What are the types of bonds that can form between non-metals?

8B Groundwork resource
Covalent bonding

8C How does a molecule's shape affect its physical properties?

8C Groundwork resource
Molecular shape and physical properties

PRACTICALS

8.1 PRACTICALS: CLASSIFICATION & IDENTIFICATION	Can you identify organic compounds from their physical properties?	Page 508
8.4 PRACTICALS: CONTROLLED EXPERIMENT	Which milk makes the strongest glue?	pro

8.1 Grouping hydrocarbon compounds

KEY IDEAS

In this topic, you will learn that:

+ hydrocarbons are organic compounds that contain carbon and hydrogen
+ hydrocarbons can have non-metal functional groups that affect their chemical and physical properties
+ hydrocarbons can be grouped into alkanes, haloalkanes, alkenes, alcohols and carboxylic acids depending on the type of functional group they contain.

organic compound
any compound that contains carbon

As you learnt in Chapter 3, carbon can share its valence electrons to form four covalent bonds with other atoms, including hydrogen, oxygen, other carbon atoms, and a range of other non-metals. Any compound containing carbon is called an **organic compound**. In fact, carbon is considered the 'backbone' of organic compounds because it can:

- form single, double or triple bonds
- be connected in chains of different length
- form compounds of different shapes.

Therefore, carbon is a unique and very versatile atom. Organic compounds are all around us: in the things we consume (e.g. sugar, caffeine and artificial flavours) and in the things that we use in our everyday lives (e.g. fuels, plastics and clothing). You will learn more about everyday applications of organic compounds in Topic 8.4.

Hydrocarbon families

hydrocarbon
an organic compound that contains one or more carbon atoms bonded to hydrogen atoms

homologous series
a family of compounds that contain the same functional groups and have similar properties; they differ by a $-CH_2$

When a compound contains one or more carbon atoms bonded to hydrogen atoms, it is called a **hydrocarbon**. Hydrocarbons can be grouped into a family (or a **homologous series**) if they have:

- the same general formula
- a similar structure
- a trend in physical properties
- the same chemical properties.

FIGURE 1 Carbon is found in many everyday objects, including the sugar in donuts, the fuel that powers cars, the plastic in bottle caps, and clothes.

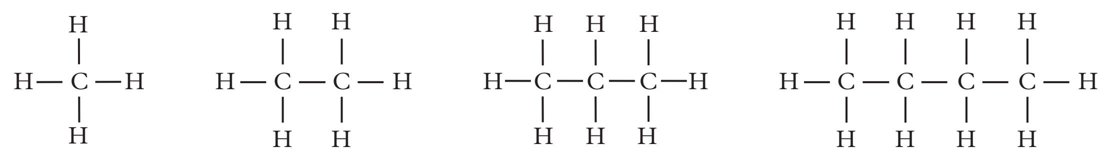
8.1
PRACTICAL: CLASSIFICATION & IDENTIFICATION
Can you identify organic compounds from their physical properties? Go to page **508**.

Alkanes

Alkanes are **saturated** hydrocarbons. This means that they only contain single carbon–carbon bonds, with the maximum possible number of hydrogen atoms attached to all carbons. The simplest of the alkanes is methane (CH_4). Figure 2 shows the first four compounds in the alkane family. As you go across the figure, each compound gains a carbon and two hydrogen atoms ($–CH_2–$). This means that they are considered a homologous series of alkanes. The names of all the alkanes end in the suffix -ane. In Topic 8.2, you will learn more about the rules for naming organic compounds.

alkane
a saturated hydrocarbon; contains only carbon–carbon single bonds

saturated
containing only carbon–carbon single bonds



Methane (CH_4) **Ethane (C_2H_6)** **Propane (C_3H_8)** **Butane (C_4H_{10})**

FIGURE 2 The first four alkanes

General formula of alkanes

A **general formula** is a simple rule that allows you to determine the molecular formula without needing to count all of the hydrogen atoms. For alkanes, the general formula is C_nH_{2n+2}, where n is the number of carbon atoms in the molecule. For example, butane contains four carbon atoms. If $n = 4$, then $2n + 2 = 8 + 2 = 10$. Therefore, the formula for butane is C_4H_{10}. For any homologous series, you can use a general formula to find the molecular formula as long as you know the number of carbon atoms.

general formula
a rule allowing you to determine the molecular formula of an organic compound; each hydrocarbon family has a different general formula

Structural isomers of alkanes

Structural isomers are compounds with the same molecular formula, number and type of atoms, but arranged in a different way. This means that their properties are also different. For example, butane (C_4H_{10}) can be arranged in two ways (Figure 3). One has four carbon atoms arranged in a single chain, and the other consists of three carbon atoms and one carbon branch (a **methyl group**). Branched chains can consist of more than one carbon. A hydrocarbon branch is generally called an **alkyl group**.

structural isomers
compounds with the same molecular formula, but a different structure

methyl group
a hydrocarbon branch consisting of one carbon and three hydrogen atoms

alkyl group
a group where a hydrogen atom has been removed from an alkane

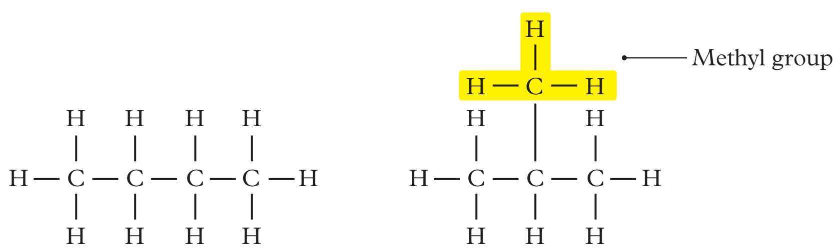


FIGURE 3 C_4H_{10} has two structural isomers.

All alkanes containing four or more carbon atoms (butane onwards) have structural isomers. The more carbon atoms in the molecule, the more isomers there will be. For example, C_7H_{16} has 10 structural isomers and C_9H_{20} has 35. Each of these are named differently from the original alkane. You will learn more about this in Topic 8.2.

Properties of alkanes

The structure of molecules is what gives them their properties. Table 1 summarises the properties of alkanes, including water solubility, melting and boiling points, whether they are **flammable**, and their reactivity.

flammable
can be burned in a combustion reaction

TABLE 1 The properties of alkanes

Property	Explanation
Insoluble in water	Alkanes are made up of only carbon and hydrogen. Because the temporary dipoles are so small, alkanes are non-polar. Non-polar compounds are not soluble in polar water.
Low boiling and melting point	Alkanes only have dispersion forces between molecules. Smaller alkanes are gases at room temperature. As chain length increases, dispersion forces increase. Therefore, the boiling and melting points increase (Figure 4).
Flammable	Although alkanes are not very reactive with other molecules, they can become explosive when heated in the presence of oxygen.
Low reactivity	All carbon atoms have four stable covalent bonds, making them unlikely to react with other molecules.

Study tip

It is very important to understand how the structure of a molecule affects its properties, including solubility and melting/boiling points. Go back to Chapters 3 and 6 to review this.

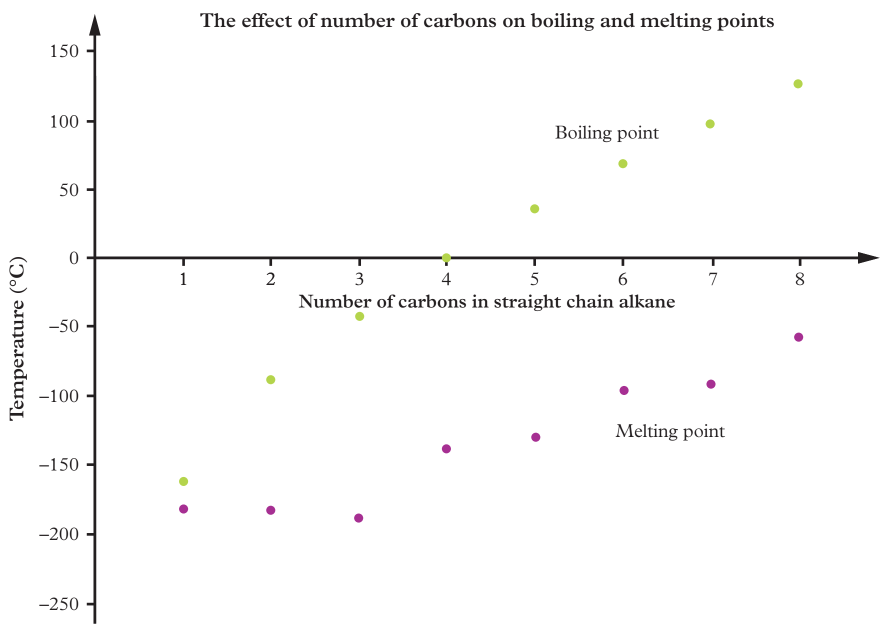
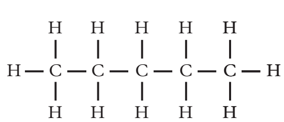
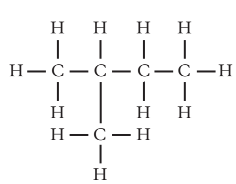


FIGURE 4 The boiling and melting points of alkanes increase as the chain length increases.

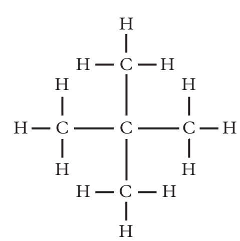
Generally, the branching of carbon chains means that the dispersion forces between the molecules are reduced. The branches prevent the molecules from packing together as tightly. This reduces their boiling points (Figure 5).



Pentane
Boiling point 36°C



2-Methylbutane
Boiling point 30°C



2,2-Dimethylpropane
Boiling point 10°C

FIGURE 5 The boiling points of isomers of pentane

Applications of alkanes

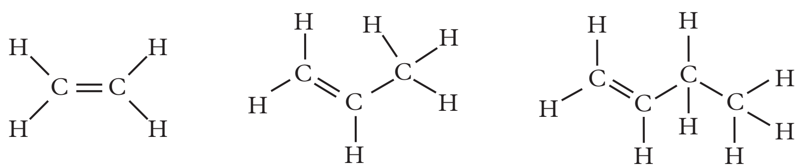
The properties of alkanes make them very useful. They are used as **raw materials** in the chemical industry. This means that they are a good starting point for synthesising many larger, more complex molecules.

Alkanes also have many other uses; for example, as non-polar solvents to dissolve non-polar solutes; fuels (e.g. gas and petrol); lubricants (long-chain alkanes (C17–C35) are very viscous); and bitumen (e.g. road surfaces (>C35)).

FIGURE 6 Alkanes have many uses, including as fuels and anti-corrosion coatings.

Alkenes

Alkenes are a family of hydrocarbons that have at least one carbon–carbon double bond. Molecules that contain at least one double (or triple) carbon–carbon bond are **unsaturated**. They do not have the maximum possible number of hydrogen atoms attached to all carbon atoms. The double bond is considered a **functional group** because it adds unique properties to the hydrocarbon. Like the alkanes, they are made up of only carbon and hydrogen. The first three alkenes are shown in Figure 7. All the alkenes end in the suffix -ene.



Ethene (C_2H_4) **Propene (C_3H_6)** **But-1-ene (C_4H_8)**

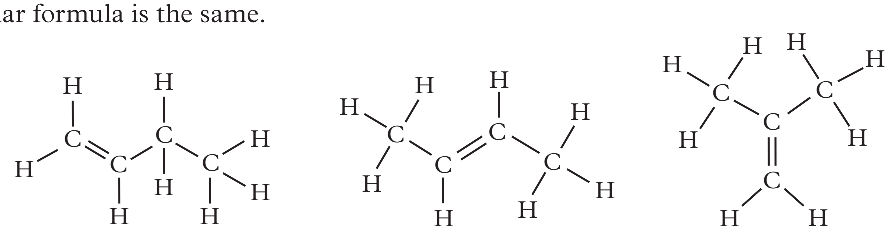
FIGURE 7 The first three alkenes

General formula of alkenes

The general formula for alkenes is C_nH_{2n}, where n is the number of carbon atoms in the molecule. For example, an alkene with six carbon atoms ($n = 6$) will have $2n$ hydrogen atoms, so $2n = 2 \times 6 = 12$. Therefore, the molecular formula for this alkene will be C_6H_{12}. Because of the carbon–carbon double bond, an alkene will have two fewer hydrogen atoms than the alkane with the same number of carbon atoms.

Structural isomers of alkenes

Structural isomers exist for alkenes with four or more carbon atoms. Isomers can result from branching or a change in the position of the carbon–carbon double bond (Figure 8). The position of the double bond has changed between but-1-ene and but-2-ene. In methylpropene, there is a methyl group branch. Structural isomers can look quite different, but their molecular formula is the same.



But-1-ene **But-2-ene** **Methylpropene**

FIGURE 8 But-1-ene has two additional structural isomers: but-2-ene and methylpropene.

raw materials
compounds used as starting materials for the synthesis of larger and more complex compounds

alkene
an unsaturated hydrocarbon; contains at least one carbon–carbon double bond

unsaturated
containing at least one carbon–carbon double bond

functional group
a group of atoms that is responsible for the characteristic reactions of a compound

Study tip

Because alkenes must contain a carbon–carbon double bond, methene containing one carbon atom, cannot exist.

Properties of alkenes

The properties of alkenes are summarised in Table 2. Table 3 compares the boiling points of the four alkanes and alkenes.

TABLE 2 The properties of alkenes

Property	Explanation
Insoluble in water	Like alkanes, alkenes are insoluble in water because they are non-polar.
Low boiling and melting point	The alkenes have lower boiling points than alkanes with the same number of carbon atoms. The double bond affects packing of the molecules and the dispersion forces between them. As with alkanes, the boiling and melting points of alkenes increases with carbon chain length (Table 3).
Flammable	Like alkanes, alkenes burn readily in the presence of oxygen.
Stable but more reactive than alkanes	The carbon–carbon double bond is reactive. They can be involved in addition reactions, which you will learn about in Chapter 9.

FIGURE 9 Alkenes can be used to form polystyrene foam.

TABLE 3 The boiling points of alkenes compared to alkanes of the same chain length

Name	Boiling point (°C)	
	-ane	-ene
Eth- (two carbon atoms)	–88.6	–104.0
Prop- (three carbon atoms)	–42.1	–47.70
But- (four carbon atoms)	–0.5	–6.30
Pent- (five carbon atoms)	36.06	30.00

Uses of alkenes

Alkenes have similar uses to alkanes, including as non-polar solvents and fuels. However, the greater reactivity of alkenes makes them too valuable to use as fuels and so alkanes are generally used for this purpose. Instead, alkenes are often used as reactants in addition reactions. They can be used to make products such as antifreeze, polystyrene, polyethylene, ethanol, acetone, acrylic fibres and plastics. You will learn about the production of polymers by addition reactions in Chapter 9.

Haloalkanes

haloalkane
an alkane that contains a halogen functional group

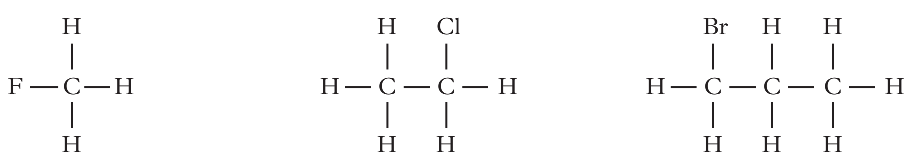
halogen
a group 17 atom, including fluorine, chlorine, bromine and iodine

Haloalkanes are alkanes that contain a **halogen** functional group. In haloalkanes, a group 17 atom (e.g. fluorine, chlorine, bromine, iodine) has replaced a hydrogen in the alkane. This is because halogens have seven valence electrons and can share a pair of electrons with carbon to form a covalent bond.

Haloalkanes are named by adding a prefix to the alkane name. For example, methane with a fluorine replacing one hydrogen is called fluoromethane. In the same way, chlorine will give the prefix chloro-, bromine is bromo- and iodine is iodo-.

Study tip

A prefix is the beginning of a word, such as meth-, eth-, propyl- or chloro-. A suffix is the end of the word, such as -ane, -ene or -ol.



Fluoromethane **Chloroethane** **1-Bromopropane**

FIGURE 10 Three examples of haloalkanes

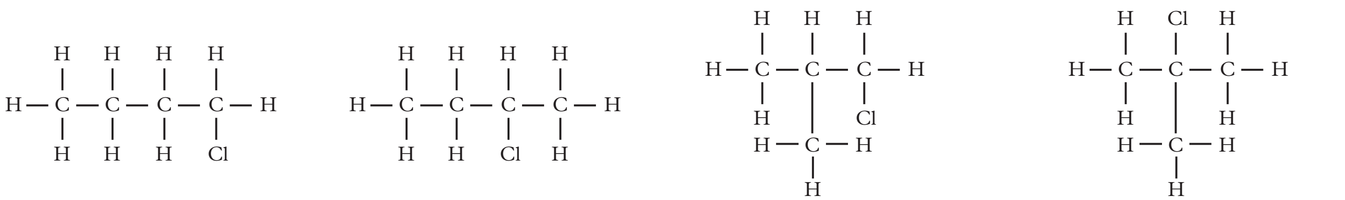
General formula of haloalkanes

The general formula for haloalkanes is very similar to that of the alkanes. The difference is there is one less hydrogen and there is an X group that represents the specific halogen that is attached. This means that the formula is $C_nH_{2n+1}X$, where n is the number of carbon atoms in the molecule. For example, a chloroalkane with three carbon atoms ($n = 3$) will have $2 \times 3 + 1 = 7$ hydrogen atoms and one chlorine. Therefore, the formula is C_3H_7Cl.

Structural isomers of haloalkanes

Different structural isomers exist for haloalkanes because the halogen can replace any hydrogen atom. Figure 11 shows four different isomers of C_4H_9Cl: they differ by whether they have a linear or a branched chain and in the position of the chlorine atom, which can be on a **terminal carbon** (end carbon) or central carbon.

terminal carbon
the end carbon in a hydrocarbon chain



1-Chlorobutane **2-Chlorobutane** **1-Chloro-2-methylpropane** **2-Chloro-2-methylpropane**

FIGURE 11 Four structural isomers of C_4H_9Cl

Properties of haloalkanes

The replacement of hydrogen with a halogen results in a dipole. The difference in electronegativity between the halogen and carbon depends on the specific halogen, but the bond between the halogen and carbon will be polar. You can find the electronegativity values for carbon and the halogens in Table 4.

This polarity in the carbon–halogen bond gives haloalkanes a variety of physical properties that are different from those of alkanes (Table 5). Table 6 on the following page lists the boiling points of some haloalkanes.

TABLE 4 Electronegativities of carbon and the halogens

Atom	Electronegativity
C	2.6
F	4.0
Cl	3.2
Br	3.0
I	2.7

TABLE 5 The properties of haloalkanes

Property	Explanation
Slightly soluble in water	The polar carbon–halogen bond allows haloalkanes to form dipole–dipole attractions with water. The different electronegativities (I < Br < Cl < F) affects the strength of dipole–dipole attractions. More electronegative halogens form stronger intermolecular forces with water molecules and are more soluble.
Low boiling and melting point	Dipole–dipole attractions also form between haloalkane molecules. Together with dispersion forces, this gives haloalkanes higher boiling and melting points. The greater the mass (F < Cl < Br < I), the stronger the dispersion forces and dipole–dipole attractions. So, haloalkanes with heavier/larger halogens have higher boiling and melting points.
Less flammable	Halogens make alkanes less flammable.
Reactive	Haloalkanes are reactive because of the electronegative halogen.

TABLE 6 The boiling points of four haloalkanes compared with the corresponding alkanes

Haloalkane	Boiling points (°C)				
	X = H	X = F	X = Cl	X = Br	X = I
CH_3X	–161.7	–78.4	–24.2	3.6	42.4
CH_3CH_2X	–88.6	–37.7	12.3	38.4	72.3
$CH_3CH_2CH_2X$	–42.1	–2.5	46.6	71.0	102.5
$CH_3(CH_2)_2CH_2X$	–0.5	32.5	78.4	101.6	130.5

The position of the halogen along the hydrocarbon chain also affects the properties of the haloalkanes. For example, the boiling point of 1-chloropropane is 46.6°C, whereas the boiling point of 2-chloropropane is 35.0°C.

Uses of haloalkanes

Haloalkanes are used in a wide variety of commercial products, including fire extinguishers, flame retardants, refrigerants, propellants, pharmaceuticals, solvents and cleaning products.

FIGURE 12 Haloalkanes are used in flame retardants and pharmaceutical products.

Some haloalkanes are considered to be toxic environmental pollutants, including chlorofluorocarbons, which were universally used as refrigerants. However, in the 1980s, their contribution to the ozone layer depletion and the greenhouse effect became known and their uses have become restricted.

Alcohols

hydroxyl
a functional group consisting of an oxygen bonded to a hydrogen (–OH)

Alcohols are hydrocarbons containing a **hydroxyl** (–OH) functional group. As you saw with haloalkanes, the –OH replaces a hydrogen in the hydrocarbon chain. The –OH is polar and allows alcohols to form hydrogen bonds. Here, you will learn about how this functional group gives alcohols very special properties compared with the hydrocarbon families we have explored so far. Alcohols are named with the suffix -ol.

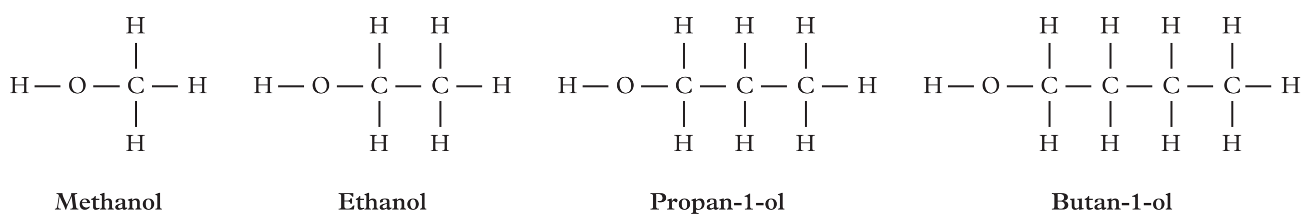


FIGURE 13 Four alcohols: methanol, ethanol, propan-1-ol and butan-1-ol

General formula of alcohols

Alcohols have the general formula $C_nH_{2n+2}O$ or $C_nH_{2n+1}OH$, where n is the number of carbon atoms in the molecule. For example, an alcohol with six carbon atoms ($n = 6$) will have $2n + 2$ carbon atoms in total. $2 \times 6 + 2 = 14$ hydrogen atoms and one oxygen atom gives the formula $C_6H_{14}O$. You can also write this as $C_6H_{13}OH$, which is more useful in VCE Chemistry. There are many functional groups with an oxygen atom, so it is better to have the OH separate so you can easily identify that it is an alcohol.

Structural isomers of alcohols

Alcohols with three or more carbon atoms can have structural isomers. These can differ in the position of the hydroxyl group. As you saw with the other hydrocarbon families, the carbon chain of alcohols can also be branched. The greater the number of carbon atoms, the greater the number of isomers.

Properties of alcohols

As with halogens, the –OH functional group of an alcohol is polar. However, unlike halogens, they can form hydrogen bonds. This has a big impact on the properties of alcohols (Table 7).

TABLE 7 The properties of alcohols

Property	Explanation
Soluble in water	The hydroxyl group can form strong hydrogen bonds with water. Therefore, alcohols are soluble.
Higher boiling and melting point	Hydrogen bonding is stronger than dispersion forces and dipole–dipole attractions. Therefore, alcohols have higher boiling and melting points than alkanes, alkenes and haloalkanes (of the same carbon length), which do not hydrogen bond.
Flammable	Alcohols burn readily in oxygen.
Slightly reactive	Alcohols are less reactive than haloalkanes.

Study tip

Remember, for a molecule to form hydrogen bonds, it has to have FON (fluorine, oxygen or nitrogen) and a hydrogen covalently bonded together.

Uses of alcohols

The most obvious use of alcohols is in alcoholic drinks. Beer, wine and spirits all contain ethanol, which is produced by the fermentation of sugars. Ethanol is also becoming more common in transport fuels. Fermenting biomass (such as corn) produces bioethanol, a renewable fuel.

Alcohols are also commonly used as solvents. Because they can dissolve many organic compounds and are non-toxic, they are found in perfumes, fragrances and cosmetics. They are also common solvents used in research laboratories, including for analytical techniques such as high-performance liquid chromatography, which you might remember from Chapter 6.

Carboxylic acids

Carboxylic acids contain a **carboxyl** (–COOH) group that can only be present on an end (or terminal) carbon. This functional group consists of a carbon with a double bond to one oxygen and a single bond to a hydroxyl group (Figure 14). Carboxylic acids are named with the suffix -oic acid.

carboxyl
a functional group consisting of a carbon double bonded to an oxygen and single bonded to a hydrogen (–COOH)

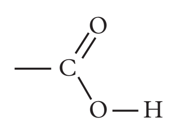
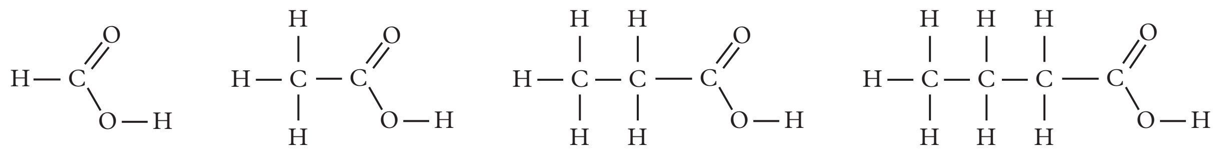


FIGURE 14 A carboxyl functional group



Methanoic acid ($HCOOH$) | **Ethanoic acid (CH_3COOH)** | **Propanoic acid (C_2H_5COOH)** | **Butanoic acid (C_3H_7COOH)**

FIGURE 15 The first four carboxylic acids

General formula of carboxylic acids

Carboxylic acids have the general formula $C_nH_{2n+1}COOH$, where n is the number of carbon atoms, excluding the carbon in the carboxyl group. For example, let's look at a carboxylic acid with five carbon atoms. One carbon will be in the carboxyl group, so $n = 4$. Therefore, excluding the carboxyl group, there will be $2 \times 4 + 1 = 9$ hydrogen atoms in the hydrocarbon chain. The molecular formula is C_4H_9COOH. Combining all the like atoms to get the formula $C_5H_{10}O_2$ is also correct. However, as with alcohols, it's easier to tell that it is a carboxylic acid if the carboxyl group is shown separately.

Structural isomers of carboxylic acids

There are only branched structural isomers of the carboxylic acids. This is because, in order to form one double bond to an oxygen and a single bond to a hydroxyl group, the carbon must be an end carbon. Branched isomers can occur when there are four or more carbon atoms in the carboxylic acid.

Properties of carboxylic acids

Like the alcohols, carboxylic acids can also form hydrogen bonds (Figure 16). This is because the carboxyl functional group is polar.

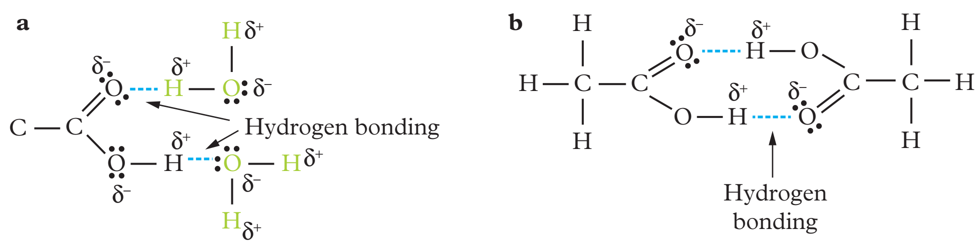


FIGURE 16 Carboxyl functional groups can form hydrogen bonds with **a** water and **b** other carboxyl groups.

The effects of this on the properties of carboxylic acids are shown in Table 8.

TABLE 8 The properties of carboxylic acids

Property	Explanation
Highly soluble in water	The carboxyl group can form strong hydrogen bonds with water. Therefore, carboxylic acids are soluble.
Very high boiling and melting point	Because there is the opportunity to form two hydrogens per molecule, carboxylic acids can form stronger hydrogen bonding than alcohols. This gives them very high boiling and melting points.
Not very flammable	Carboxylic acids are not as flammable as the other hydrocarbon families.
Slightly reactive	Carboxylic acids are less reactive than haloalkanes. They are also weak acids that can donate a proton (H^+). This means they can participate in acid–base reactions. You will learn about this in Chapter 11.

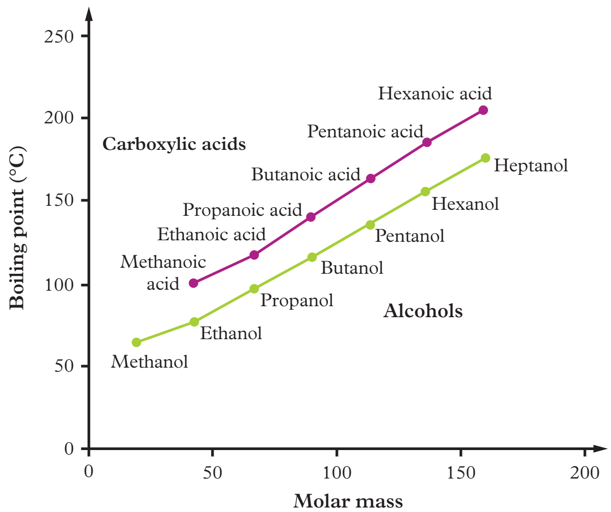


FIGURE 17 The boiling points of carboxylic acids are higher than those of alcohols with the same molar masses.

Uses of carboxylic acids

Carboxylic acids are very common in nature. They have a sour taste. One example is citric acid (Figure 18), which gives citrus fruits their acidic properties. Another example is ethanoic acid (CH_3COOH), which is found in vinegar.

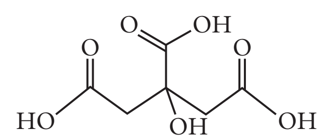


FIGURE 18 Citric acid is found in citrus fruits.

In nature, ants and other insects release corrosive carboxylic acids, such as formic acid (CHOOH), as a defence mechanism.

Because of their reactivity, carboxylic acids can be used in the production of polymers, biopolymers, adhesives, coatings and pharmaceuticals. Reactions with carboxylic acids can also form esters, which are used as artificial scents and flavours in food and cosmetics.

Comparing hydrocarbon families

The key information you learnt in this topic is summarised in Table 9.

TABLE 9 Comparing hydrocarbon families

	Alkanes	Alkenes	Haloalkanes	Alcohols	Carboxylic acid
Functional group	N/A	Carbon–carbon double bond	Halogen (F, Cl, Br, I)	Hydroxyl (–OH)	Carboxyl (–COOH)
General formula	C_nH_{2n+2}	C_nH_{2n}	$C_nH_{2n+1}X$	$C_nH_{2n+1}OH$	$C_nH_{2n+1}COOH$
Solubility in water	Insoluble	Insoluble	Slightly soluble	Soluble	Highly soluble
Melting and boiling point	Low	Lower than alkanes	Higher than alkanes	Higher than haloalkanes	Very high
Flammable	Yes	Yes	Less flammable than alkanes	Yes	Least flammable
Reactivity	Low	Low but more reactive than alkanes	Moderate	Low but more reactive than alkanes	Moderate but less reactive than haloalkanes
Applications	• Solvents • Fuels • Lubricants • Anti-corrosion coatings	• Solvents • Fuels • Raw material for addition reactions	• Fire extinguishers • Refrigerants • Pharmaceuticals • Cleaning products	• Solvents • Beverages • Biofuels • Fragrances	• Food additives • Used to produce polymers, adhesives, pharmaceuticals, etc

8.1 CHECK YOUR LEARNING

Describe and explain

1 Describe what a hydrocarbon is.

2 Using an example, explain how a functional group can affect the properties of a compound.

3 Explain what a structural isomer is.

Apply, analyse and compare

4 Use the general formulas to write a chemical formula for:

- **a** an alkane containing six carbons
- **b** a carboxylic acid containing three carbons
- **c** an alcohol containing eight carbons
- **d** an alkane containing three carbons
- **e** an alkene containing three carbons.

5 Rank the following in order of highest to lowest boiling point and explain your reasoning.

$CH_3CH_2CH_2CH_2CH_3$

$CH_3CH_2CH=CHCH_3$

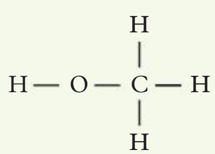
$CH_3CH_2CH_2CH_2CH_2OH$

$CH_3CH_2CH_2CH_2CH_2Cl$

6 Compare the boiling points of carboxylic acids with those of the other hydrocarbon families. Refer to their structure in your response.

7 Identify the homologous series that each organic compound belongs to.

a



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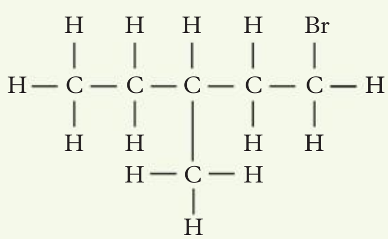

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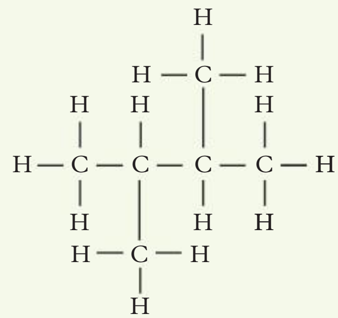

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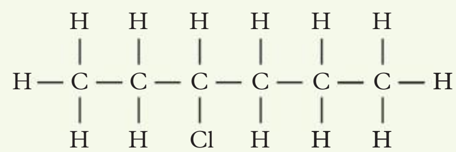

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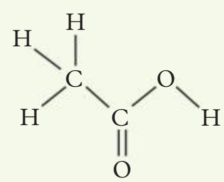

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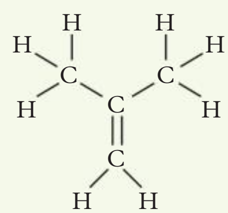


Design and discuss

8 Draw two structural isomers of C_5H_{12} and discuss how their structures affect their boiling points.

9 Evaluate this statement: 'Alcohols are more soluble than alkanes in water'.

8.2 Representing organic compounds

KEY IDEAS

In this topic, you will learn that:

+ organic compounds are named according to a set of rules called the International Union of Pure and Applied Chemistry (IUPAC) systematic nomenclature or naming rules
+ organic molecules can be represented in four ways: molecular formula, structural formula, semi-structural formula and skeletal structure.

IUPAC systematic nomenclature
a set of rules set by the International Union of Pure and Applied Chemistry to consistently name organic compounds

A set of rules for naming organic compounds is outlined by the International Union of Pure and Applied Chemistry (IUPAC). This is known as the **IUPAC systematic nomenclature**. These naming rules were created so that chemists can communicate, name and identify organic compounds in a consistent way.

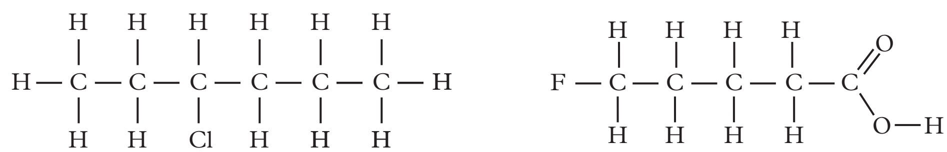


FIGURE 1 The IUPAC names for these molecules are 3-chlorohexane and 5-fluoropentanoic acid. In this topic, you will learn how to apply the IUPAC rules to name a range of organic compounds.

Study tip

Be careful when identifying the longest chain. Sometimes a structure can look as though it has an alkyl group, but it is just drawn with a bend in the chain.

IUPAC systematic nomenclature

You can name an organic compound by following these steps:

- **Step 1:** Determine the longest chain of carbon atoms – this will be the '**main chain**' (or parent chain). Use the correct prefix based on the total number of carbon atoms (Table 1).
- **Step 2:** Locate the different functional groups and determine which end is nearest to an alkyl branch, functional group or double bond – this is your '**starting point**'. If a compound contains two or more of these, the order of priority from lowest to highest is: branch < halogen < double bond < hydroxyl/carboxyl. For example, if an organic compound has a double bond and an alkyl branch, the starting point carbon will be the one closest to the double bond.
- **Step 3:** Number the carbon atoms from the 'starting point' to the terminal carbon.

TABLE 1 Main chain prefixes

Number of carbons	Prefix
1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-

main chain
the longest hydrocarbon chain

starting point
the carbon that is the nearest to an alkyl branch, functional group or double bond; this is the first carbon

Study tip

When finding the starting point, a double bond or functional group takes priority over an alkyl branch if they are equal distance from a terminal carbon.

Study tip

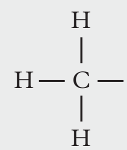
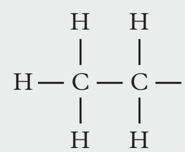
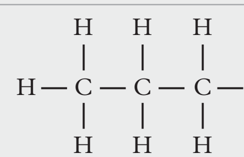
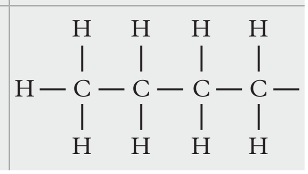
When you are numbering the longest chain, make sure you start from the end closest to a functional group, double bond or alkyl branch.

- **Step 4:** Identify and name any alkyl group branches (Table 2) and functional groups (Table 3).

TABLE 2 Functional group names

Functional group	Prefix/suffix
Alkane (C–C)	-ane
Alkene (C=C)	-ene
Haloalkane (–F, –Cl, –Br, –I)	Fluoro-, chloro-, bromo-, iodo-
Alcohol (–OH)	-ol or hydroxy- (as a prefix if there is another functional group that takes priority as a suffix)
Carboxylic acid (–COOH)	-oic acid

TABLE 3 Alkyl branch group names

Number of carbons	Alkyl branch name	Structural formula
1	Methyl-	
2	Ethyl-	
3	Propyl-	
4	Butyl-	

Study tip

When naming an organic compound, a hyphen (-) is used between numbers and words. A comma (,) is used between two numbers. No dash, comma or space is needed between two words.

Study tip

Prefixes are named in alphabetical order, regardless of their priority.

counting prefix
letters put at the start of a chemical name to indicate the number of branches or functional groups; e.g. 2 = di, 3 = tri, 4 = tetra

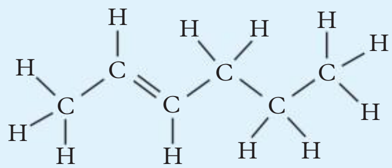
- **Step 5:** Determine the number of the carbon atom that each functional group is attached to. You will use these numbers in the name. They are separated from words by a hyphen (e.g. 1-chloro-).
- **Step 6:** When there are two or more of the same branch type or functional group in the molecule, use a **counting prefix** (e.g. di-, tri- and tetra-). Use a comma to separate the numbers (e.g. 2,2-dichloro-).
- **Step 7:** Put it all together in the correct order! Name all prefixes in alphabetical order.

Let's practise using these rules with Worked examples 8.2A, 8.2B and 8.2C.

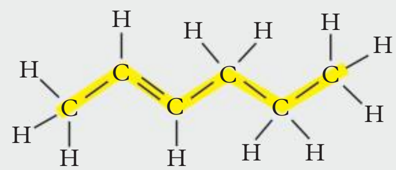
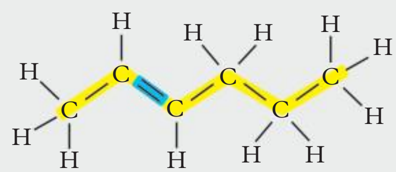
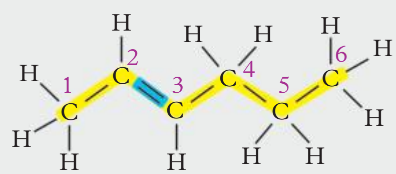
8.2A WORKED EXAMPLE

NAMING BASIC ORGANIC COMPOUNDS

Identify the IUPAC systematic name for the following organic compound.



Solution

Think	Do
Step 1: Determine the longest chain of carbon atoms.	The main chain is six carbons long (yellow highlight), so the prefix is ‘hex-’. 
Step 2: Locate the different functional groups and determine which end is nearest to an alkyl branch, functional group or double bond.	There is one double bond (blue highlight). 
Step 3: Number the carbon atoms from the ‘starting point’ to the terminal carbon.	The carbons are numbered from left to right (purple numbers), since the left-hand side carbon is closest to the double bond. 
Step 4: Identify and name any alkyl group branches or functional groups.	There is a carbon–carbon double bond, so the suffix is ‘-ene’.
Step 5: Determine the number of the carbon that each functional group is attached to.	The double bond connects carbons 2 and 3. We will always choose the lowest number, so the position of the double bond is carbon 2. This makes the corrected suffix ‘-2-ene’.
Step 6: Put it all together in the correct order! Remember that numbers and letters are separated by a hyphen (-).	Hex-2-ene

Study tip

In VCE Chemistry, you will only come across organic compounds with up to two different functional groups. This means that it’s important for you to remember the order of priorities: hydroxyl/carboxyl > double bond > halogen > alkyl branch.

8.2B Worked example
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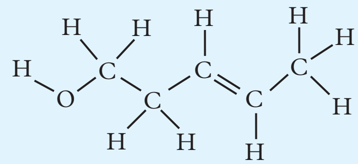
8.2B Worked example
Video demonstration

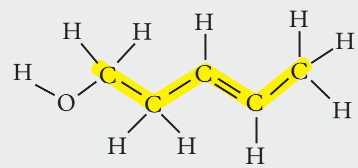
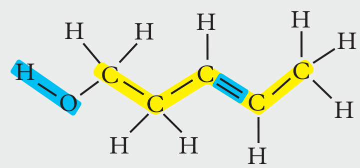
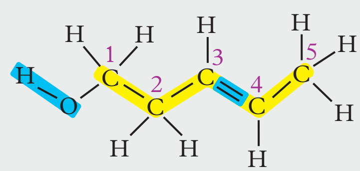
8.2C WORKED EXAMPLE

NAMING ORGANIC COMPOUNDS WITH MULTIPLE FUNCTIONAL GROUPS

Identify the systematic name of the following organic molecule.



Solution

Think	Do
Step 1: Determine the longest chain of carbon atoms.	The main chain is five carbons long (yellow highlight), so the prefix is 'pent-'. 
Step 2: Locate the different functional groups and determine which end is nearest to an alkyl branch, functional group or double bond.	There is a double bond and a hydroxyl group (blue highlight). 
Step 3: Number the carbon atoms from the 'starting point' to the terminal carbon.	The carbons are numbered from left to right (purple numbers). This is because the hydroxyl group takes priority over the double bond. The left-hand side carbon is closest to the hydroxyl group. 
Step 4: Identify and name any alkyl group branches or functional groups.	The carbon–carbon double bond makes this compound an alkene, so use the suffix '-ene'. The hydroxyl group makes it an alcohol, so we add the suffix '-ol'.
Step 5: Determine the number of the carbon that each of them is attached to.	The double bond connects carbons 3 and 4. We will always choose the lowest number, so the position of the double bond is carbon 3. This makes the corrected suffix '-3-ene'. The hydroxyl group is attached to carbon 1. Therefore, the corrected prefix is '-1-ol'.
Step 6: Put it all together in the correct order! Remember that numbers and letters are separated by a hyphen (-).	Pent-3-en-1-ol Note that the last 'e' on the '-ene' suffix is dropped.

Representations of organic compounds

There are four common ways to represent organic molecules:

- molecular formula
- structural formula
- semi-structural formula
- skeletal structure.

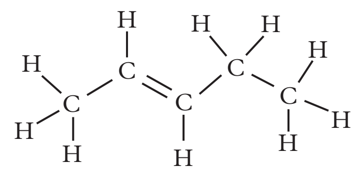
In this topic, you will learn how to draw the structural and semi-structural formulas for organic compounds. You don't need to know how to draw skeletal structures until Unit 4, but it is helpful to be able to recognise them now.

Study tip

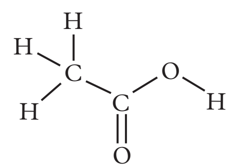
Go back to Chapter 3 if you need a refresher on molecular and structural formulas.

Structural formula

You might remember from Chapter 3 that a structural formula shows all of the atoms in a compound and how they are bonded together. For organic compounds, they can also show the position of the atoms and functional groups (Figure 2). This cannot be captured with just a molecular formula (e.g. C_5H_{10} and $C_2H_4O_2$).



Pent-2-ene



Ethanoic acid

FIGURE 2 The structural formulas of pent-2-ene (C_5H_{10}) and ethanoic acid ($C_2H_4O_2$), which show all the bonds, and the positions of double bond or functional groups

Semi-structural formula

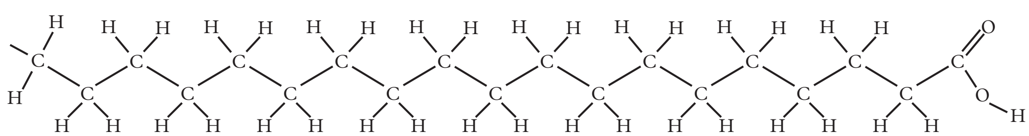
Semi-structural formulas show the position of all of the atoms and functional groups. However, it does not show the bonds. This type of representation is very helpful when you have long hydrocarbon chains. The structural formula can take up a lot of time and space to draw.

This can be simplified into a **condensed semi-structural formula.** Brackets are put around repeating units and the number of repetitions is indicated as a subscript. In Figure 3, there are 17 repeating CH_2 units in a row, so this is represented as $(CH_2)_{17}$.

semi-structural formula
a way of representing the structure of an organic compound, in which the positions of all the atoms and functional groups, but not the bonds, are shown

condensed semi-structural formula
a simpler form of the semi-structural formula, in which brackets are used to show repeating units

Structural formula



Semi-structural formula

$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2COOH$

Condensed semi-structural formula

$CH_3(CH_2)_{17}COOH$

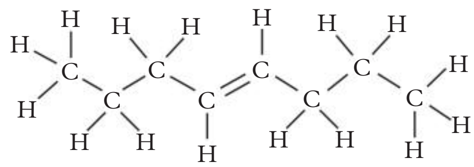
FIGURE 3 The structural formula, semi-structural formula and condensed semi-structural formula for a long-chain carboxylic acid ($C_{19}H_{38}O_2$)

Study tip

Some semi-structural and condensed semi-structural formulas show the position of carbon–carbon double bonds. For example: $CH_3(CH_2)_2CH{=}CH(CH_2)_2CH_3$. It is acceptable for you to draw them with or without the double bond.

Another example is oct-4-ene (Figure 4), in which the CH_2 units are separated by two CH units. In this case, the semi-structural formula is condensed in two parts. So, CHCH isn't condensed because it represents a double bond.

Structural formula



Semi-structural formula

$CH_3CH_2CH_2CHCHCH_2CH_2COOH$

Condensed semi-structural formula

$CH_3(CH_2)_2CHCH(CH_2)_2COOH$

FIGURE 4 The structural and skeletal formulas of oct-4-ene

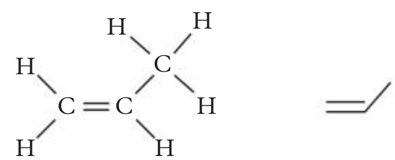
Hydroxyl (–OH) groups in the middle of the chain are represented by using brackets in the semi-structural formula. For example, butan-2-ol would be $CH_3CH(OH)CH_2CH_3$.

Skeletal structure

skeletal structure
a simpler form of the structural formula, in which carbon atoms are represented as vertices and hydrogen atoms bonded to these carbon atoms are not shown

The **skeletal structure** of an organic compound looks very similar to the structural formula (Figure 5). There are two main differences:

- the carbon atoms are represented as vertices (corners)
- any hydrogen atoms that are connected to a carbon are not shown.



Structural formula **Skeletal structure**

FIGURE 5 The structural, semi-structural and condensed semi-structural formulas for oct-4-ene

Comparing representations of organic compounds

The features of the different representations of organic compounds are shown in Table 4, using butanoic acid as an example.

Worked example 8.2D walks you through how to draw an organic compound if you are given its semi-structural formula. You can also challenge yourself to examine IUPAC names in Challenge 8.2.

8.2D Worked example
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8.2D Worked example
Video demonstration

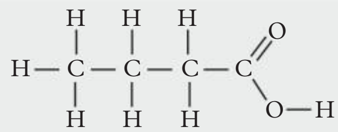
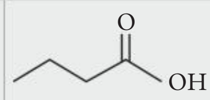
8.2 Challenge
Find me in your obook pro

TABLE 4 Representations of organic molecules, using butanoic acid as an example

	Molecular formula	Structural formula	Semi-structural formula	Skeletal structure
Number and type of atoms	✓	✓	✓	✓
Position of functional groups		✓	✓	✓

(continued)

TABLE 4 continued

	Molecular formula	Structural formula	Semi-structural formula	Skeletal structure
Type of covalent bonds		✓		✓
Example	$C_4H_8O_2$		CH_3CH_2 CH_2COOH	

8.2 CHECK YOUR LEARNING

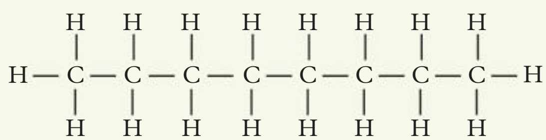
Describe and explain

1 Explain why the IUPAC systematic naming rules exist.

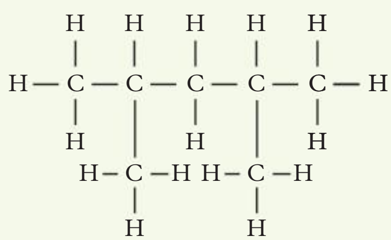
Apply, analyse and compare

2 Name the following structures according to the IUPAC naming rules.

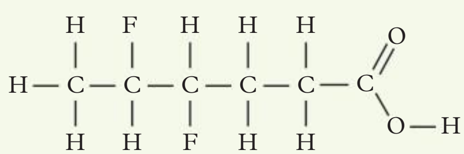
a



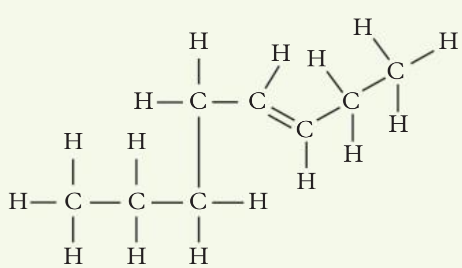
b



c



d



3 Convert the following semi-structural formulas into structural formulas.

a $CH_3CH_2CHClCH_2CH_3$

b $CH_2CHCH_2CH(OH)CH_3$

c $CH_3CH_2CH_2COOH$

d $CH_3(CH_2)_4CHCHCH_3$

e $CH_3CH_3CHCH_2CH(OH)CH_3$

4 Name the structures you drew in Question 3 by using the IUPAC naming rules.

5 Draw the structural and semi-structural formulas for the following molecules based on their IUPAC names.

a Heptan-4-ol

b 3,3-Dimethylbut-1-ene

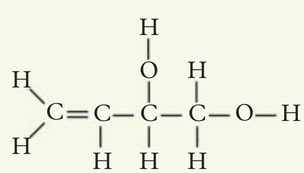
c 1-Chloro-3-fluorobutane

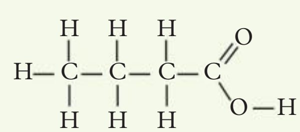
d 2,3-Dimethylpentanoic acid

Design and discuss

6 Discuss the advantages and disadvantages of molecular, structural and semi-structural formulas.

7 A chemist comes across two organic compounds, shown below. Discuss why she would not be able to identify them if she only had their molecular formula.





8.3 Renewable sources of organic compounds

KEY IDEAS

In this topic, you will learn that:

+ fossil fuels are the main source of organic compounds
+ fossil fuels are non-renewable, not sustainable and harmful to the environment
+ plant-based biomass is a renewable resource that could be an alternative source of organic compounds
+ relying less on fossil fuels leads to sustainable development.

Study tip

Remember that an organic compound is any compound consisting of carbon covalently bonded to other atoms.

Carbon is a very versatile element that can form the many organic compounds that we find in our everyday lives, including solvents, fuels, pharmaceuticals, clothing, food products, adhesives, dyes and paints. For most of history, organic compounds have been derived (made) from **fossil fuels**. However, fossil fuels are a limited resource, and their use makes them harmful to the environment.

fossil fuel
a mixture of hydrocarbon compounds that are mined from the Earth's crust; includes coal, natural gas, crude oil

FIGURE 1 Seventy per cent of fossil fuels are used for consumer products, like electronics, sports equipment, glasses, and the pen and ink you are using to make notes with now.

renewable
resources that are available in a continuous/unlimited supply or can replenish naturally

This means that it is critical for us to stop relying on fossil fuels and start looking at ways to replace them with more **renewable**, sustainable and environmentally friendly sources of organic compounds. In this topic, we will look at the problems with fossil fuels, what plant-based biomass is, and how it could be an important replacement for fossil fuels.

Fossil fuels

Fossil fuels are mixtures of hydrocarbons that we mine from the Earth's crust. They are made from decomposed plants and animals that lived and died millions of years ago. The three types of fossil fuels are coal, crude oil and natural gas.

Fossil fuels are **non-renewable**. We are consuming them at a faster rate than they are made naturally (over millions of years). This means that it will become more challenging to source fossil fuels to make the organic compounds that we use every day. As well as this, the burning of fossil fuels for energy releases **by-products** such as carbon dioxide and water vapour. These gases contribute to the **greenhouse effect** and **climate change**.

For these reasons, scientists are looking at renewable sources of organic compounds to replace fossil fuels. That's where plant-based biomass comes in.

FIGURE 2 The use of fossil fuels is very harmful to the environment.

non-renewable
resources that take an extremely long time to replenish and are available in a limited supply

by-product
a secondary product formed in a chemical reaction

greenhouse effect
a phenomenon where greenhouse gases (e.g. methane, carbon dioxide, water vapour) trap heat in the atmosphere and cause the Earth to warm

climate change
the gradual increase in the temperature of the Earth's surface, oceans and atmosphere, and consequent changes in climate generally caused by the enhanced greenhouse effect

Plant-based biomass

Biomass is material from living organisms or from organisms that were once alive. Therefore, like fossil fuels, biomass comes from plants and animals. However, unlike fossil fuels, biomass sources are renewable. This means that we can grow them at the same or a faster rate than we can consume them. Because of this, biomass is considered a renewable source of energy and organic compounds.

There are many different types of biomass, including:

- wood and wood processing wastes – firewood, woodchips, sawdust
- agricultural crops – corn, soybeans, sugar cane, algae
- waste materials – crop and food processing residues
- animal manure and human sewage
- recycled products – paper, cardboard, cotton, and wool products.

Generally, biomass is sourced from plants. **Plant-based biomass** contains stored chemical energy from the Sun. Plants convert this energy to glucose or other sugars, through the process of **photosynthesis**. You will learn about the chemical pathways through which biomass is converted to energy in Units 3 and 4.

FIGURE 3 Biomass can be sourced in many ways, such as from algae and recycled products.

FIGURE 4 Biomass can be turned into pellets and used as a fuel.

biomass
material from animals and plants that can be fermented to derive fuel and other organic products

plant-based biomass
specifically, biomass that comes from plants

photosynthesis
a chemical process in which energy from sunlight is converted to sugars and oxygen

Towards sustainability with biomass-derived organic compounds

Moving towards plant-based biomass as a primary source of organic compounds rather than fossil fuels directly addresses United Nations Sustainable Development Goal 12 'Responsible consumption and production'.

By replacing fossil fuels as an energy source and in consumer products, we will take an important step towards sustainable development. This also relates to the following green chemistry principle:

> *Use of renewable feedstocks: raw materials or feedstocks should be made from renewable (mainly plant-based) materials, rather than from fossil fuels whenever practicable.*

One example is the production of polyethylene. This is used in many products such as cling wrap, shopping and garbage bags, wire and cable insulation, plastic squeeze bottles and toys. Polyethylene is currently processed from fossil fuels, but we can replace this with by the **fermentation** of plant-based biomass. The steps for production of polyethylene from fossil fuels and biopolyethylene from biomass are shown in Figure 5, but you will not need to know the details of these processes until Units 3 and 4.

fermentation
a chemical process in which sugars are broken down in the absence of oxygen

By managing our use of resources and shifting towards more sustainable practices, we will then move from a linear economy to a circular economy.

Study tip

Drawing links between chemistry and sustainability is particularly important for VCE Chemistry. Go back to pages 32–35 in Chapter 1 to refresh your understanding of sustainable development, green chemistry principles, and linear and circular economies.

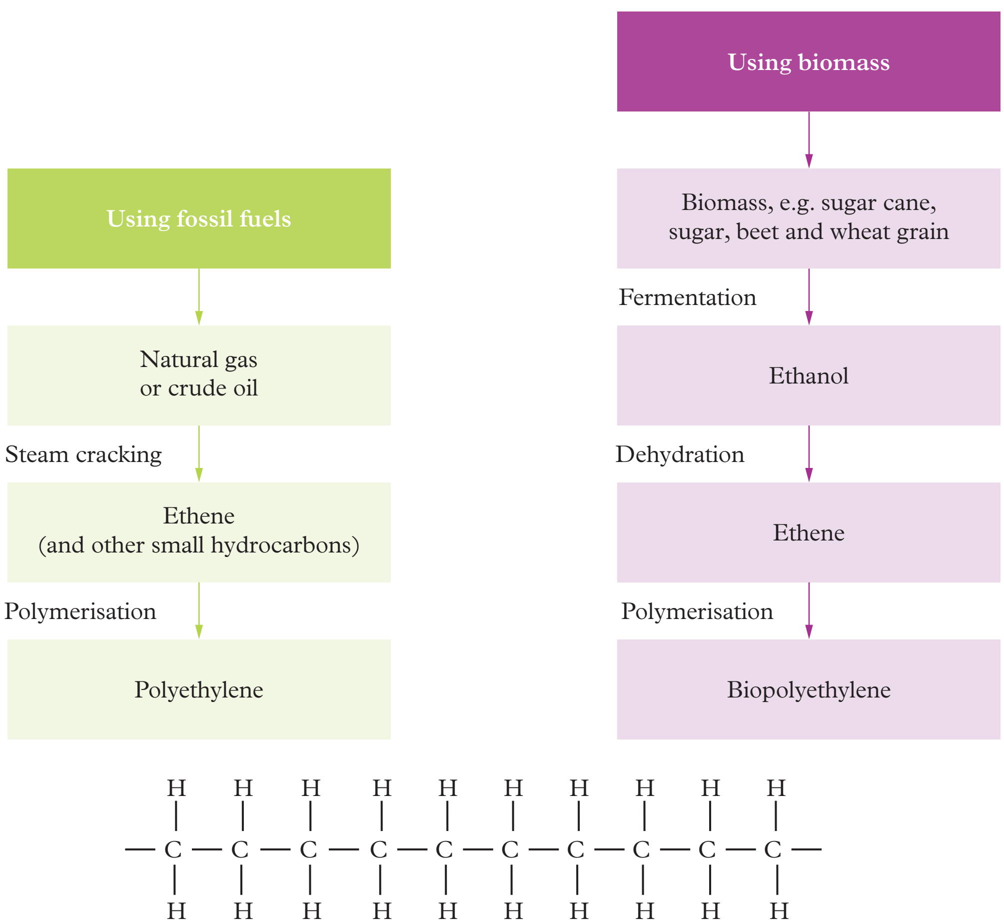


FIGURE 5 Polyethylene can be made from fossil fuels or biomass.

8.3 REAL-WORLD CHEMISTRY

Cleaner than clean energy – making a solar cell without fossil fuels

Solar energy is one solution to the world's reliance on fossil fuels. It involves using photovoltaic (or solar) cells to transform sunlight into energy we can use. However, many of the components of a solar cell are organic compounds sourced from fossil fuels. One major component is the protective plastic backing sheets or backsheets. These sheets protect the panels from moisture. They are currently made from polymers derived from petroleum.

One company has proposed using cotton rags and castor beans as a starting point for making these polymers. The cotton rags can be turned into a film of cellulose, a natural fibre. Castor beans can be processed to form nylon film. The cellulose and nylon film are blended to make BioBacksheets. These BioBacksheets appear to last as long (or even longer) than backsheets produced from fossil fuels and keep out just as much moisture. Another benefit is that, unlike fossil fuels, the cost of these alternative sources is low and does not fluctuate.

Apply your understanding

1 Compare the use of fossil fuels, cotton rags and castor beans.
2 Explain how this is an example of a linear economy shifting towards a circular economy.
3 Identify the green chemistry principle(s) that are related to using BioBacksheets.
4 Explain how the principle(s) you identified in Question 3 address the United Nations Sustainable Development Goals.

FIGURE 6 Solar panels are made from organic compounds sourced from fossil fuels.

FIGURE 7 Used cotton rags are being processed into natural cellulose to make backsheets for solar cells.

8.3 CHECK YOUR LEARNING

Describe and explain

1 Explain what a fossil fuel is and where it comes from.
2 Describe the different types of biomass.
3 Explain what plant-based biomass is.

Apply, analyse and compare

4 Two of the green chemistry principles covered in VCE Chemistry are:
- designing safer chemicals
- use of renewable feedstocks.

Explain how they relate to the concept of replacing fossil fuel products with plant-based biomass products.

Design and discuss

5 Research and select one new product that uses plant-based biomass to replace fossil fuels and discuss how it is an example of a linear economy changing to a circular economy.
6 The green chemistry principle 'Design for degradation' states that:

Chemical products should be designed so that at the end of their use they break down into harmless degradation products and do not persist in the environment.

Discuss how this green chemistry principle works with and relates to the green chemistry principle 'Use of renewable feedstocks'.

8.4 Organic compounds in everyday life

KEY IDEAS

In this topic, you will learn that:

- organic compounds are in many products that we use every day
- some organic compounds can have many benefits, but some also are hazardous to humans and the environment

In Topic 8.3, you learnt that organic compounds are all around us. By relying less on fossil fuels to produce many of these products, we can progress towards a circular and more sustainable economy. Despite this, one important thing to remember is that no matter where they are sourced from, organic compounds can have both benefits and hazards.

Products containing organic compounds

In this topic, we will explore the many different products that are made from organic compounds and discuss their harms and benefits to humans and the environment. This includes:

- solvents
- paints and dyes
- food additives
- cosmetics
- medicines
- pesticides
- car parts
- artificial hearts.

Study tip

Most organic solvents come from fossil fuels. By replacing fossil fuels with renewable feedstocks, we may be able to minimise the toxicity of the solvents. This addresses two green chemistry principles: 'Designing safer chemicals' and 'Use of renewable feedstocks'.

Organic compounds in solvents

The chemical reactions and pathways used to synthesise many of our everyday products require organic solvents. These reactions include the production of dyes, plastics, polymers, food products, textiles, printing inks, and pharmaceutical and agricultural products. Organic solvents are also in paints and paint products, adhesives, glue and cleaning agents.

Study tip

Designing biodegradable organic compounds that do not bioaccumulate addresses the green chemistry principle 'Design for degradation'.

Benefits of organic compounds in solvents

Organic compounds are useful in a diverse range of products. One example is toluene, which is cheap, easily sourced and used in many ways. It is a raw material for the production of nylon, as well as paints, lacquers, glues and adhesives. Toluene is also used as a cleaning and drying agent in the rubber, lumber, dry cleaning, motor, aviation and chemical industries.

Another example is cyclohexane, which is also a very widely used solvent. Cyclohexane is used in the production of perfumes, nylon, plastics and paint. It is also used as a solvent to dissolve lacquers, resins, fats, waxes, oils, bitumen and crude rubber.

Hazards of organic compounds in solvents

Long-term exposure to organic solvents is harmful to humans. Toluene vapours have a severe impact on the central nervous system and cause **neurological disorders**. Cyclohexane affects the central nervous system similarly. Contact with cyclohexane in liquid and vapour forms can also seriously damage the eyes.

neurological disorder
a disorder that affects the brain or nervous system

8.4 PRACTICAL: CONTROLLED EXPERIMENT

Which milk makes the strongest glue?
Go to your **Student obook pro**.

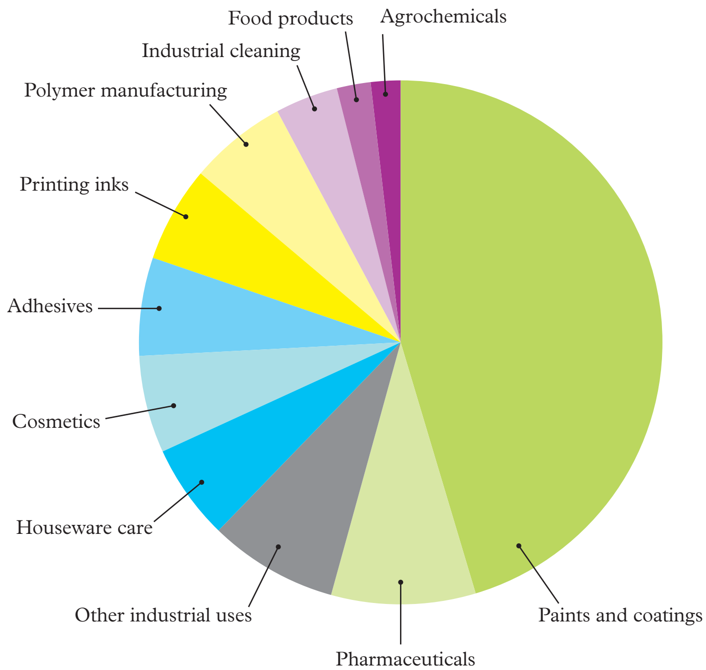


FIGURE 1 The average use of solvents across different industries and products.

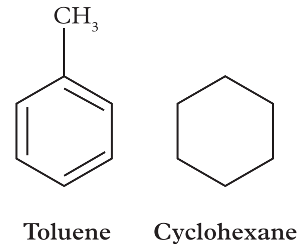


FIGURE 2 Skeletal structures for toluene and cyclohexane

Toluene evaporates quickly when exposed to air. It also breaks down quickly into other toxic chemicals, which can damage aquatic life and plants. Cyclohexane itself does not cause a lot of harm to the environment. However, it can **bioaccumulate** in fish that are exposed to contaminated waterways. On reaction with nitrogen dioxide, ozone or other organic compounds in the air, cyclohexane can produce a **photochemical smog**.

bioaccumulate
when a substance concentrates in the bodies of living things

photochemical smog
a haze formed when volatile organic compounds react with nitrogen dioxide, ozone or other volatile organic compounds in the air

Organic compounds in paints

Paints are liquid mixtures of organic compounds. Not only are the coloured **pigments** organic, but so are all of the components that make up a paint mixture.

paint
a liquid mixture that, when dry, enhances the appearance of objects or protects them from the environment

pigment
a chemical compound that gives colour

Benefits of organic compounds in paints

Organic compounds in paints have a variety of roles. These include:

- resins – are binders that hold together all the components and pigments, and allow the paint to bind to surfaces
- additives – enhance the properties of the paint; for example, helping the paint glide more easily onto a wall or canvas, decreasing the drying time, and making the paint mould- or scuff-resistant
- solvents – act as carriers to help the pigments and resins mix together
- pigments – give paint its colour and sheen.

Paints are commonly used to enhance the appearance of objects or protect them from the environment. Most obviously, you can find paint on the surfaces of buildings and furniture. They are also a popular medium for art.

FIGURE 3 Paints are mixtures of organic compounds.

Hazards of organic compounds in paints

Organic compounds in paints can be harmful to humans and the environment. In particular, some components of paint form vapours very easily. These **volatile organic compounds** (VOCs) can be very toxic to people who breathe them in regularly. Chronic (long-term) exposure can lead to symptoms such as headaches, memory loss, dizziness and problems with vision. VOCs in the air can also react with nitrogen oxides to form ozone. Formation of ozone at ground level can produce smog and cause diseases in plants.

volatile organic compound
an organic compound that can easily evaporate and become vapour

Study tip

There has been a shift towards sourcing organic compounds for paints from plant-based biomass instead of fossil fuels. Remember that this is a very important step towards Sustainable Development Goal 12 'Responsible consumption and production'.

8.4 REAL-WORLD CHEMISTRY

Azos are to dye for

Synthetic dyes are commonly used to colour fabrics. One particular group of synthetic dyes is azo dyes. Although they give long-lasting, bright and vibrant colours to fabrics, they are toxic and have been banned in manufacturing in many countries, including Australia and the European Union. Some of the 'fast fashion' countries still use azo dyes. Many of these fashion products are imported into Australia, which means you're still likely to encounter fabrics containing azo dyes.

There are two main types of azo dyes – azo direct dyes and azo reactive (acid) dyes. As their names suggest, azo direct dyes are applied directly onto the fibre and adsorb onto the material. Meanwhile, azo reactive dyes react with the fibre structure to form covalent bonds.

Azo direct dyes have been linked to increased risk of cancer, mutations and negative reproductive effects. Azo reactive dyes have been linked to increased allergy risk. Neither of the dyes is **biodegradable** in the environment. This means that they can accumulate in the marine food chain, then affect humans who consume food sourced from the sea. Therefore, azo dyes are classified as substances of very high concern.

biodegradable
the ability of a substance to be decomposed by bacteria or other living organisms, therefore avoiding pollution

Apply your understanding

1 Discuss how accumulation of azo dyes in the environment can affect humans and animals.
2 Discuss why some countries still use azo dyes, even when they are banned in so many other places.
3 Identify the green chemistry principle(s) that are related to using alternative dyes.
4 Explain how the principle(s) you identified in Question 3 address the United Nations Sustainable Development Goals.

FIGURE 4 Azo dyes come in many different vibrant colours.

Organic compounds in food additives

Organic compounds are common additives in foods. Their use is highly controlled, with many restrictions on what chemicals can be added to food. In Australia, this is regulated by Food Standards Australia New Zealand. We also follow rules set by the US Food and Drug Administration.

Benefits of organic compounds in food additives

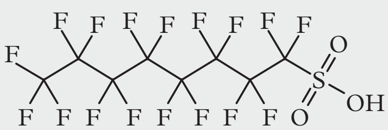
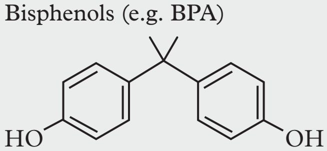
Food additives can be used to enhance the flavour, nutritional value, appearance and texture or safety of foods. This makes you, the consumer, more likely to buy the food product.

food additive
a compound that is added to food to enhance flavour, nutritional value, appearance, texture or safety

Hazards of organic compounds in food additives

All food additives go through testing to ensure their safety; however, it is not always possible to avoid contamination during manufacture. Hazardous organic compounds can be introduced from the packaging or other sources (e.g. the water that is used to grow or produce the product). Examples are shown in Table 1.

TABLE 1 Organic contaminants from food manufacture and their hazards

Organic contaminant	Source	Has the following benefits	Poses the following hazards
Perfluoroalkyl chemicals 	Greaseproof paper and cardboard packaging	• Repel water and oil • Are resistant to heat and chemical reactions • Useful surface coating	• If pregnant mothers are exposed, their babies can have low birthweight, with defects in their fertility, immune systems and thyroid • Do not biodegrade, so they can bioaccumulate and biomagnify in wildlife
Bisphenols (e.g. BPA) 	Lining of food and soda cans	• Protective coating for plastics • Tough, lightweight and shatter resistant • Biodegradable	• Can act like the hormone oestrogen and interfere with puberty and fertility • Can cause problems with the immune system and nervous system in humans • Harmful to aquatic organisms

Organic compounds in cosmetics

Cosmetics include all products that help you maintain or improve your appearance, such as shampoos, conditioners, hand washes, body washes, personal care products and make-up. They are available in many, many different brands, which contain many, many different ingredients. Therefore, it might not be surprising that many different organic compounds are found in cosmetics.

cosmetic
a product used to help maintain or improve appearance

Benefits of organic compounds in cosmetics

Organic compounds have all kinds of roles in cosmetic products. Some of these are:

- active ingredients that restore the hair or skin
- dyes and pigments that add colour
- compounds that make the products waterproof
- detergents to remove dirt and grease
- compounds that add fragrance.

Study tip

There is plenty of room for us to work towards the green chemistry principle 'Designing safer chemicals'. This will help to reduce harm to humans and the environment.

FIGURE 5 Organic compounds are found in many personal care products.

Hazards of organic compounds in cosmetics

Considering that many of these products are applied directly to your body, it is important to understand any potential risks. Like food products, the chemicals added to cosmetic products are highly controlled. Despite this, some hazardous organic compounds can still be found in cosmetics. We will look at two examples: phthalates and formaldehyde.

Phthalates are a class of chemicals that can be hidden in the list of ingredients under the term 'fragrance'. They are used as preservatives to increase the lifetime of products. Phthalates allow scents to linger and help lotions adhere better to the skin. However, if they get into the soil and groundwater, they can enter the root systems of plants and be consumed by animals and humans. In the body, they can disrupt the endocrine system, interfering with **metabolism** and normal hormone function.

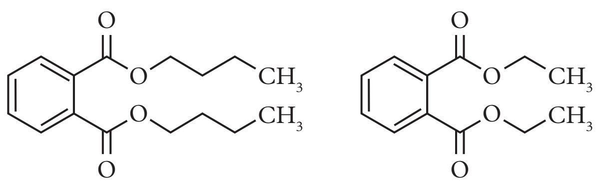
Formaldehyde is used as a **preservative** to protect personal care products from harmful bacteria. Unfortunately, formaldehyde is classified as a group 1 carcinogen, meaning that high exposure over a long period of time can cause cancer. It can also trigger asthma, cause dermatitis, hair loss, and irritation of the skin, eyes and nose. In the environment, formaldehyde is very toxic to aquatic life. It can also make small mammals sick, affecting their ability to breed and reducing their lifespan.

metabolism
the chemical processes in the body that convert energy from food into energy that our cells can use

preservative
a substance that prolongs the lifetime of a product by protecting it from harmful bacteria

Study tip

Bio-acetone is one example of a cosmetic product made from plant-based biomass (e.g. corn) instead of fossil fuels. This contributes to a circular economy through the green chemistry principle 'Use of renewable feedstocks'.



DBP **DEP** **Formaldehyde**

FIGURE 6 Skeletal structures of two phthalates, dibutyl phthalate (DBP) and diethyl phthalate (DEP), and formaldehyde

Organic compounds in medicines

When you think about organic compounds in **medicines**, natural and herbal remedies such as ginkgo and echinacea may come to mind first. However, the **active ingredient** in many modern medicines you know are also organic compounds or come from organic compounds. In fact, most medicines on the market are derived from nature, including aspirin (which comes from willow trees) and morphine (which comes from the opium poppy).

medicine
a compound or mixture of compounds used to treat disease

active ingredient
the compound in a medicine that is responsible for its medicinal effect

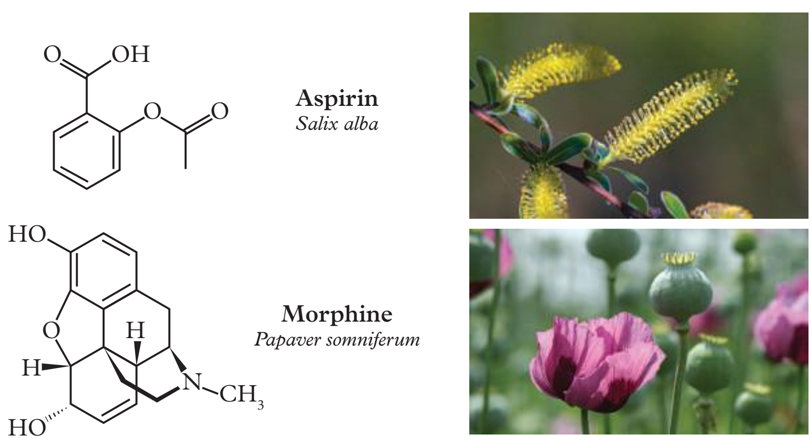


FIGURE 7 Aspirin and morphine are organic compounds that can be derived from plants.

Benefits of organic compounds in medicines

Organic compounds can have useful medicinal properties. For example:

- aspirin is a pain killer and blood thinner; dilute aspirin is also used in skincare products because it is a good treatment for acne
- morphine is used to treat chronic and acute pain.

You will learn more about medicinal chemistry in Units 3 and 4.

Hazards of organic compounds in medicines

The pharmaceutical industry is one of the most highly regulated industries in the world. This helps protect people from harm. Medicinal products are approved for market depending on their benefit to a patient compared with their risk. However, these risks can often be more severe than what we see with other controlled products like foods and cosmetics. This means that there are important hazards that we must be aware of. We will look at aspirin and morphine as examples.

Aspirin is derived from salicylic acid, which can cause severe eye, respiratory and skin irritation. It can also cause burns if it contacts wet skin. Ingestion of salicylic acid upsets the gastrointestinal system, causing symptoms such as nausea, vomiting and diarrhoea.

Morphine can cause similar gastrointestinal symptoms. It can also result in a hormone imbalance and is highly addictive.

Study tip

Many organic medicines can be sourced from plant-based biomass. This is more sustainable than synthesising them from non-renewable sources of organic compounds (e.g. fossil fuels). Moving towards natural, plant-based products addresses the green chemistry principles and encourages a circular economy.

Organic compounds in pesticides

Pesticides are any chemical substances or mixtures used to protect plants or animals. They help to control, repel, prevent or destroy unwanted plants or pests. Those that eliminate plants are called **herbicides** and those that kill pests can be **insecticides** or rodent killers. They also include chemicals that prevent disease in plants. You are likely to have come across pesticides such as bug spray to repel mosquitoes. Many of the active ingredients in pesticides are organic compounds.

pesticide
a compound or mixture of compounds used to protect plants or animals by controlling unwanted plants or pests

herbicide
a pesticide that specifically controls plants

insecticide
a pesticide that specifically controls insects

livestock
agricultural animals kept for produce such as meat, milk, eggs and fur

Benefits of organic compounds in pesticides

Pesticides are generally designed to target a wide range of invasive and toxic weeds or pests. For example, glyphosate is a commonly used herbicide in agriculture to protect crops such as fruits, vegetables, corn, cotton and canola. Chlorpyrifos is an insecticide that effectively kills many kinds of pests, such as termites, mosquitoes, roundworms and fire ants.

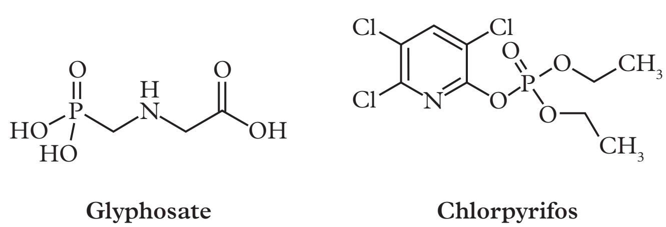


FIGURE 8 Skeletal structures for glyphosate and chlorpyrifos

Hazards of organic compounds in pesticides

Generally, the organic compounds found in pesticides are harmless to humans. However, this is not always the case. Some organic compounds can bioaccumulate. If the plants or animal feed sprayed with pesticide are eaten by **livestock**, they may transfer into cows' milk or meat. These chemicals can make their way into water systems and into aquatic animals. Residues may also contaminate fruit and vegetable crops.

enzyme
a substance produced in living things that speeds up chemical reactions

total artificial heart
a device that replaces the ventricles of a heart; used for patients whose ventricles are no longer functional

ventricle
a chamber of a heart

Chlorpyrifos can take weeks to years to break down. It bioaccumulates and can enter the food chain. It is toxic to many animals, so its presence could have disastrous effects on an ecosystem. Chlorpyrifos is also toxic to humans. It blocks an **enzyme** that controls messages between nerve cells and disrupts the nervous system.

Glyphosate breaks down in the environment and does not bioaccumulate. However, it poses potential risks to terrestrial and aquatic plants and birds. In humans, glyphosate exposure has been linked to respiratory dysfunction.

Organic compounds in artificial hearts

A **total artificial heart** is a pump that is surgically installed to replace diseased or damaged heart **ventricles**. The ventricles help pump blood from the heart around the body. A machine is connected to the artificial heart from outside the body (Figure 9). This controls the artificial heart by pumping compressed air into the tubing lines. It is recharged like a normal battery by plugging it into a power point or car charger.

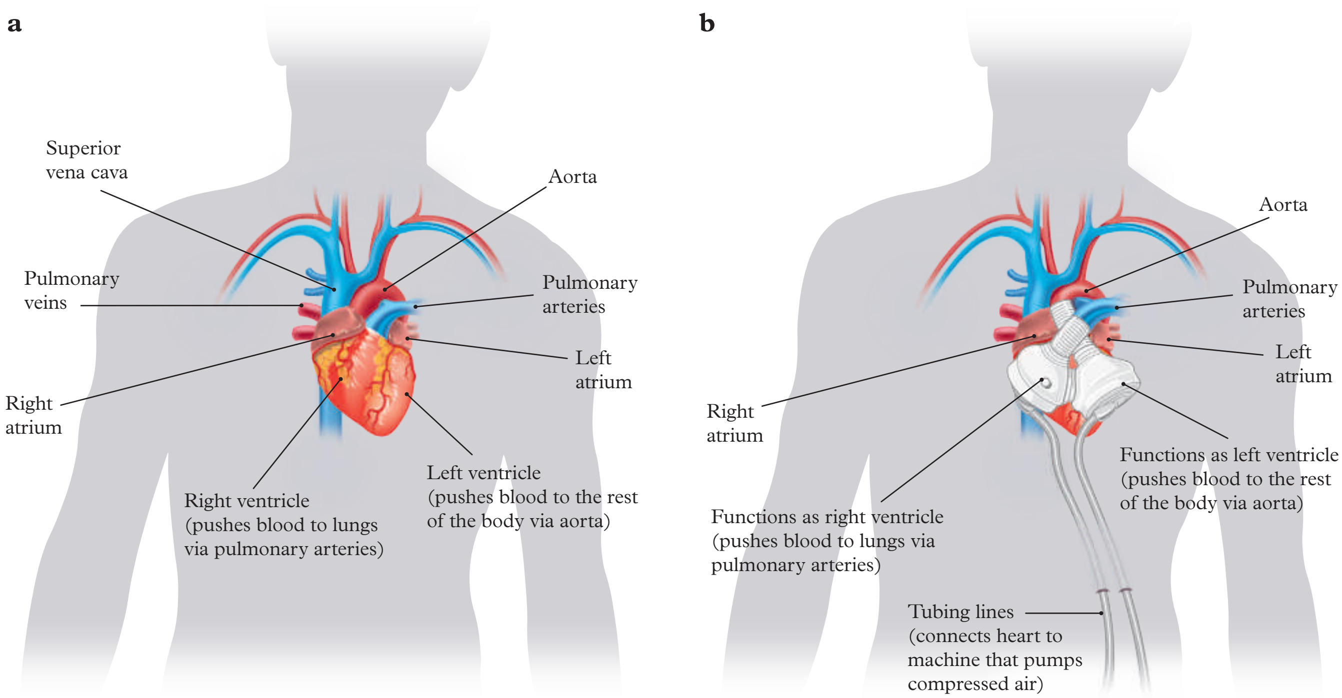


FIGURE 9 **a** The structure of a normal heart. **b** A total artificial heart, including the tubes that exit the body which are connected to a machine that powers and controls the artificial heart.

biopolymer
a polymer that occurs in nature or is made by a living organism, e.g. silk, DNA, cellulose, proteins; also known as natural polymer

biocompatible
not harmful to the human body

Benefits of organic compounds in artificial hearts

Scientists and researchers have made various prototypes and adaptations of artificial hearts. This includes trialling many different materials, such as plastics, metals, microporous materials, biosynthetic materials and **biopolymers**. The organic compounds that are used in artificial hearts are very high quality and must be **biocompatible** (not harmful to the human body). Two examples are:

- membranes of multisegmented polyurethane – a flexible material used to form the surfaces that are in contact with blood; can withstand many cycles of stresses and deformation
- polycarbonate urethane – used in the tubing of an artificial heart.

Hazards of organic compounds in artificial hearts

Artificial hearts are foreign objects introduced into the body. They can have a high risk of negative side effects. No matter how biocompatible the organic compounds are, there is always a risk of rejection and incompatibility in different people.

Research continues to search for organic compounds that are safer to use. One example is polycarbonate urethane, which appears to have less of an inflammatory response in the human body than some other polymers (e.g. ultra-high molecular weight polyethylene).

Most organic compounds used in artificial hearts are considered biochemically inert (non-reactive) and do not pose a threat to the environment. However, some chemicals used in the manufacturing of polyurethanes can be hazardous to the environment.

Study tip

By continually researching and finding safer alternatives to existing compounds, we are addressing the green chemistry principle 'Designing safer chemicals'.

8.4 SKILL DRILL

Using reasoning to evaluate ideas about sustainability

Key science skill: Analyse, evaluate and communicate scientific ideas

After reading through Topic 8.4, a student concludes that, in order to protect the environment, we need to stop using all chemicals that can bioaccumulate.

Practise your skills

1 Evaluate and identify the limitations of the student's conclusion.
2 Describe the implications of following the student's suggestion.
3 Discuss the impact of making judgments based on sociocultural, economic, political, legal and/or ethical factors and not solely on scientific evidence.

Need help evaluating and analysing scientific ideas? See Topic 1.8 (page 24).

8.4 CHECK YOUR LEARNING

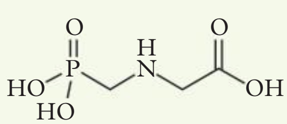
Describe and explain

1 Describe what an organic compound is.

2 Describe the difference between a natural organic compound and a synthetic organic compound.

Apply, analyse and compare

3 The structure of glyphosate is shown below.



a Explain why it is soluble in water.

b Discuss how this property would be beneficial for a compound in pesticides.

Design and discuss

4 Select two different organic chemicals discussed in this chapter that have benefits to society but also pose a hazard risk to humans and the environment. Discuss whether the benefits of each outweigh the hazards.

5 Discuss why the organic compounds added to food and cosmetics are so carefully controlled, but for pesticides and solvents they are not.

6 The green chemistry principle 'Design for degradation' states that:

Chemical products should be designed so that at the end of their use they break down into harmless degradation products and do not persist in the environment.

Discuss how this principle should be used to consider replacements for some of the chemicals disussed in this chapter.

CHAPTER 8 Review

Chapter summary

8.1

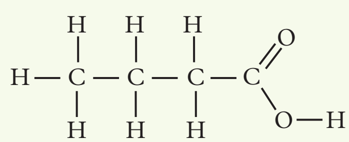
- Any compound that contains carbon is an organic compound. Hydrocarbons contain only carbon and hydrogen. Different functional groups can add unique properties to hydrocarbons.
- Families (or homologous series) of organic compounds have the same general formula, a similar structure, a trend to their physical properties and the same chemical properties.
- Alkanes are hydrocarbon chains connected by carbon–carbon single bonds. They have the general formula C_nH_{2n+2} and are non-polar.
- Alkenes are hydrocarbon chains that contain a carbon–carbon double bond. They have the general formula C_nH_{2n} and are also non-polar.
- Haloalkanes contain a –F, –Cl, –Br or –I in place of a hydrogen atom within the hydrocarbon chain. They have the general formula $C_nH_{2n+1}X$, where X is any halogen, and they are polar.
- Alcohols have a hydroxyl (–OH) functional group. They have the general formula $C_nH_{2n+1}OH$ and are polar.
- Carboxylic acids have a carboxyl (–COOH) functional group. They have the general formula $C_nH_{2n+1}COOH$ and are polar.
- Structural isomers are compounds with the same molecular formula, number and type of atoms, but a different arrangement of atoms and resulting properties.

8.2

- The IUPAC (International Union of Pure and Applied Chemistry) naming rules help to systematically name and identify organic compounds.
- Organic compounds can be represented by their molecular formula, structural formula, semi-structural formula or skeletal structure.

Molecular formula $C_4H_8O_2$

Structural formula



Semi-structural formula $CH_3CH_2CH_2COOH$

Skeletal structure

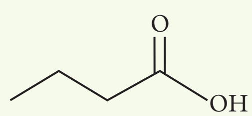


FIGURE 1 The molecular formula, structural formula, semi-structural formula and skeletal structure of butanoic acid.

- The molecular formula only gives the number and type of atoms, and not how they are bonded together.
- The structural formula shows the exact layout of an organic molecule, including all the bonds and atoms.
- The semi-structural formula shows the atoms and their positions, but not the bonds.
- The skeletal structure shows the exact layout of an organic molecule, but carbon atoms are shown as vertices and hydrogen atoms connected to the carbon atoms are not shown.

8.3

- Fossil fuels are mixtures of hydrocarbons made from decomposed plants and animals. They are a non-renewable resources found and mined from the Earth's crust.
- Biomass can come from plant or animal sources. They are classified as renewable because we can grow them at the same or a faster rate than we consume them.

8.4

- Organic compounds are found in products we use every day. Some of them are completely harmless, but others can pose significant health and environmental hazards.

Chapter checklist

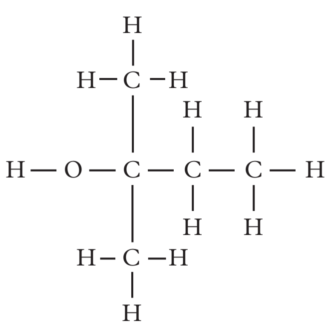
Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can …	Confidently	Mostly	Not really	Revision link
• identify different hydrocarbon compounds, including alkanes, alkenes, haloalkanes, alcohols and carboxylic acids	☐	☐	☐	Go back to Topic 8.1
• explain how the structure of a hydrocarbon compound relates to its properties	☐	☐	☐	Go back to Topic 8.1
• represent organic compounds using their molecular formula, structural formula and semi-structural formula	☐	☐	☐	Go back to Topic 8.2
• name compounds according to the IUPAC naming rules	☐	☐	☐	Go back to Topic 8.2
• describe the importance of using plant-based biomass as an alternative source of organic compounds that are traditionally derived from fossil fuels	☐	☐	☐	Go back to Topic 8.3
• describe the different benefits and hazards of organic compounds found in everyday products, such as synthetic fabrics, foods, medicines, pesticides, solvents, cosmetics, car parts, artificial hearts, adhesives, and dyes and paints	☐	☐	☐	Go back to Topic 8.4

Revision questions

Multiple choice

1 Identify the correct IUPAC name for the following molecule.



A 2-Methylbutan-2-ol

B 2-Ethylpropan-2-ol

C 3,3-Dimethylpropan-3-ol

D 1,1-Dimethylpropan-1-ol

2 Which of the following does *not* describe a homologous series?

A They only contain compounds with carbon and hydrogen.

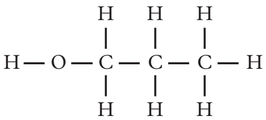
B They have the same general formula and differ by a $-CH_2$.

C There is a trend in the physical properties.

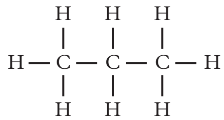
D They have the same chemical properties.

Revision questions

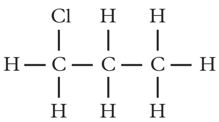
The following four molecules relate to Questions 3 and 4.



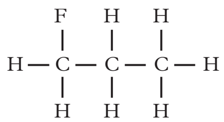
Propan-1-ol



Propane



1-Chloropropane



1-Fluoropropane

3 Identify the molecule that has the highest boiling point.

A Propan-1-ol
B Propane
C 1-Chloropropane
D 1-Fluoropropane

4 Identify the molecule that has the lowest boiling point.

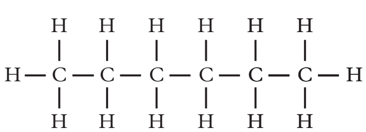
A Propan-1-ol
B Propane
C 1-Chloropropane
D 1-Fluoropropane

5 A compound has the molecular formula $C_{14}H_{28}$. Which hydrocarbon family does it belong to?

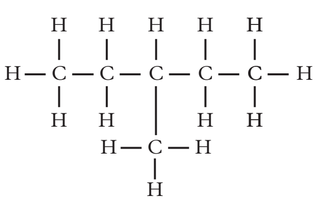
A Alkane
B Alkene
C Haloalkane
D None of the above

6 Structures I–IV are organic compounds.

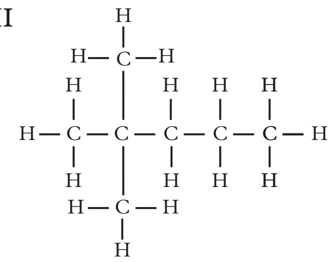
I



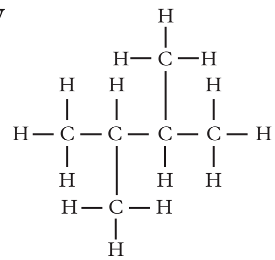
II



III



IV



Identify the compounds that are isomers.

A I and II
B II and III
C I, II and IV
D II, III and IV

7 The molecule with the semi-structural formula $CH_3CH(CH_3)CH_2CHClCH_2Cl$ has the IUPAC name:

A 4,5-dichloro-2-methylpentane.
B 2-methyl-4,5-dichloropentane.
C 1,2-dichloro-4-methylpentane.
D 4-methyl-1,2-dichloropentane.

8 The following are the IUPAC names of four compounds.

I 2-Chloro-3,3-dimethylpentane
II 3-Chlorohexane
III 2-Chloro-3-methylpentane
IV 2-Chloro-2,3-dimethylbutane

Identify the compounds that are isomers.

A I, II and III
B I, II and IV
C I, III and IV
D II, III and IV

Short answer

Describe and explain

9 Explain why carbon can form so many compounds.

10 Explain why the melting point of fluoromethane (CH_3F) is higher than that of methane (CH_4).

11 Hexane has a boiling point of 68.7°C and hex-1-ene has a boiling point of 63.0°C.

a Draw the structural formula of each compound.

b Explain the difference in their boiling points.

12 Explain why IUPAC rules state that '1-chloropropane' should include a number in the name, but 'chloroethane' does not.

13 Explain why ethanol has a higher boiling point than ethane.

14 Determine whether carboxylic acids can have isomers in which the position of the carboxyl functional group changes. Explain why or why not.

15 Explain how United Nations Sustainable Development Goal 12 'Responsible consumption and production' relates to the green chemistry principle 'Use of renewable feedstocks'.

16 The boiling points of two isomers of C_5H_{10} are shown in the table.

Isomer	Boiling point (°C)
Pent-1-ene	30
3-Methylbut-1-ene	20

a Draw the structural formula of each isomer.

b Explain the difference in boiling point.

Apply, analyse and compare

17 Look at the following molecular formulas and classify the compounds into the correct hydrocarbon family.

a C_3H_6

b C_5H_{12}

c C_4H_9OH

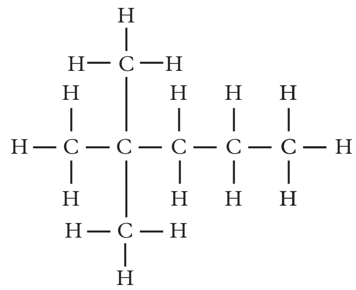
d CH_3F

e $C_{20}H_{40}$

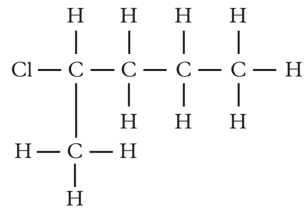
f $C_6H_{11}COOH$

18 Apply the IUPAC naming rules to name the following compounds.

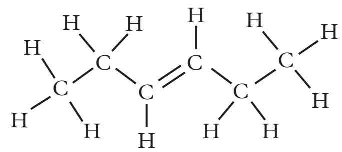
a



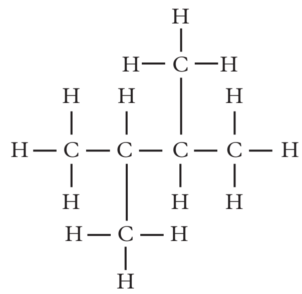
b



c



d



19 Draw the correct structural formula for:

a 2,3-dimethylpentane

b 3-ethylpentane

c pent-2-ene

d octan-1-ol

e pentanoic acid

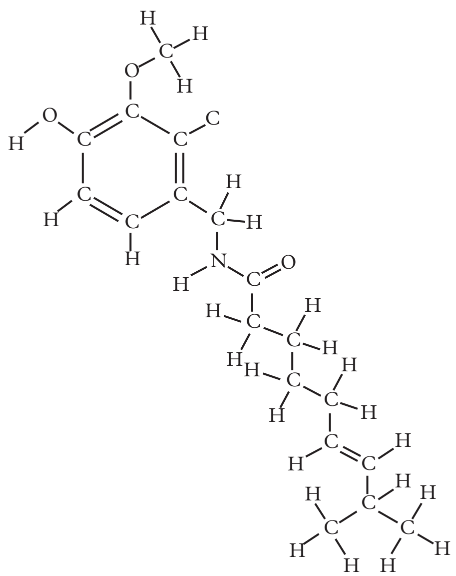
f 3,3-diflourohexane.

20 Convert the structures you drew in Question 19 into:

a molecular formulas

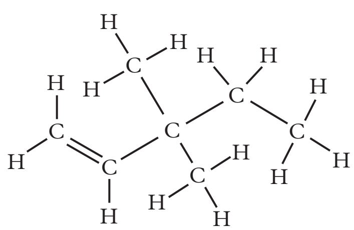
b semi-structural formulas.

21 Identify and name all of the functional groups in the following compound that you have learnt so far.

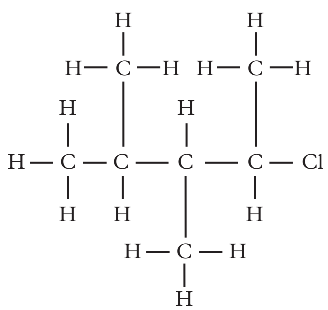


22 Apply the IUPAC rules to name the following compounds.

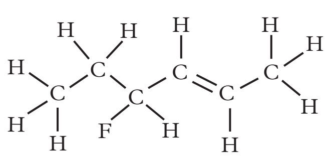
a



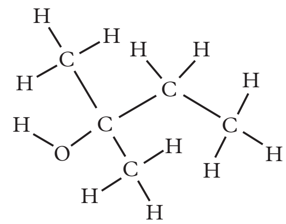
b



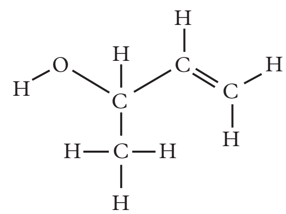
c



d



e



23 Using the IUPAC names provided, draw the correct structural formula of:

a 1-chloro-2-fluoro-4-methylpentane

b 3,3-dimethylbut-1-ene

c 2-methylbutan-2-ol

d 3-ethyl-3-methylpentanoic acid

e 3-chlorobutan-1-ol.

24 Convert the structures you drew in Question 23 into:

a molecular formulas

b semi-structural formulas.

25 There are many isomers with the molecular formula $C_5H_{11}Cl$.

a Draw all the isomers of $C_5H_{11}Cl$.

b Apply the correct IUPAC naming rules to name each isomer.

26 Compare the solubility of ethanol in water with that of octanol in water. Identify which is more soluble and explain why.

27 Compare the solubility of propane in water with propanoic acid in water. Identify which is less soluble and explain why.

28 The following are names of six compounds.

I 3,3-Dimethylpent-4-ene

II 3-Fluorohex-4-ene

III 2-Chlorohexan-6-oic acid

IV 3-Methylbutan-3-ol

V 2-Methyl-1-chloropent-1-ene

VI 1,2-Dichloro-3-methyloctan-4-ol

a Represent each compound by its structural formula.

b Determine whether the names are the correct IUPAC names.

c Explain why the other compounds are not named correctly.

d Write the correct IUPAC name for each structure that is incorrectly named.

Design and discuss

29 Design a flow chart that will help you use the IUPAC naming rules.

30 Five organic compounds, their molar masses and boiling points are shown in the table.

Molecule	Molar mass (g mol^{-1})	Boiling point (°C)
Propane	44.10	–42.1
Propene	42.081	–47.7
Fluoroethane	48.060	–37.7
Ethanol	46.07	78.0
Methanoic acid	46.025	100.5

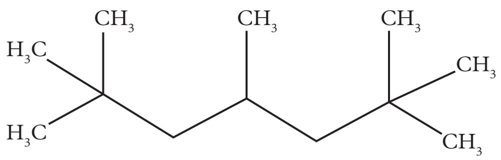
a Discuss how five molecules with such similar molar masses can have such different boiling points.

b Compare the boiling points of pentane and methanoic acid and explain why they are different.

c Compare the boiling points of ethanol and fluoroethane.

d Predict where the boiling point of propanol would be in this table. Explain your reasoning.

31 Isododecane was primarily synthesised from fossil fuels. It is now being derived from biomass.



Isodecane

Discuss how synthesis from biomass is an example of a linear economy transitioning into a circular economy.

32 Discuss how replacing raw materials sourced from fossil fuels with biomass relates to United Nations Sustainable Development Goal 12 'Responsible consumption and production'.

33 The green chemistry principle 'Design for degradation' states that:

Chemical products should be designed so that at the end of their use, they break down into harmless degradation products and do not persist in the environment.

Discuss how this green chemistry principle can be applied to a chemical such as chlorpyrifos, which is found in insecticides, takes years to break down and bioaccumulates.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 9

Polymers and society

KEY KNOWLEDGE

→ the differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers
→ the formation of addition polymers by the polymerisation of alkene monomers
→ the distinction between linear (thermoplastic) and cross-linked (thermosetting) addition polymers with reference to structure and properties
→ the features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching
→ the categorisation of different plastics as fossil fuel-based (HDPE, PVC, LDPE, PP, PS) and as bioplastics (PLA, Bio-PE, Bio-PP); plastic recycling (mechanical, chemical, organic), compostability, circularity and renewability of raw ingredients
→ innovations in polymer manufacture by condensation reactions, and the breakdown of polymers using hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy

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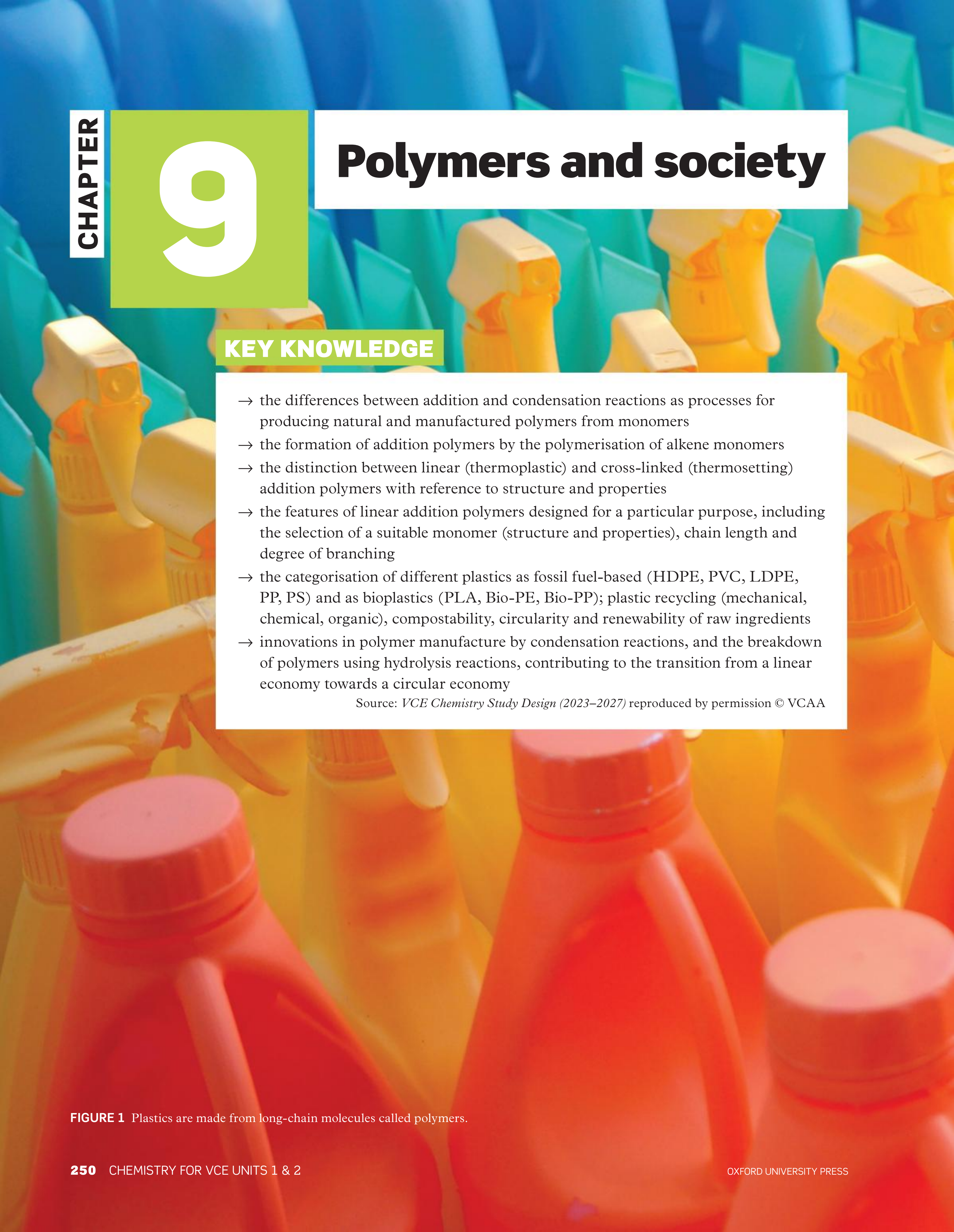


FIGURE 1 Plastics are made from long-chain molecules called polymers.

GROUNDWORK

In Chapter 9, you will learn about what polymers are, the structures and properties of polymers and how polymers are used in society.

This chapter will build on concepts you have learnt about in Chapter 8 and some concepts you may have covered in Year 10. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

9A What is an organic compound?

9A Groundwork resource
Organic compounds

9B What is a polymer?

9B Groundwork resource
Monomers and polymers

9C What is biomass and what is a renewable feedstock?

9C Groundwork resource
Renewable feedstocks

3D Why is it important to move towards a circular economy?

9D Groundwork resource
Circular economy

PRACTICALS

9.3 PRACTICAL: CONTROLLED EXPERIMENT	What effects do cross-links have on the properties of polymers?	Page 510
9.5 PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT	How can the design of copolymers be used to meet product requirements?	Page 512

9.1 Addition and condensation polymerisation reactions

KEY IDEAS

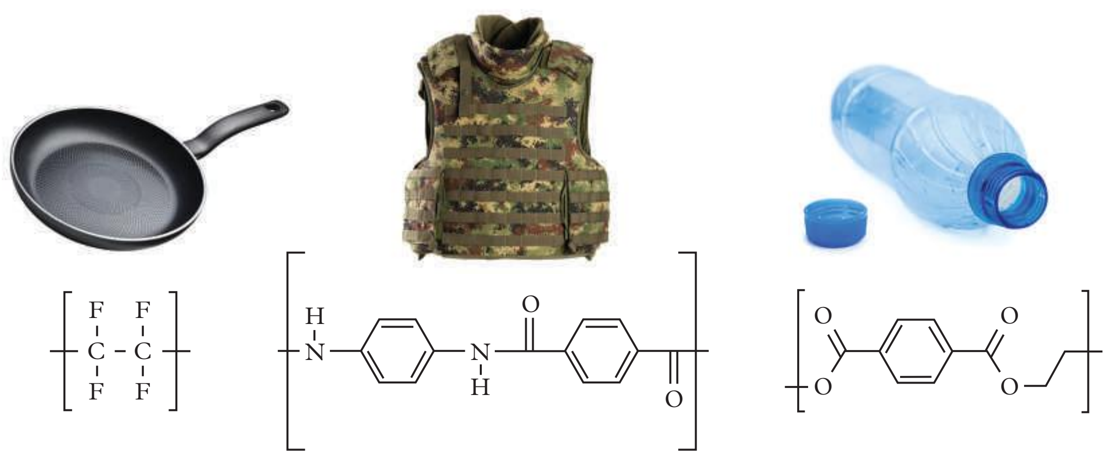
In this topic, you will learn that:

+ polymers can be synthetic or natural
+ addition polymers are made by addition reactions
+ condensation polymers are made from condensation reactions.

polymer
a large molecule made up of many repeating units called monomers

If you look around you, you may notice you are surrounded by different varieties of plastics. Plastics are made from **polymers**, which are large, long-chain molecules held together by covalent bonds. Many polymers are made naturally, such as components of fingernails and DNA; others are made synthetically – they are manufactured.

Manufactured polymers are often produced with properties to meet a specific purpose. For example, the polymer Kevlar makes a lightweight, heat-resistant and strong fibre. Because of its properties, Kevlar is used in bulletproof vests and racing car tyres. Alternatively, the polymer Teflon is waterproof and reduces friction, making it ideal for non-stick cookware coatings. The structure of polymers can be very similar or vastly different, depending on the properties and uses they are designed for.



Teflon Kevlar PET

FIGURE 1 Polymers and their different applications. Teflon is used in non-stick coatings, Kevlar in bulletproof vests and PET for water bottles.

Polymer structure

monomer
a small molecule that reacts with other molecules that are the same or different to form a polymer

Polymers are large molecules made by joining together small covalently bonded molecules called **monomers**. The term 'polymer' comes from the Greek words of *poly*, meaning many, and *mer*, meaning parts. Polymer length can range from the combination of two to millions of monomers.

The process in which monomers combine to form polymers is called **polymerisation**. The two types of polymerisation reactions we will explore in this topic are addition and condensation polymerisation.

polymerisation
a chemical reaction in which monomers are connected to form a polymer

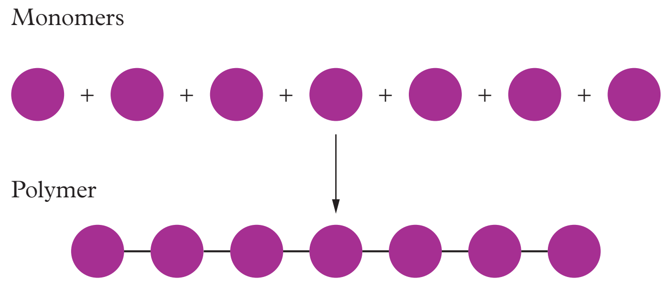


FIGURE 2 Monomers join together to form polymers.

Copolymers

The monomers that combine to make a polymer do not always have to be the identical molecules. When the monomers that make up a polymer are different molecules, the polymer they make is called a **copolymer**.

copolymer
A polymer made up of two or more different monomers

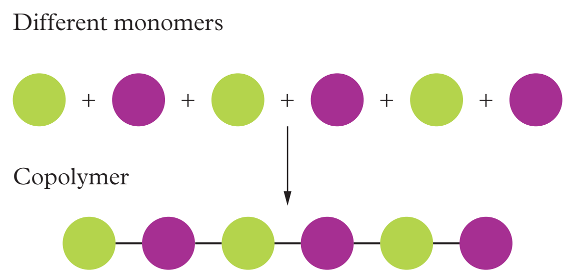


FIGURE 3 A copolymer is made of different monomers.

Addition polymerisation

Addition reactions occur when two or more molecules are added together to form one new molecule. In addition polymerisation reactions, polymers are formed from alkene monomers. For example, the **synthetic polymer** polyvinylchloride (PVC) is formed from chloroethene monomers. The addition polymerisation that forms PVC is shown in Figure 4.

addition reaction
a reaction in which two or more molecules add together to form one larger molecule and no other products

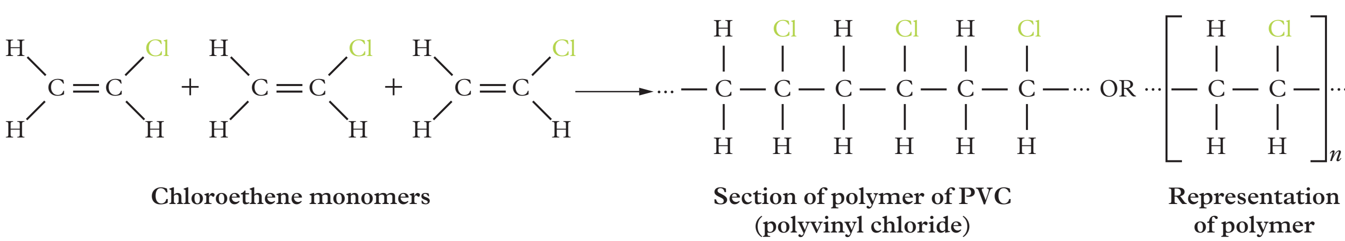


FIGURE 4 The addition polymerisation reaction between chloroethene monomers to form polyvinylchloride, commonly known as PVC

synthetic polymer
an artificially produced or synthesised polymer; generally derived from fossil fuels, but increasingly sourced from biomass

For addition polymerisation to occur, any double carbon–carbon bonds in the alkene monomers must first be broken. Breaking of these bonds often requires the help of a catalyst. Once the carbon–carbon double bonds are broken, the newly single bonded carbon atoms can form single carbon–carbon bonds with the other monomers. The linked single carbon–carbon bonds are what form the polymer chain.

Given the extreme length of some polymers, drawing out a polymer in full is not always feasible. Instead, we can use a representation of a section of the polymer.

This representation includes:

- large square brackets surrounding the basic unit of the polymer that repeats
- a subscript *n*, which indicates the number of times the section repeats.

In Figure 4 on the previous page, you can see how this notation has been used to shorten the drawing of PVC.

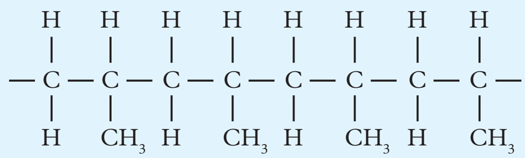
Study tip

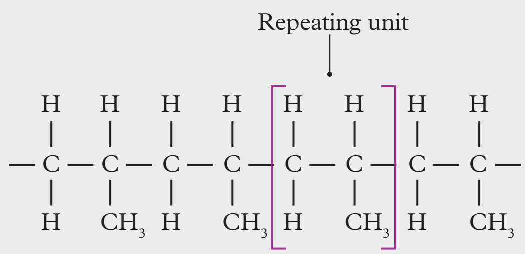
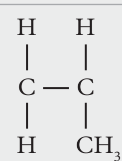
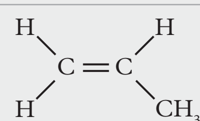
The monomers for an addition polymer need to have a double carbon–carbon bond. The monomers for a condensation polymer need to have two or more functional groups to react together.

9.1 WORKED EXAMPLE

IDENTIFYING AND DRAWING THE MONOMER OF AN ADDITION POLYMER

Identify and draw the monomer that forms the following addition polymer.



Think:	Do:
Step 1: Identify the repeating unit of the polymer.	
Step 2: Draw the repeating unit on its own.	
Step 3: Draw a double bond between the two carbons of your repeating unit; this will give you the monomer.	
Step 4: Identify and write the name of the monomer from your drawing in step 3.	Propene

Condensation polymerisation

condensation reaction
a reaction in which two molecules with functional groups react to form a larger molecule and a small molecule by-product

Condensation reactions occur when two molecules combine to form a larger molecule while also producing a small by-product (such as water). In condensation polymerisation reactions, polymers are formed from monomers that have at least two functional groups per molecule. The functional group of one monomer reacts with a functional group on another monomer and produces a by-product, which breaks off. Covalent bonds can then form between monomers and link the polymer together.

Natural condensation polymers

Many condensation polymers, including polypeptides, starch and cellulose, are formed naturally. You will study these condensation polymers in greater detail during Unit 4, but for now you can see how the functional groups on their monomers allow the condensation polymerisation to take place.

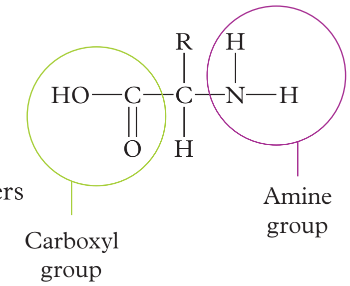


FIGURE 5 Amino acids, which have carboxyl (COOH) and amine (NH_2) functional groups, are the monomers that form polypeptides.

Polypeptides

One example of a group of naturally formed condensation polymers is polypeptides (proteins). Figure 6 shows how amino acid monomers react to produce a polypeptide through condensation polymerisation. The NH_2 of one amino acid monomer reacts with the COOH of another monomer to produce water (the by-product). A covalent bond then forms between the outer carbon and nitrogen atoms of the monomers, linking together the polypeptide polymer.

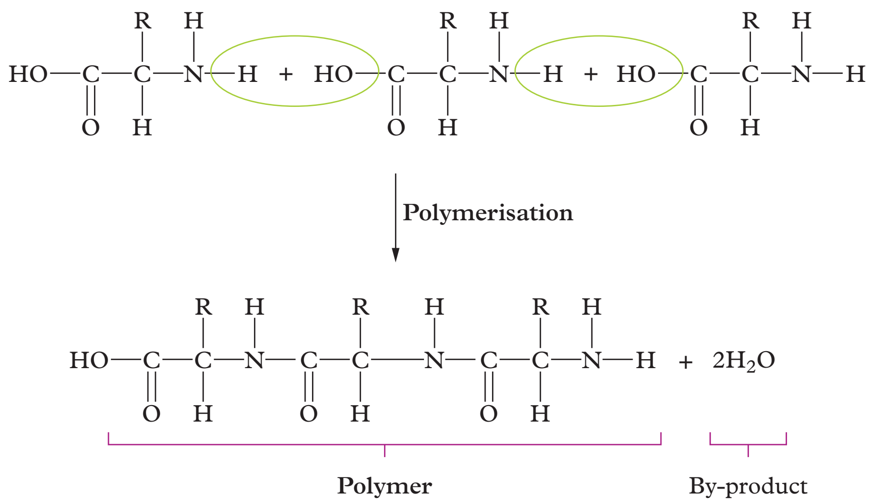


FIGURE 6 Amino acids undergo condensation polymerisation to produce a polypeptide.

Starch and cellulose

Starch and cellulose are also formed from condensation reactions. In Figures 7 and 8, you can see they are both made from glucose monomers.

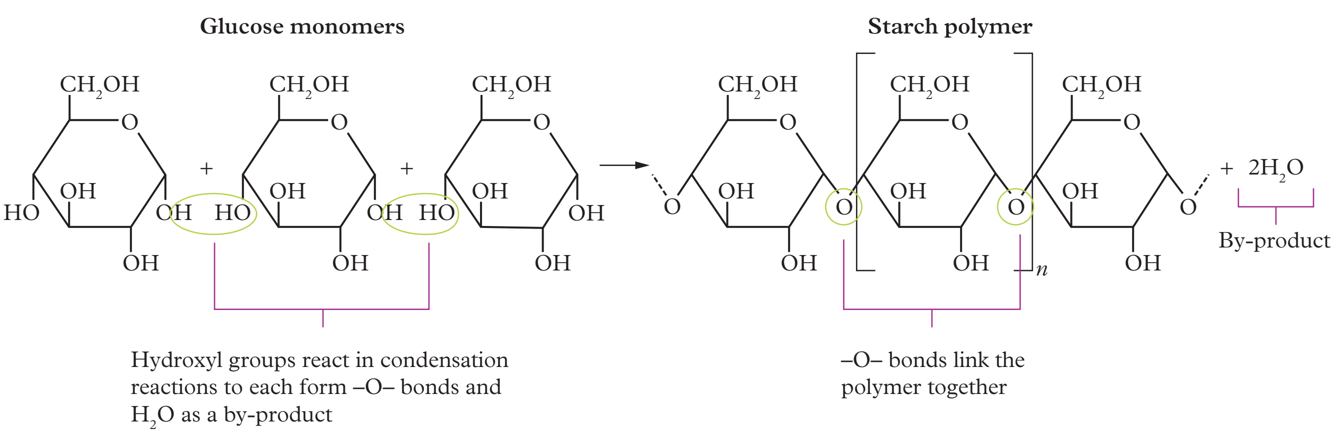


FIGURE 7 Glucose monomers react to form the natural polymer starch.

The glucose monomers that form starch and cellulose are isomers; this is why they give two different polymers. However, the same reaction forms both polymers. Each glucose monomer has multiple hydroxyl (OH) groups, which react to form water as a by-product and –O– bonds between monomers. The –O– bond links create either the starch or the cellulose polymer.

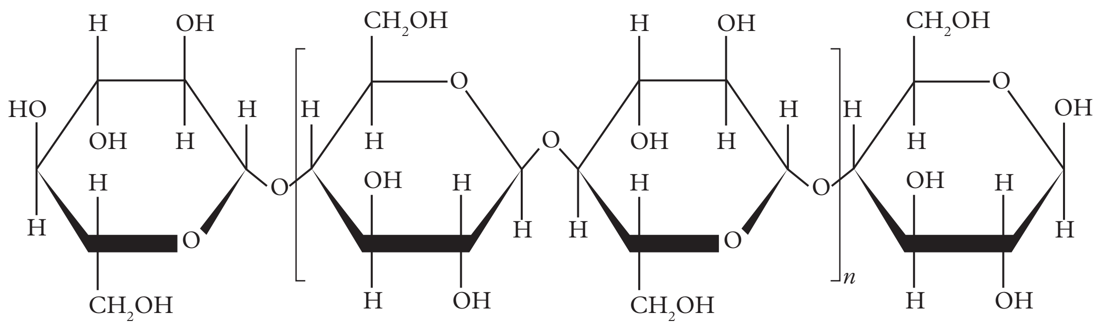


FIGURE 8 The structure of the natural polymer cellulose

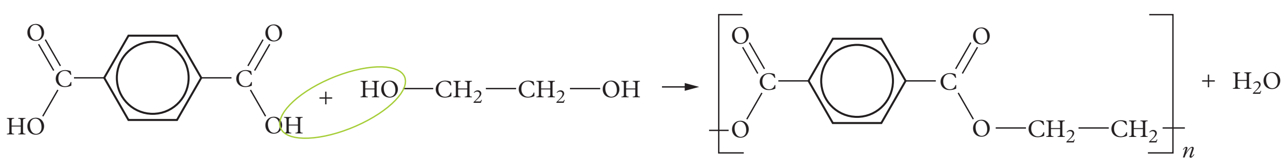
> **Study tip**
> A condensation reaction is a reaction between two functional groups and produces a small molecule as a by-product.

Manufactured condensation polymers

Condensation polymers can also be manufactured or produced synthetically. Manufactured condensation polymers are essential in many products we use or interact with daily. Two manufactured condensation polymers we will look at in greater detail are polyethylene terephthalate and nylon.

Polyethylene terephthalate

Polyethylene terephthalate (PET) is a strong, lightweight manufactured polymer often used in food packaging. PET is a copolymer, meaning its monomers have different formulas. Carboxylic acid and hydroxyl functional groups on monomers react in condensation reactions to form water and –COO– or ester bonds that link polymer chain. PET is one example of a polyester, a polymer that contains ester functional groups in their main chain.



Terephthalic acid **Ethylene glycol** **Polyethylene terephthalate (PET)**

FIGURE 9 PET is formed from multiple condensation reactions between alternating monomers.

Nylon

Nylon is another example of a manufactured condensation polymer. The first commercial uses of nylon include toothbrush bristles, women's stockings and parachutes for World War II. Many different nylon polymers exist, each with their own specific combination of monomers.

Like PET, nylon is a copolymer. One reaction to produce nylon is shown in Figure 10. In Figure 10, one monomer (adipic acid) has two carboxyl functional groups, and the other monomer (hexamethylenediamine) has two amine functional groups. These functional groups react to form a –CONH– between each of the units in the polymer, producing water as a by-product.

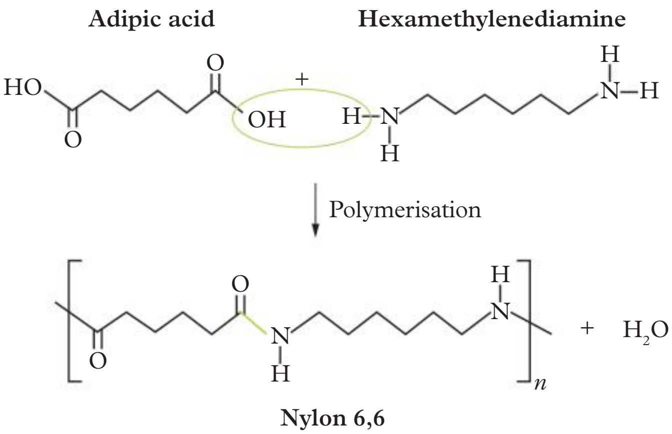


FIGURE 10 The condensation polymerisation reaction to form nylon

FIGURE 11 Nylon was used to make parachutes for World War II.

9.1 CHECK YOUR LEARNING

Describe and explain

1 Explain why monomers can be called the building blocks of polymers.

2 Describe the condensation polymerisation reaction between glucose monomers.

3 Describe the differences in the monomers that are used in condensation and addition polymerisation reactions.

Apply, analyse and compare

4 Analyse the two polymerisation reactions I and II shown below.

  a Identify which type of polymerisation reaction is occurring – addition or condensation.

  b Discuss the differences between the two reactions.

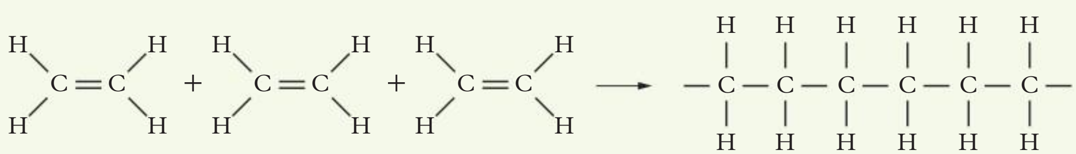
5 Compare the reactions to form a natural condensation polymer and a synthetic condensation polymer. Discuss the similarities and differences in the polymerisation reactions.

Design and discuss

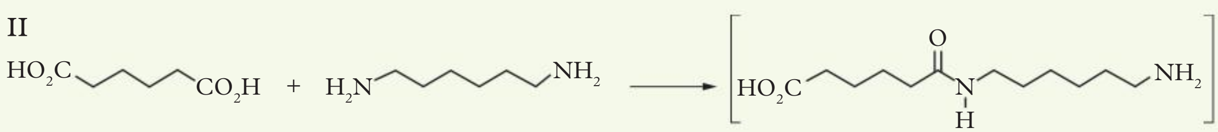
6 Discuss what would be required for an addition copolymer to be formed.

7 Draw the polymer that is formed from an addition polymerisation reaction between three ethene monomers.

I



II



9.2 Formation of addition polymers

KEY IDEAS

In this topic, you will learn that:

- alkenes can form addition polymers because they have a carbon–carbon double bond
- addition polymers have different properties and uses based on their monomers
- addition copolymers can be formed from two or more monomers and have properties from each monomer.

Addition polymers

In the previous topic, you learnt that alkenes react in addition reactions to form addition polymers. In this topic, we will look at a variety of alkene monomers and the reactions they undergo to form different addition polymers and addition copolymers.

Polyethylene

The simplest of the alkenes is ethene. The addition reaction of ethene monomers forms a polymer called polyethylene. In this reaction, the double carbon–carbon bond in ethene monomers breaks and the new monomers that now have single carbon–carbon bonds covalently bond to form polyethylene. The addition polymerisation reaction to form polyethylene is shown in Figure 1.

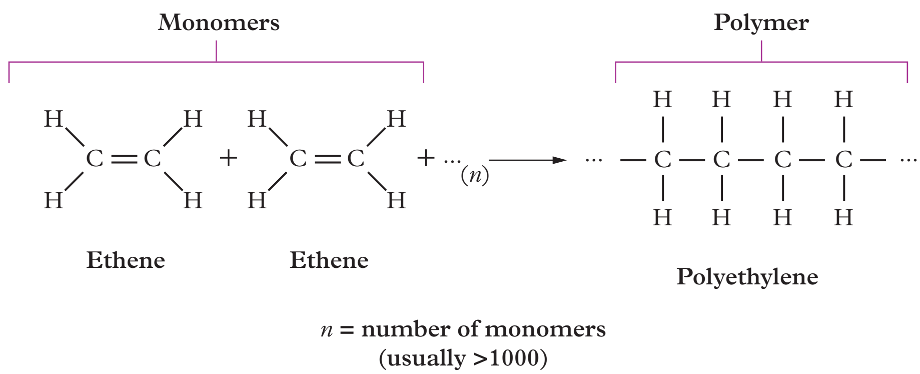


FIGURE 1 The addition reaction of the monomer ethene to form polyethylene

There are many variations of polyethylene. Later in the chapter you will look at various types of polyethylene and their different structures and properties.

Addition polymers from alkenes

There are hundreds of simple and complex addition polymers. Table 1 summarises a variety of simple addition polymers, their monomers, properties and uses.

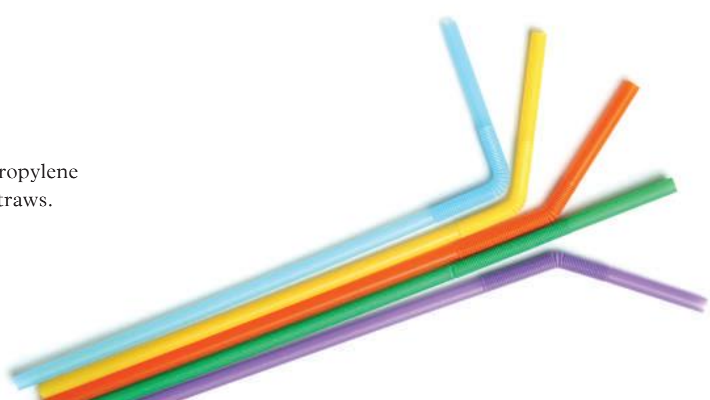
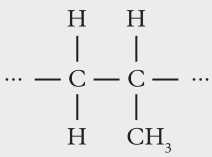
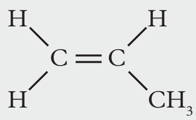
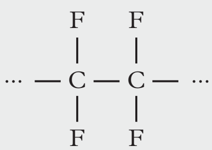
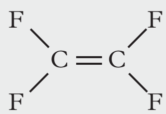
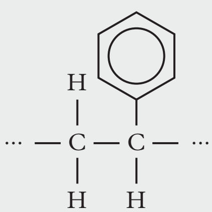
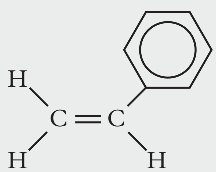
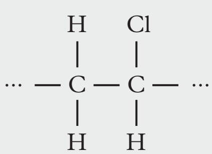
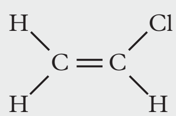
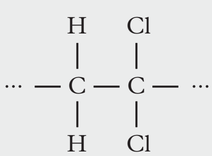
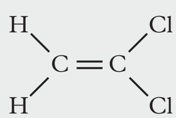
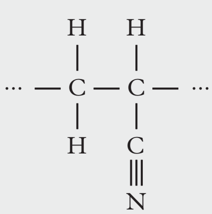
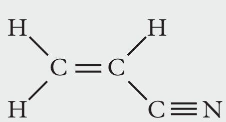


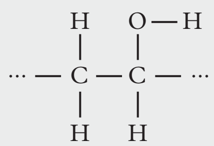
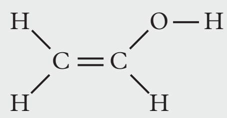
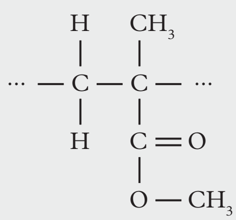
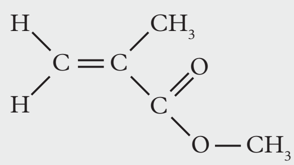
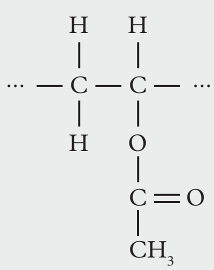
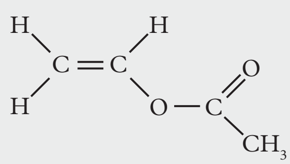
FIGURE 2 Polypropylene is used to make straws.

TABLE 1 Addition polymers

Polymer	Monomer	Properties	Uses
Polypropylene (PP) 	Propene 	• Durable • Flexible and easy to mould • Clear but can have pigment added	• Straws • Nappies and sanitary products • Disposable cups and cutlery • Packaging • Artificial grass and carpets
Polytetrafluoroethylene (Teflon) 	Tetrafluoroethylene 	• Non-stick • Insoluble in all common organic solvents • High melting point of approximately 327°C • Outstanding thermal, electrical and chemical resistance	• Non-stick coating for pots and pans • Low-friction bearings, slides and gear plates • Electrical insulator in some circuit boards • Chemically resistant pump parts and valves
Polyphenylethene or polystyrene 	Phenylethene or styrene 	• Chemically resistant to dilute acids and bases, no odour or taste • Excellent thermal and electrical insulators • Hard but not brittle • Readily soluble in organic solvents • Poor oxygen and UV resistance	• Toys • Packaging • Food packaging • Disposable cups • Cosmetic containers • Insulation foam • Building blocks
Polyvinylchloride (PVC) 	Chloroethene or vinyl chloride 	• Clear but can be dyed • Rigid and hard • Low impact strength • Compatible with plasticisers - to improve flexibility • Poor heat and light resistance	• Water and sewage pipes • Window frames • Wire and cable insulation • Footwear • Sporting goods • Toys
Polyvinylidene chloride (PVDC) 	Dichloroethene 	• Excellent barrier against oxygen, water and aromas • Chemically resistive to alkalis and acids • Insoluble in oil and organic solvents	• Plastic food wrap • Clothes • Tape • Shower curtains • Filters • Doll hair • Shoe insoles
Polypropene nitrile (acrylic) 	Propene nitrile 	• Strong • Can be formed into fibres • UV resistant	• Very hard, coarse fabric used for awnings, soft tops of cars and in brake linings • Used to reinforce concrete • Often used as a copolymer and made into a softer fabric used in clothing

(continued)

TABLE 1 Continued

Polymer	Monomer	Properties	Uses
Polyvinyl alcohol (PVOH) 	Ethenol or vinyl alcohol 	• Hard • Colourless • Odourless • Highly water soluble • Biodegradable • Resistant to solvents and oils	• As a binder, thickening agent, emulsifier in paper • Adhesives can be remoistened • Cosmetics • Food and pharmaceuticals • Water soluble packaging material • Medical application
Polymethyl methacrylate 	Methyl methacrylate 	• Stiff • Hard • Excellent weatherability • Transparent and colourless • Easily moulded	• As Perspex – a lightweight and shatter resistant replacement for glass, signs and brake lights for cars • Also used as a copolymer to improve polymer properties
Polyvinyl acetate (PVA) 	Vinyl acetate 	• Reasonably inexpensive to synthesise • Adheres and bonds to many different porous surfaces • Non-toxic • Odourless • Good light stability, meaning it is resistant to yellowing • Sensitive to water	• As an adhesive • As a plasticiser and thickener for paints, textile finishes, plastics, cement and chewing gum

Addition copolymers

Copolymers, which are made from two or more different monomers, can also be produced by addition reactions. By having a combination of different monomers, different properties are conferred on the polymer.

Acrylonitrile–butadiene–styrene (ABS) and polyvinyl chloride acetate (PVCA) are two examples of copolymers (Figure 3).

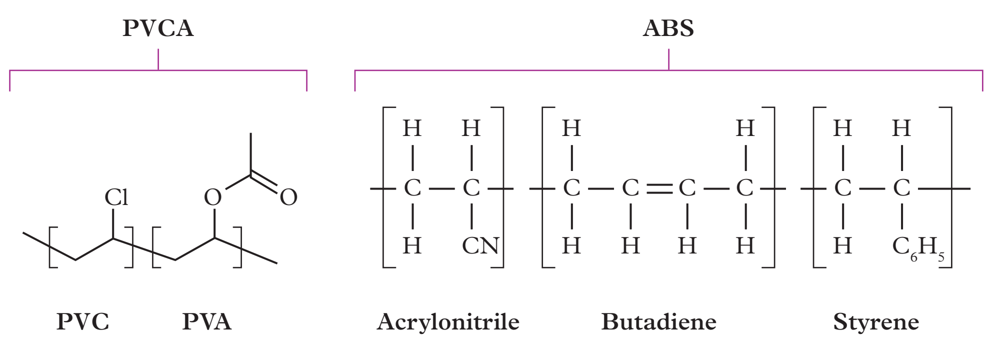


FIGURE 3 ABS and PVCA are both addition copolymers.

Each monomer brings a specific property to the polymer. For example, the PVC monomer brings durability, but has poor UV resistance, so adding the PVA monomer makes the PVCA polymer better in this respect. You can read more about ABS in Real-world chemistry 9.2.

9.2 REAL-WORLD CHEMISTRY

Addition copolymers

A well-known addition copolymer is ABS (acrylonitrile–butadiene–styrene).

The ABS monomers have the following properties:

- acrylonitrile: ageing, chemical and heat resistance
- butadiene: impact strength, low temperature properties, toughness
- styrene: lustre, mouldability, rigidity.

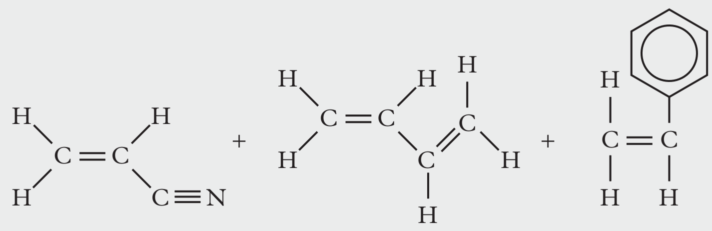


FIGURE 4 The monomers that form the polymer ABS

ABS is used to make Lego blocks, lamps and mirror housings and is used in automotive consoles and panels. ABS is also used in household electronics such as computer keyboards, printer housings, small appliances and wall plugs. Food-grade ABS can be used in kitchen appliances and food processing equipment.

The ABS copolymer is generally opaque, cheap to make, strong, durable and heat resistant.

Apply your understanding

1. Identify which properties of ABS come from each of its monomers.
2. ABS is used to produce Lego. Discuss the properties of the toy bricks that are derived from each monomer.
3. Discuss how changing the proportions of each monomer could change the properties of the ABS polymer.

FIGURE 5 ABS is used to make Lego blocks

9.2 CHECK YOUR LEARNING

Describe and explain

1 Explain what must happen to the double carbon–carbon bond of alkene monomers before the monomers combine to make addition polymers.

2 Identify the properties of polyvinyl alcohol (PVOH) that would benefit its use to make dissolvable wrapping for dishwashing tablets.

Apply, analyse and compare

3 Analyse the section of an addition polymer in Figure 6 and identify the monomer that is used to synthesise it

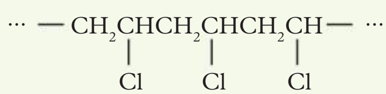
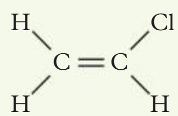


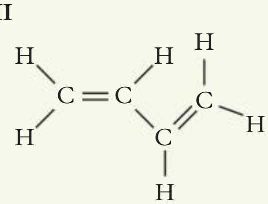
FIGURE 6 An addition polymer

4 Analyse the following molecules.

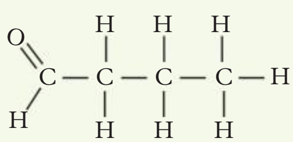
I



II



III



a Identify if they can undergo addition polymerisation reactions.

b If they can undergo addition polymerisation reactions, draw a section of the polymer they would form.

5 Calculate the molecular mass of the addition copolymer PVCA that is made up from 20 000 monomer units, of which 20% are vinyl acetate and 80% are vinyl chloride.

Design and discuss

6 The addition copolymer polyvinyl chloride acetate (PVCA), as shown in Figure 7, is made from the two monomers PVC and PVA.

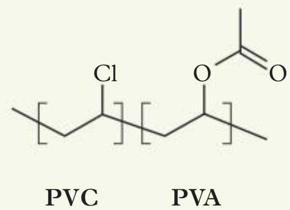


FIGURE 7 Polyvinyl chloride acetate

The properties of PVCA include high transparency, and being odourless and tasteless, very tough, and durable. Because of its excellent heat seal properties, PVCA is commonly used in resins.

a Discuss the properties each monomer contributes to the polymer.

b The combination of monomers varies but can be 80–95% vinyl chloride and 5–20% vinyl acetate. Discuss the overall properties of PVCA when the percentage of monomers changes.

c PVCA can be used in credit cards. Discuss the properties of PVCA that would assist this use.

Study tip

Remember that in polymer reactions, *n* = the number of monomers or repeating units. It is usually >1000, but can be any number depending on the type of polymer.

9.3 Linear and cross-linked addition polymers

KEY IDEAS

In this topic, you will learn that:

+ polymers can be categorised as thermoplastic or thermosetting depending on their properties
+ there are advantages and disadvantages to both thermoplastic and thermosetting polymers.

There are different categories of addition polymers classed by their structures and physical properties. These categories include:

- linear polymers
- cross-linked polymers
- elastomers.

In this topic, we will look at each of these categories of addition polymers and examine how their structure and properties relate to their different applications.

Linear polymers

Linear polymers have no strong chemical bonding between polymer chains. This means only hydrogen bonding, dipole–dipole attractions or dispersion forces attract the linear polymers chains together, as shown in Figure 1.

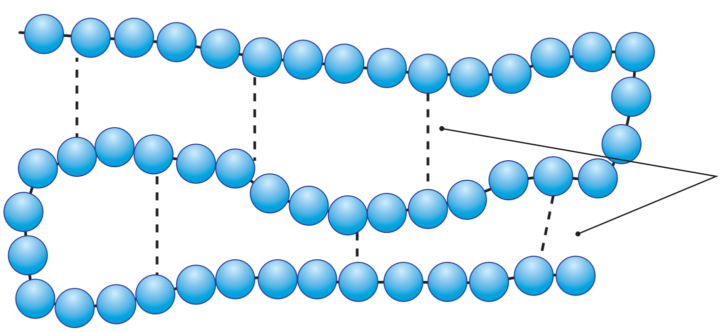


FIGURE 1 Linear polymers, or thermoplastics, have intermolecular forces between the polymer chains.

thermoplastic
a polymer that has weak intermolecular forces between the polymer chains; also known as linear

cure
a process in polymer manufacturing in which material is toughened

Because linear polymers lack strong bonds between their chains, when heated they will soften and eventually become fluid and liquid. The energy to overcome the intermolecular forces between the linear polymer chains is much lower than what is required to overcome covalent or other chemical bonds. When heated, the intermolecular forces are easily broken and linear polymers will soften.

Linear polymers are also known as **thermoplastic** polymers. When thermoplastic polymers are **cured**, they are hard at room temperature. But once they are heated, they become soft. Such properties mean that products made from thermoplastic polymers can easily be remoulded and recycled without affecting the properties of the polymer chains.

9.3
CONTROLLED EXPERIMENT

What effects do cross-links have on the properties of polymers? Go to page **510**.

Applications of thermoplastic polymers

FIGURE 2 The filament for 3D printers is a thermoplastic polymer.

Although the properties of thermoplastic polymers can vary significantly, materials made from thermoplastic polymers generally resist shrinkage and have considerable elasticity and strength. These three properties make thermoplastics desirable in a wide variety of applications including:

- plastic shopping bags
- insulating electrical cabling
- 3D printer filament
- liquid storage tanks
- ropes
- plastic food wrapping.

Thermoplastics are used in the construction, electronics, medical and biomedical, food and beverage, chemical, automotive and plumbing industries. Some of the more common polymers that are produced as thermoplastic polymers are polyvinylchloride (vinyl or PVC), polypropylene (PP), polystyrene (PS), polyethylene (PE), polycarbonate (PC) and polyethylene terephthalate (PET).

Advantages and disadvantages of thermoplastic polymers

Like any material, thermoplastics have their advantages and disadvantages for specific use. These advantages and disadvantages are summarised in Table 1.

TABLE 1 The advantages and disadvantages of thermoplastic polymers

Advantages	Disadvantages
• Resistant to impact, chipping, corrosion and some chemicals and detergents • Adhere well to metal • Can reshape without affecting the properties • Recyclable • Good electrical insulation properties	• Can soften when heated, which makes them undesirable for some applications • Are more expensive than thermosetting polymers

Cross-linked polymers

Unlike linear polymers, cross-linked polymers have covalent bonds between polymer chains. The strong covalent bonds that exist between polymer chains are known as cross-links, as shown in Figure 3. The energy required to break these links is much higher than the energy required to break the weaker intermolecular forces in thermoplastic polymers. Cross-linked polymers are also known as **thermosetting** polymers.

thermosetting
a polymer that has strong covalent bonds between the polymer chains; also known as cross-linked

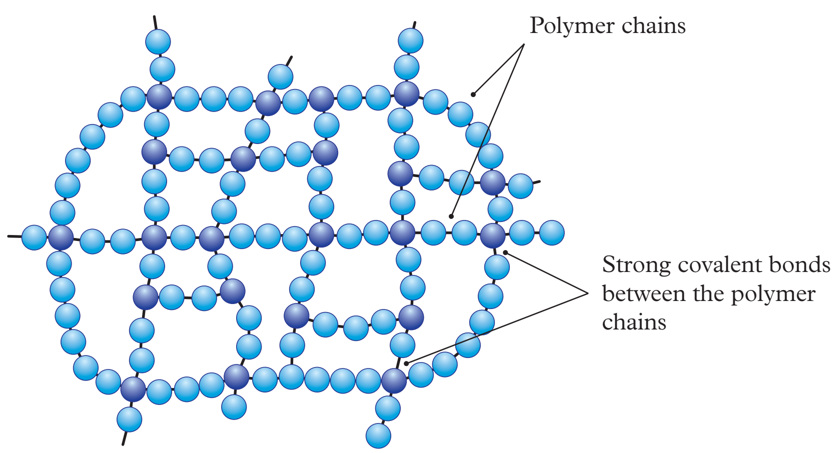


FIGURE 3 Thermosetting polymers have strong covalent bonds between the polymer chains.

Once a thermosetting polymer is cured or set into the desired product, it will not melt when heated and it stays solid because of its strong covalent bonds. This makes thermosetting polymers excellent in applications where heat is a factor; for example, in insulation in electronic appliances. Thermosetting polymers are also much more resistant to chemicals, which makes them excellent for chemical processing equipment and chemical storage containers.

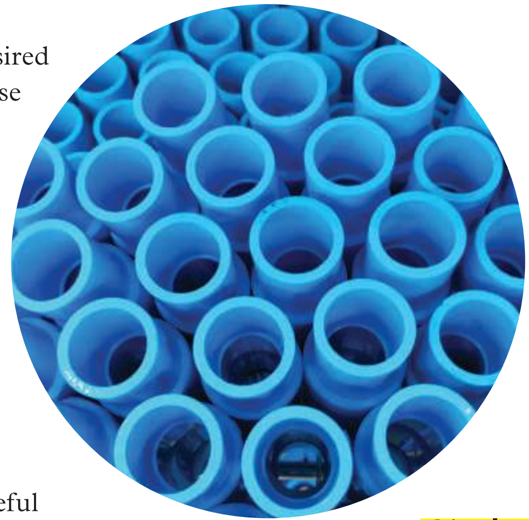


FIGURE 4 Chemical pipes are made from thermosetting polymers.

Applications of thermosetting polymers

The combination of chemical, heat and electrical resistance and structural robustness makes thermosetting polymers useful in many applications, such as:

- electrical and medical equipment, covers and housings
- heavy construction equipment such as door and window frames and panels
- livestock feeding troughs
- chemical pipes and fitting.

Industries that use thermosetting polymers include aerospace and defence, automotive, construction and energy production. Some more common polymers that are used as thermosetting polymers include polyepoxides (epoxy resins), phenol–formaldehyde (PF or phenolics), polysiloxane (silicones), polyethylene terephthalate (polyesters) and polyurethanes.

Study tip

The terms 'thermoplastic' and 'linear' polymer are interchangeable, as are the terms 'thermosetting' and 'cross-linked' polymers.

Advantages and disadvantages of thermosetting polymers

Like any material, thermosetting polymers have their advantages and disadvantages for specific use. The advantages and disadvantages of thermosetting polymers are summarised in Table 2.

TABLE 2 The advantages and disadvantages of thermosetting polymers

Advantages	Disadvantages
• Resistant to chemicals, water, heat, corrosion, deformation and impact • Excellent heat and electrical insulation properties • Can be moulded into different thicknesses and tolerances • Cheaper to produce than thermoplastic polymers • Wide choice of colour and surface finishes	• Cannot be reshaped or remoulded once set • Not recyclable

9.3
Skill drill
Find me in your obook pro

Elastomers

elastomer
a polymer that has elastic or stretching properties

Elastomers are different from both thermoplastic and thermosetting polymers because they have occasional cross-links between the chains, but far fewer than thermosetting polymers. This means that chains can move across and past each other but can also return to their original position.

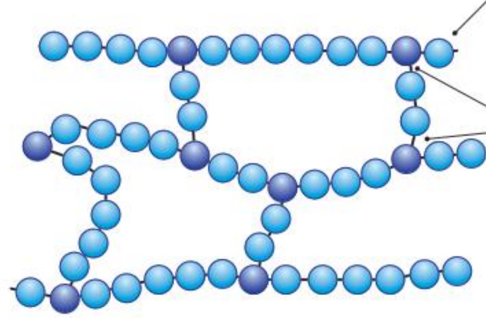


FIGURE 5 An elastomer has some cross-links between polymer chains that allow it to stretch and move.

Rubber bands and elastics and rubber products such as tyres are made from elastomers. The cross-links stop elastomers from melting when heated, but they also make recycling elastomers difficult.

FIGURE 6 Rubber gloves are made from elastomers.

FIGURE 7 The cross-links in rubber bands allow them stretch and then return to their original shape

9.3 CHECK YOUR LEARNING

Describe and explain

1 Describe the bonding between polymer chains in a thermoplastic polymer.

2 Describe the bonding between polymer chains in a thermosetting polymer.

3 Describe the bonding between polymer chains in an elastomer.

Apply, analyse and compare

4 Compare the properties of a thermosetting and a thermoplastic polymer.

Design and discuss

5 Discuss the properties that make thermoplastic polymers useful for 3D printer filaments.

6 Discuss the properties that make thermosetting polymers useful for the housings of electrical equipment.

7 Discuss the properties that make an elastomer useful for the inner tube in a bike tyre.

9.4

Features of linear addition polymers

KEY IDEAS

In this topic, you will learn that:

+ the type of monomer, chain length, branching and crystallinity affect the properties of polymers.

Throughout this chapter, you have learnt that there are many types of polymers, each with specific properties that are considered when they are designed and manufactured. The polymer used to make a pen casing is different from the polymer used to create cable insulation. Polymers that are designed and produced are often done so with a specific purpose and set of properties in mind.

Designing polymers for a purpose

The first step in designing a polymer for a purpose would be to select the properties that you want the polymer to have. Once you have identified these properties, many other areas need to be considered, including:

- monomer selection
- crystallinity
- degree of branching
- polymer chain length.

Monomer selection

There is a wide variety of monomers, each with specific properties. For example, methacrylate is used to make the polymer polymethyl methacrylate (PMMA). Methyl methacrylate produces a polymer that is stiff and transparent and does not allow oxygen to pass through it. This polymer is excellent for Plexiglass or Perspex but does not work well as a packing material because it is tough and rigid.

Polystyrene, which is produced by the addition reaction of styrene monomers, is useful as a packing material. However, it would not work well to make brake light casings for a car because it is not transparent and weather resistant like PMMA.

The selection of monomers is important because it leads to the eventual properties of the polymer. The use of copolymers and designer polymers is becoming more desirable because the properties from different monomers can be added together to make them more versatile.

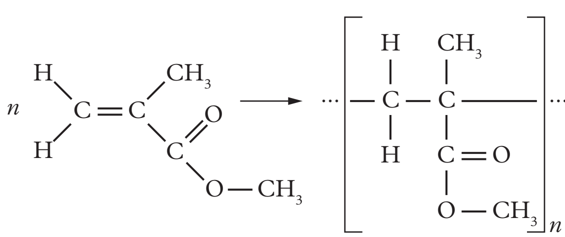


FIGURE 1 PMMA is made from methyl methacrylate monomers.

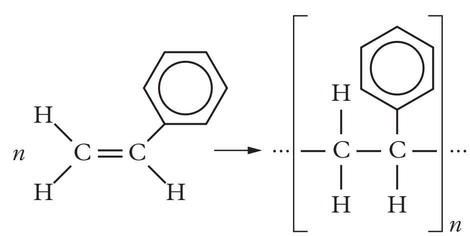


FIGURE 2 The styrene monomer is used to make polystyrene.

9.4 REAL-WORLD CHEMISTRY

Designer polymers

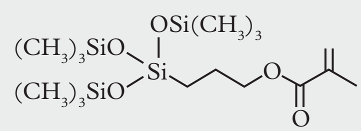
Contact lenses were originally made from polymethyl methacrylate PMMA. Using PMMA meant lenses were transparent and lightweight, so the wearer could see through them. However, PMMA would not allow oxygen to pass through, so it irritated the eyes while being rigid and uncomfortable. A designer polymer was required – a polymer that was lightweight and transparent, allowed oxygen to pass through and was flexible.

The first polymer made to address this issue was the hydrogel polymer. Based on the monomer 2-hydroxyethyl methacrylate HEMA (Figure 3), the polymer is hydrophilic and can absorb several times its mass in water. The result is a contact lens that is soft, comfortable and compatible with the biology of the eye. HEMA-based contact lenses are available today, but they still don't allow enough oxygen to pass through to the cornea, resulting in red eyes, swelling and blurred vision.

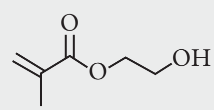
The next design of the polymer was to incorporate silicon-containing polymers, such as 3-tris(trimethylsiloxy)silylpropyl methacrylate. This allowed about six times more oxygen to pass through the contact lenses to the cornea. Unfortunately, adding the hydrophobic silicon monomers to the hydrophilic hydrogels was not straightforward. Silicon also added a stiffness to the lenses and repelled water, causing contact lenses to dry out and not allow the wearer to blink easily.

The solution was to create a designer polymer with the two. This meant that the surface of the lenses needed to have hydrogel channels that trap water to make it moist and comfortable while the core region of the lens contained more silicon channels to transport oxygen to the cornea.

1 Explain the difference between a copolymer and a designer polymer.
2 Discuss the properties of the monomers methyl methacrylate and HEMA that make them beneficial to synthesising contact lenses.
3 Hydrogel polymers are also used in nappies and sanitary pads. Discuss the properties of these that would make them useful for these purposes.



3-Tris(trimethylsiloxy)silypropyl methacrylate



2-Hydroxyethyl methacrylate (HEMA)

FIGURE 3 Structures of the monomers used in contact lens polymers

FIGURE 4 Contact lenses require a flexible polymer that is light weight and transparent and allows oxygen to pass through.

Crystallinity

Crystallinity refers to the way polymer chains are packed. Polymer structure can be:

- **crystalline** – where polymers are regularly organised into lines
- **amorphous** – where polymers are randomly packed.

crystalline
having a set, specifically ordered 3D structure

amorphous
lacking a clearly defined structure or shape

Many polymer materials contain both crystalline and amorphous regions, and the properties of the polymer depends on the percentage of each region.

The more organised and lined up polymer chains are, the more crystalline the structure is. This brings the chains closer together, increases the dispersion forces, and strengthens the material. It also makes the material more opaque because light cannot pass through the densely packed chains.

The amorphous regions of a structure are where chains are tangled. The more randomly packed the polymer structure, the more amorphous regions a structure has. Structures with amorphous regions are less able to pack closely together. These structures have weaker intermolecular forces between the chains and are therefore less rigid, weaker and often transparent.

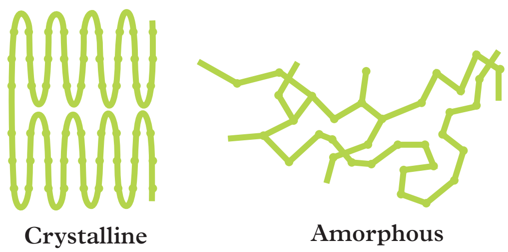


FIGURE 5 Amorphous and crystalline packed polymers

Chain length

As chain length increases, the molecular mass of the polymer increases, which results in the intermolecular forces between the polymer chains increasing. This means polymers with longer chains are stronger and have higher melting and boiling points.

Degree of branching

The degree of branching on the polymer chains also affects the intermolecular forces between them. The more branches a polymer has, the more difficult it is for chains to pack closely together, which weakens intermolecular forces. If there are fewer branches along the chain, the chains can pack in closer and intermolecular forces between chains are stronger.

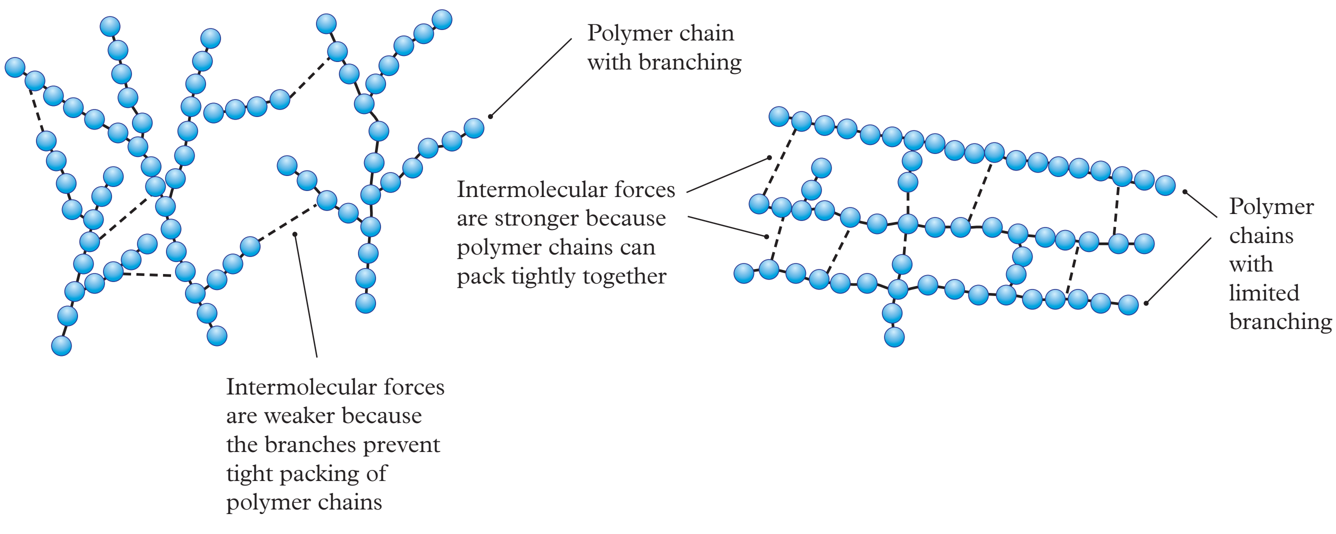


FIGURE 6 Polymer chains with different degrees of branching

A wide variety of polyethylene polymers

Polyethylene is a great example of an addition polymer that has been designed by considering its intended uses. Earlier in the chapter, we looked at the addition reaction to form polyethylene from the monomer ethene. Polyethylene is a thermoplastic polymer that can have many different crystalline structures and a vast range of applications depending on the type. It is one of the most widely produced plastics, with tens of millions of tonnes produced each year.

There are many types of polyethylene. Three that we will look at are:

- low-density polyethylene (LDPE)
- high-density polyethylene (HDPE)
- ultra-high molecular weight polyethylene (UHMW or UHMWPE)

Low-density polyethylene (LDPE)

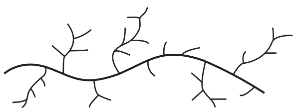


FIGURE 7 LDPE has frequent branching off the main chain.

LDPE is formed under high temperatures and pressures. LDPE forms rapidly, which allows for the formation of branches off the main linear chain, as you can see in Figure 7.

Because LDPE has many branches, chains cannot pack together tightly and dispersion forces between the chains are weak. The properties and uses of LDPE are a result of these weak intermolecular forces.

LDPE is flexible and has excellent ductility, but low tensile strength. These properties are useful for:

- plastic bags
- light packaging materials: six-pack rings, waterproof carton linings, plastic wraps, snap-on lids
- wash bottles
- corrosion protection layer for work surfaces
- computer hardware covers and packaging.

High-density polyethylene (HDPE)

FIGURE 8 HDPE has limited branching off the main chain.

HDPE is a robust, moderately stiff plastic with a highly crystalline structure. Its production takes place at lower temperatures and pressures than that of LDPE and also uses a highly specific transition metal catalyst. Because of the milder conditions that form HDPE, there is limited branching in the HDPE polymer chain, as shown in Figure 8.

Because there is less branching along HDPE chains, polymers produced from HDPE are more rigid because chains can pack closely together. This results in stronger dispersion forces between the chains. Such properties make HDPE application common in:

- strong packaging materials: bottle caps, plastic milk bottles, drums, bulk containers for industrial use
- filaments for 3D printers
- fibres for ropes, nets and industrial fabrics
- vehicle fuel tanks
- boat parts
- pipes and tubing
- plastic chairs and tables
- playground structures: slides, swing seats
- consumer products: garbage and recycling bins, ice cube containers, toys, ice chests.

TABLE 1 Comparison of LDPE and HDPE properties

Property	LDPE	HDPE
Chemical structure	More branching	Less branching, more linear
Flexibility	Less crystalline structure; therefore, more flexible	High crystalline structure; therefore, tougher and more rigid
Chemical resistance	Resistant to most alcohols, acids and bases Low-to-no resistance to oxidising agents and organic solvents	Superior resistance to organic solvents, alcohols, acids and bases
Strength	Weaker than HDPE because of the branching and weaker intermolecular forces	Strong because of the linear structure and stronger intermolecular forces
Heat resistance	Density drops at higher temperatures but retains strength and flexibility over a wide range of temperatures	Useful above 100°C
Transparency	Highly transparent due to low crystalline structure	Low transparency due to increased level of the crystalline structure

Study tip

You don't have to remember all the polymer names or acronyms, but it does come in handy to remember some of the more common ones.

9.4 Challenge
Find me in your obook pro

Ultra-high molecular weight polyethylene (UHMWPE)

UHMWPE is a very dense version of polyethylene. UHMWPE consists of extremely long polymer molecules, which means it has much higher molecular weights than HDPE and LDPE.

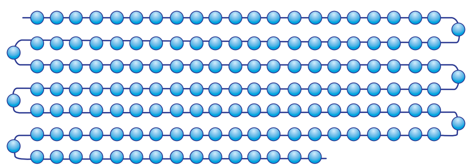


FIGURE 9 UHMWPE has long linear polymer chains.

Longer polymer chains have stronger dispersion forces between them because of increased molecular mass. This makes UHMWPE much stronger than both HDPE and LDPE. UHMWPE can be spun into threads with tensile strengths greater than steel and is frequently incorporated into bulletproof vests and artificial hip joints.

9.4 CHECK YOUR LEARNING

Describe and explain

1 Explain how polymer chain length affects the properties of polymers.
2 Explain how increased branching on polymer chains decreases the strength of a polymer.
3 Describe the properties of a polymer that mainly consists of amorphous regions.

Apply, analyse and compare

4 Compare the structures, intermolecular forces and the properties of LDPE and HDPE.
5 Analyse the section of the copolymer shown in Figure 10 and identify the two monomers used to produced it.

Design and discuss

6 Design a table that summarises the advantages and disadvantages of LDPE and HDPE.
7 Gore-Tex is a designer polymer that is used as a replacement for nylon in jackets. Gore-Tex is made from layers of individual polymers. Its structure is shown in Figure 11.
   a Discuss the properties that make nylon unsuitable as a jacket material.
   b Discuss the properties of each of the polymer layers of Gore-Tex.
   c Discuss what each polymer contributes to the overall success of Gore-Tex as a jacket material.

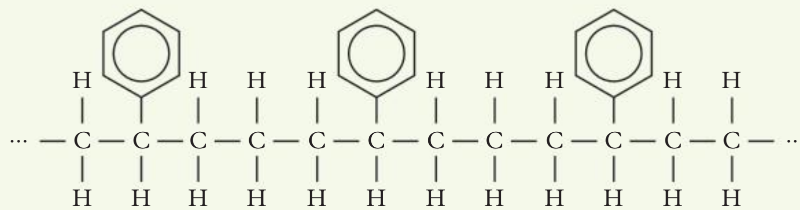


FIGURE 10 A section of a copolymer

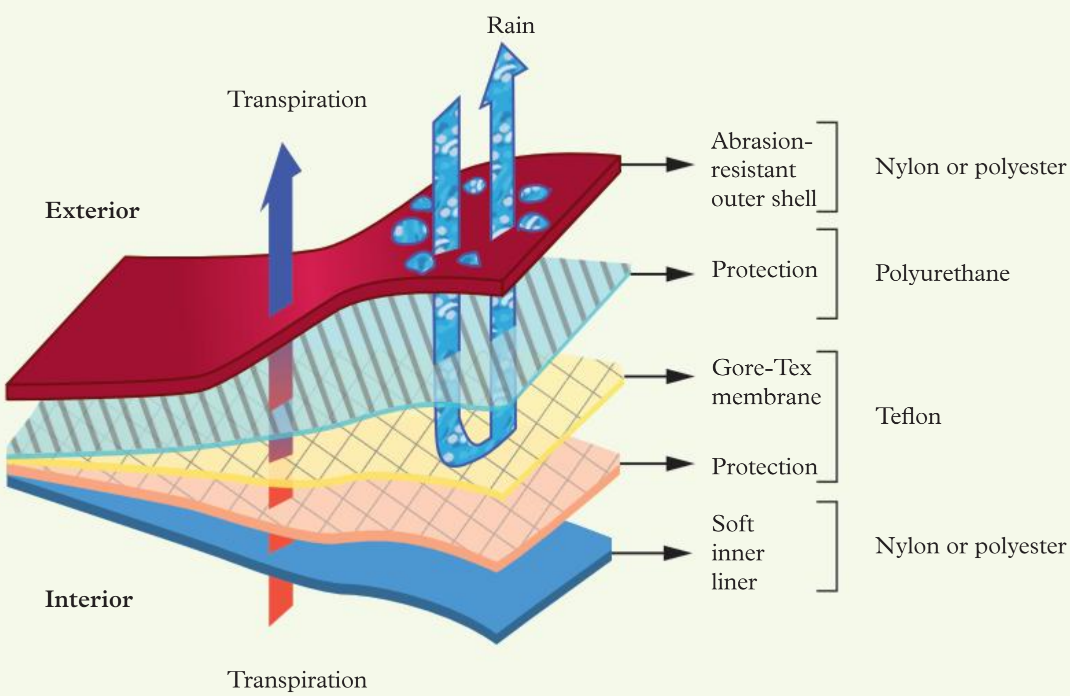


FIGURE 11 The structure of Gore-Tex

9.5 Categorising plastics

KEY IDEAS

In this topic, you will learn that:

+ there are different categories of both fossil-fuel-based plastics and bio-based plastics
+ the different categories of plastics may or may not be recyclable
+ there are different methods of recycling plastics.

Categories of plastics

The source of materials used to make a plastic determines whether they are considered fossil-fuel-based plastics or bioplastics. Plastics that sit within these two categories are further categorised by their structure and recyclability.

Fossil-fuel-based plastics

fossil-fuel-based plastic
a plastic made from petrochemicals

Fossil-fuel-based plastics are polymers made from petrochemicals. These plastics make up a large proportion of the plastics that we use every day. There are seven categories of fossil-fuel-based plastics, each with their own symbol. The seven categories are based on the polymers that form the plastic, their uses and their recyclability. Table 1 summarises the seven categories of fossil-fuel-based plastics and shows some of the next-life uses of the various categories of plastics.

TABLE 1 Categories of fossil-fuel-based plastics

Polymer	Properties	Recyclability	Uses	Next life
Polyethylene terephthalate (PET or polyester) 	Clear tough plastic	Recyclable	• Soft-drink and water bottles • Food packaging • Fruit juice containers • Oil bottles • Spread containers • Mouthwash bottles • Shampoo and conditioner bottles	Used to make more PET products
High-density polyethylene (HDPE) 	Common white and coloured plastic	Recyclable	• Milk, water and juice bottles • Buckets • Yoghurt tubs • Soap dispensers • Detergent bottles • Bleaching agent bottles • Some grocery bags	Garden furniture, pipes and more milk bottles
Polyvinyl chloride (PVC) 	Hard rigid clear plastic	Recyclable at specific places	• Pipe and window fittings • Thermal insulators • Car parts • Shampoo and cleaner bottles	Used to make more PVC products

(continued)

9.5
PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can the design of copolymers be used to meet product requirements? Go to page **512**.

TABLE 1 Continued

Polymer	Properties	Recyclability	Uses	Next life
Low-density polyethylene (LDPE) 	Soft and flexible plastic	Recyclable at specific places	• Frozen food bags • Bread bags • Food bags • Shopping bags • Food wrapping • Squeeze bottles	Bin liners, plastic furniture and floor tiles
Polypropylene (PP) 	Hard but flexible plastic	Recyclable	• Yoghurt and margarine tubs • Sauce bottles • Meat trays • Fibres • Ropes and filaments for carpets • Vehicle upholstery • Luggage • Toys	Clothing fibre, food containers and speed humps
Polystyrene (PS) 	Rigid brittle plastic	Recyclable at specific places	• Takeaway containers • Disposable cutlery • Protective packaging material • Insulation • Meat trays • Disposable cups and plates	As more packaging
Other 	All other plastics, including acrylic and nylon, mixtures of the other six plastics	Not easily recyclable	• Sunglasses frames • Large water cooler bottles • Some sports drink bottles	Goes to landfill

FIGURE 1 HDPE is used to make detergent bottles, LDPE is used to make shopping bags and PP is used to make takeaway containers.

Bioplastics

Bioplastics are polymers made from renewable sources. There are many categories of bioplastics; some are still in the early development stages as we move away from being reliant on fossil-fuel-based products. Many companies are now considering the different raw materials and products that can be produced, used and made from renewable sources such as cellulose, vegetable fats, oils, food waste and more. Table 2 summarises three different bioplastics, the common feedstocks used to produce them, and their current uses, properties and limitations.

bioplastic
a plastic made from organic renewable materials

TABLE 2 Some common bioplastics

Polymer	PLA (polylactic acid)	Bio-PE (bio-polyethylene)	Bio-PP (bio-polypropylene)
Feedstock	• Corn (major source) • Sugar beet • Potatoes • Wheat • Maize • Tapioca	• Sugar cane • Corn and wheat • Other biomass	• Corn • Sugar cane • Vegetable oil • Other biomass, including wood pulp
Replacement for	• Low-density and high-density polyethylene (LDPE and HDPE) • Polystyrene (PS) • Polyethylene terephthalate (PET) • Polypropylene (PP)	• Polyethylene	• Polypropylene
Current uses	Disposable cutlery, plastic coating on paper cups, food containers and food storage	Films (storage bags, pouches, packaging films), drink containers, automotive fuel tanks, injection moulded parts and tubes	Disposable drink containers, coffee cup coatings, coffee lids
Properties and limitations	• High tensile strength • Brittleness and low crystallinity lead to low heat stability and limited applications	• Good oxygen barrier and water vapour barrier properties • Good fat and odour barrier properties that are sufficient for use in food packaging	• Low water vapour barrier • Poor mechanical properties • Bad processability • Brittleness • Still in the development stages

Plastic recycling

Properties of plastics that encourage their use in products include being durable, chemically resistant and lightweight. Such properties also mean that plastics persist in the environment for a very long time once we dispose of them. Although we already have a selection of methods for recycling plastics, scientists and companies are continually coming up with more energy efficient methods of recycling plastics. Three methods of plastic recycling are:

- chemical recycling
- mechanical recycling
- organic recycling.

FIGURE 2 Polylactic acid is used to make disposable cutlery.

When plastics are recycled, and new plastics are formed, we reduce our reliance on using natural resources and instead reuse those we have already sourced. Chemical and mechanical recycling are both examples of a once completely linear economy becoming a more circular economy, where resources can be continually reused for similar purposes.

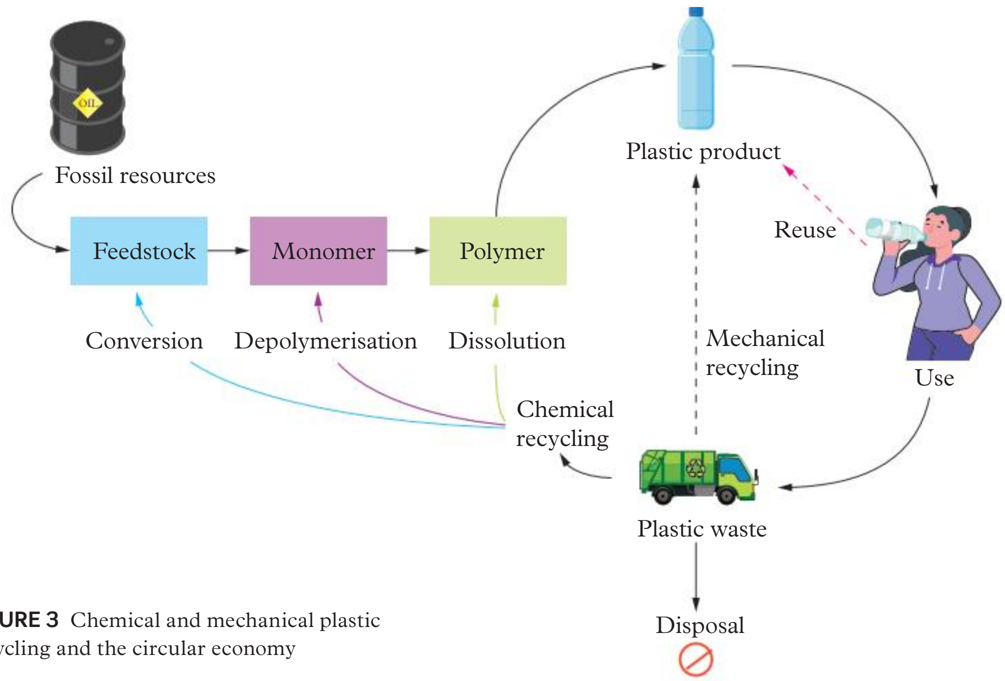


FIGURE 3 Chemical and mechanical plastic recycling and the circular economy

Chemical recycling

Chemical recycling is when polymer chains are split into their original monomers or a new feedstock. Once split, monomers or feedstock can then be transformed into secondary raw materials to produce products that are the same quality, similar quality or lesser quality than the original polymer plastic.

There are three types of chemical recycling:

- conversion recycling
- depolymerisation recycling
- dissolution recycling.

Conversion

Conversion recycling uses heat and chemicals to break down the plastic into liquid or oil-like feedstock, by a process called pyrolysis. Pyrolysis requires large amounts of heat and energy from burning fossil fuels or biomass and produces carbon dioxide. Products from conversion recycling can be used to make some equivalent polymers, but are often downcycled, to make lower quality polymers.

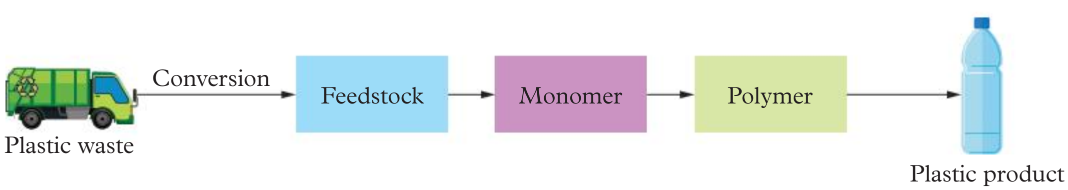


FIGURE 4 Conversion chemical recycling produces new feedstock for polymers.

Depolymerisation

Depolymerisation recycling uses different combinations of chemicals, solvents and heat to break down the polymers into their original monomers. This process is often referred to as chemolysis or solvolysis. Products from depolymerisation recycling are used for similar purposes as those made from raw materials from fossil fuels and used for similar products.

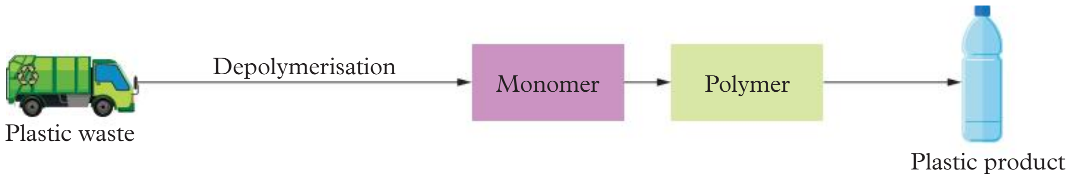


FIGURE 5 Depolymerisation is a type of chemical recycling that restores polymers into their original monomers.

Dissolution

Dissolution recycling uses heat, pressure and a solvent to dissolve the plastic waste into a solution of the polymers, additives and impurities. The wastes and polymers are separated and recovered from the solution. The structure of the original polymer is not altered, so the products made by dissolution recycling are of the same quality as the original products.

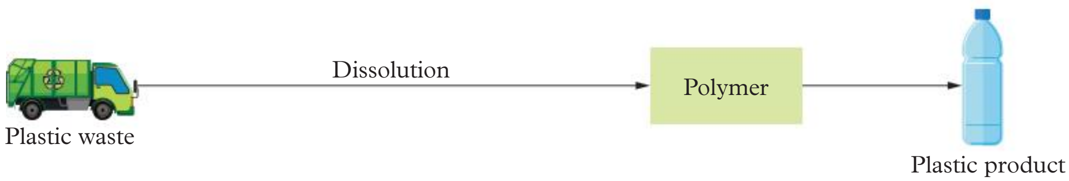


FIGURE 6 Dissolution chemical recycling is the extraction of the original polymer.

Mechanical recycling

Mechanical recycling is a process used to recover plastic waste by mechanical processes that include:

- sorting
- washing
- drying
- grinding
- compounding
- re-granulating or pelletising.

The order and number of times each of these processes occur will differ depending on the type of plastic. All thermoplastics can be mechanically recycled with often limited impact on the quality of the recycled plastics. Plastics that can be mechanically recycled to high-quality raw materials include polyethylene (PE), polypropylene (PP) and polyethylene terephthalate (PET).

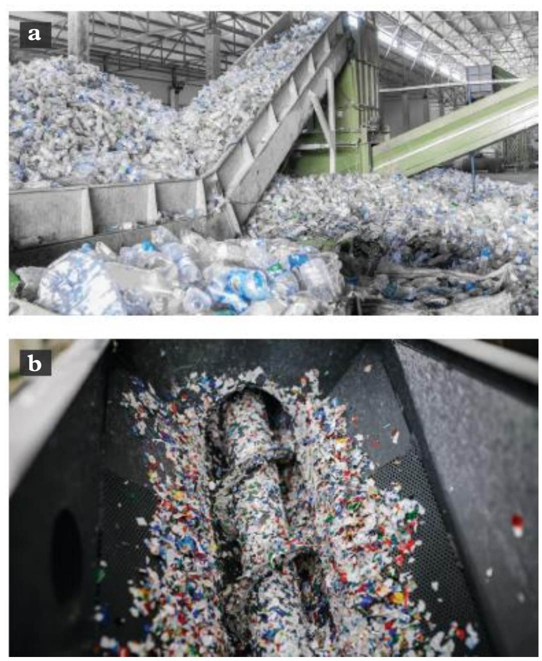


FIGURE 7 Plastic waste **a** ready to be recycled and **b** being mechanically processed

TABLE 3 The challenges and benefits of recycling

Challenges	Benefits
• Plastics need to be sorted into type, which may be automated or done manually. This requires additional time and increased costs. • Plastic waste can degrade with use or over time, which affects the quality of the recycled polymer end product • Plastic waste, with use or over time, can absorb or adsorb persistent organic pollutants (POPs). • Multilayer polymer products are difficult to sort and separate and often cannot be mechanically recycled. • There is a loss in quality after many cycles of recycling results in lower grade polymers.	• Recycling reduces our reliance on fossil fuels. • Less waste is sent to landfill. • Recycled polymers can be used to make high-quality consumer products. • Recycling reduces greenhouse gas emissions • Recycling reduces plastic pollution and related environmental impacts.

Organic recycling

Organic recycling refers to the recycling of bioplastics, rather than plastics from fossil fuels. With the increasing uses and applications of bioplastics, most biopolymers are currently being used for single-use purposes such as food packaging, garbage bags, food wrap and disposable cutlery, plates and cups.

Bioplastics cannot be recycled at standard plastic recycling plants because they are made from different polymers from those in fossil-fuel-based plastic. What happens to bioplastics after their use is largely governed by whether they are biodegradable or **compostable**.

compostable
can break down and safely decompose in compost, and can be beneficially added to soil, leaving no trace

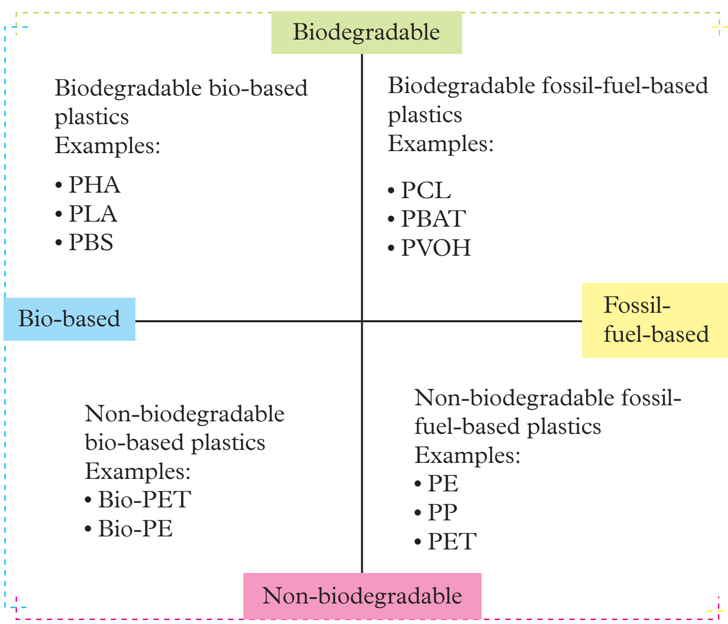


FIGURE 8 Biodegradability of bioplastics and fossil-fuel-based plastics

Biodegradable

Plastics that are biodegradable are decomposed by living organisms such as bacteria or fungi, and under the right conditions will break down and disintegrate back into the earth. The process may take months or years depending on factors such as chemical structure and surrounding conditions. Not all biopolymers or bioplastics are biodegradable. Even some biodegradable plastics cannot completely break down, and can leave behind residues or microplastics that can accumulate in the food chain.

Bio-polyethylene, which is made from a renewable feedstock such as sugar cane, is not biodegradable, just as polyethylene from fossil fuels is not biodegradable. However, some fossil-fuel-based polymers are biodegradable, such as polybutylene adipate terephthalate (PBAT) or polyvinyl alcohol (PVOH).

Composting

Some biopolymers are manufactured to be compostable. A fully compostable substance must be able to break down fully without leaving any traces. Fully compostable substances must also not have any toxic components in their structure to prevent leaving behind any toxic traces after they degrade.

When subjected to the correct conditions, compostable plastics break down quickly and should only leave behind a nutrient-rich organic material called humus. The humus can then be added to gardens and plants and creates a healthy soil environment for plant growth.

In Australia, there are two certifications for compostable plastics – industrial composting and home composting. You can see the labelling for these in Figure 9.

For bioplastics to be completely compostable, the right environments must be sustained. Certain compostable bioplastics require specific moisture and temperature levels and the correct mechanical and microbial actions. These bioplastics are directed to industrial composting facilities. Bioplastics that are home compostable require lower temperatures and often break down quicker.

FIGURE 9 Different composting labels in Australia

9.5 SKILL DRILL

Polymer sustainability

Key science skill: Analyse, evaluate and communicate scientific ideas

Use your knowledge of polymers and the sustainability concepts to analyse each scenario with reference to the sustainability concepts.

1 Evaluate how bioplastics that are compostable address the green chemistry principles and discuss if this is an example of a circular economy.

2 Evaluate how growing crops such as sugar cane to make the feedstock for polymers is a green chemistry principle. Analyse how it affects the sustainable development goals if you use farmland for plastics rather than for growing food.

3 Discuss how recycling fossil-fuel-based plastics by chemical and mechanical recycling is an example of a circular economy and evaluate if it relates to the green chemistry principles.

Need help discussing sustainability? See Topic 1.11 (page 32).

9.5 CHECK YOUR LEARNING

Describe and explain

1 Explain the concept of biodegradable plastics.

2 Describe the process used for mechanical recycling of plastics.

3 Explain why there are different categories for recycling plastics.

Apply, analyse and compare

4 Compare the three types of chemical recycling. Explain which you think is the best type of recycling and justify your answer.

5 Analyse the green chemistry principles and evaluate how they apply to plastic recycling.

6 Compare the products formed from chemical and mechanical recycling of plastics.

Design and discuss

7 Discuss how mechanical and chemical plastic recycling are examples of a circular economy.

8 Discuss if organic recycling is an example of a circular economy. Justify your reasoning.

9.6 Innovations in polymer manufacturing

KEY IDEAS

In this topic, you will learn that:

+ polymer recycling is changing a linear economy into a circular economy
+ some polymers that are produced by condensation reactions can be broken down by hydrolysis reactions
+ there are some exciting innovations in breaking down PET by different types of hydrolysis reactions.

FIGURE 1 Plastic manufacturing from fossil fuels has long been a linear economy.

Before recycling, the polymer manufacturing industry was a linear economy. This is when materials are sourced, used and then disposed of as waste without any means of reusing or recycling materials.

With the great innovations in plastic recycling that have emerged over the last few decades, we are starting to see a shift in the polymer manufacturing industry. What was once a strong linear economy has begun shifting into a circular economy, where after use and disposal, materials are recycled and reused to synthesise new products. In this topic, you will specifically look at some of the innovations around the condensation polymer polyethylene terephthalate (PET).

Condensation reaction to form PET

Polyethylene terephthalate (PET) is formed from a condensation reaction between terephthalic acid and ethylene glycol monomers. The carboxyl and hydroxyl functional groups on the monomers react to form water and the PET polymer, as shown in Figure 2.

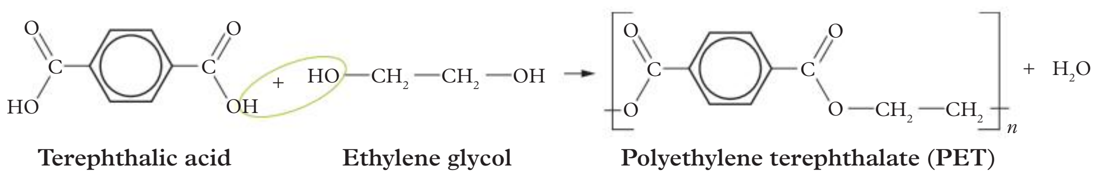


FIGURE 2 The condensation reaction to form the polymer PET

Terephthalic acid has two carboxyl functional group and ethylene glycol has two hydroxyl functional groups. Under the influence of heat and a specific catalyst, the functional groups on each side of the monomers react together, as you can see in Figures 2 and 3.

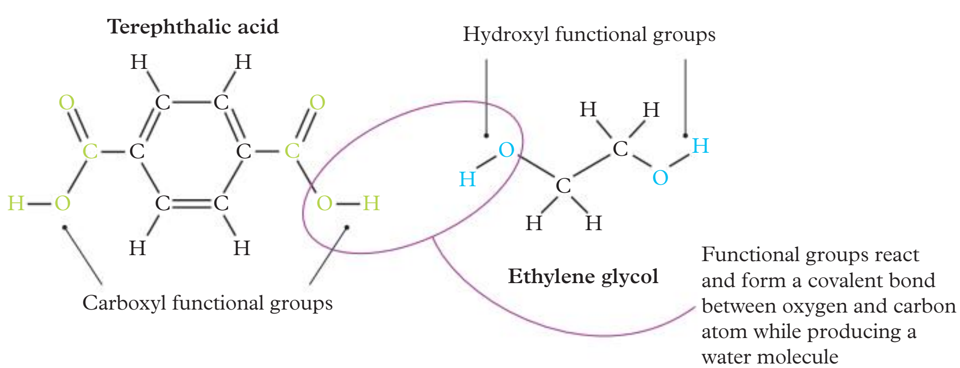


FIGURE 3 The hydroxyl group and carboxyl groups reacting in a condensation reaction

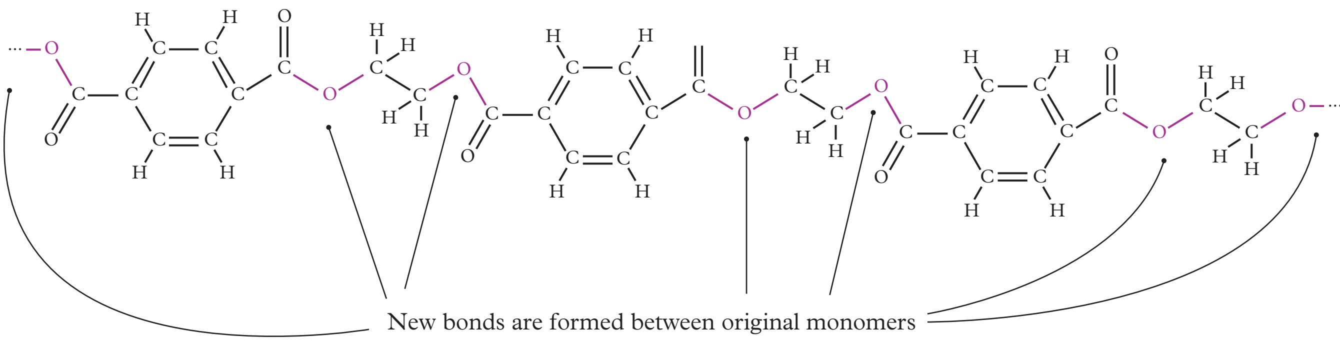


FIGURE 4 PET is a long polymer chain formed from multiple condensation reactions.

PET is a very versatile polymer – it can be drawn and spun into fibres and used for fabrics for clothing, upholstery and carpets or as fibre filling for insulated clothing, furniture or pillows. Some industrial uses of PET are in the automotive industry as a component in tyres, drive belts and seat belts. In other industries it is used for conveyer belts and as reinforcement for hoses. It is also often used in disposable medical garments and in nappies.

But its most common use, which you may have experienced, is in single-use soft-drink bottles. PET's properties of high strength and rigidity, and its impermeability to gas and liquid, make it perfect for plastic drink bottles and other food storage containers.

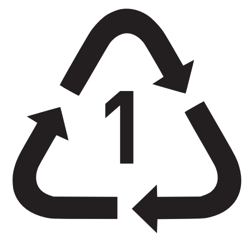


FIGURE 5 PET is used to make plastic bottles and is categorised as a number 1 plastic.

PET recycling

PET is categorised as a number 1 fossil-fuel-based plastic and is completely recyclable. Currently, the main method of PET recycling is a mechanical recycling process, which involves several steps (Table 1).

TABLE 1 The steps involved in PET recycling

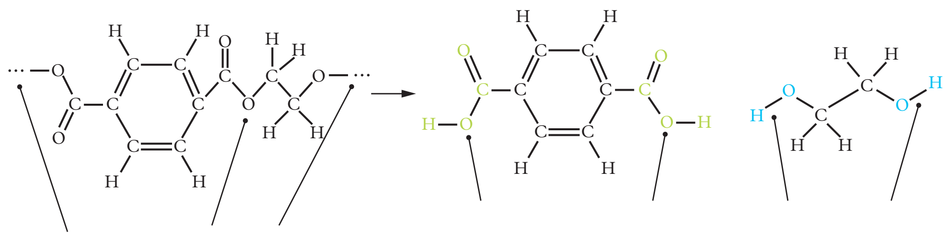
Step	Description
1 Initial sorting	PET is sorted from other plastics and sent to a PET recycling facility.
2 Washing	The bottles are pre-washed, and labels are removed using steam and chemicals.
3 Further sorting	Near-infrared (NIR), metal detectors, flotation and manual sorting equipment is used to identify and remove any non-PET materials.
4 Grinding	PET is ground into particles called 'flakes'. The purity of the flakes is extremely important to the recycled PET integrity.
5 Further washing	The flakes are thoroughly cleaned by detergents and disinfectants.
6 Rinsing and drying	The flakes are rinsed to ensure no contaminants or cleaning agents remain, and then dried.
7 Compounding	The flakes are heated to melt then into a new feedstock, and any foreign materials are also filtered out.
8 Pelletising	Extruded raw material is passed over a series of screens to form pellets of a uniform size that can then be re-introduced into the plastic manufacturing streams.

The end product is solid recycled PET that can be used to form new PET plastic products as part of a circular economy. Unfortunately, not all PET is recycled. Current PET recycling plants cannot always keep up with the large amounts of plastic needing to be recycled. Additionally, sometimes PET is not correctly disposed of, leaving PET plastics to end up in landfill or, worse, oceans and waterways.

hydrolysis reaction
a chemical reaction in which water breaks down a larger molecule into smaller molecules

Innovations in PET recycling using hydrolysis reactions

Because PET plastics are produced by a condensation reaction, they can be broken down using the reverse of this reaction – **hydrolysis reaction**. In this hydrolysis reaction, water reacts with the polymer and breaks the linking covalent bond. The products of the reaction are the initial monomers used to build the polymer. You can see this reaction in Figure 6.



Molecules of H_2O are added and the bond between monomers is broken

After the addition of water molecules, the carboxyl and hydroxyl groups re-form and the monomers are back to their original structure

FIGURE 6 The hydrolysis reaction of the PET polymer, in which water molecules are added to break the bonds and re-form the original monomers

9.6 CHALLENGE

Modelling a circular economy

Using information from this topic, create a diagram that shows how PET recycling would contribute to a circular economy.

Bacterial breakdown of PET

Japanese researchers have discovered a bacterium that can break down and consume PET. The researchers discovered and named the species, *Ideonella sakaiensis*, when they were analysing microbes living on PET debris in wastewater and soil. The bacterium uses two enzymes to break down PET.

Ideonella sakaiensis breaks down PET by hydrolysis reactions, which is similar to how humans break down carbohydrates or proteins naturally. The bacteria complete their hydrolysis reactions using enzymes that help separate the polymer back into its monomers of terephthalic acid and ethylene glycol. They then digest both compounds. Scientists are hoping to either use the bacteria to break down PET plastics or mimic the process using the enzymes.

It takes the bacteria about six weeks to consume a thumbnail-sized piece of plastic. Scientists are looking at transferring the genes that produce the enzymes into a faster growing bacterium, such as *Escherichia coli*, to speed up the reaction. *E. coli* also secretes terephthalic acid instead of consuming it. This makes the process more practical for recycling purposes, as monomers could be collected and put back into the PET manufacturing streams.

Study tip

A hydrolysis reaction (which breaks down a larger molecule into smaller molecules with the addition of water) is the opposite of a condensation reaction (which uses smaller molecules to produce a larger molecule and water).

Catalytic hydrothermal liquification (Cat-HTR)

An Australian company based in New South Wales has devised a more chemical approach to recycling not only PET plastic but mixed plastics. This process is called catalytic hydrothermal liquefaction or Cat-HTR.

In the Cat-HTR reactors, a hydrolysis reaction breaks down the polymers into monomers. The addition of water and heat to the mixed plastics hydrolyses the long polymers chains, adding hydrogen into the monomers where required. The process produces different valuable and useful chemicals and synthetic oil.

The chemicals and oil can both be put back into the plastic manufacturing stream to create new plastics, but the oil can also be used as a fuel. The oil created is combusted to form heat energy to heat the water and reactants in the process. The company can also sell the oil to other companies, or countries as a fuel source.

This process is an example of a circular economy. Despite not being a completely closed loop, one waste stream is being used to create completely new products that can be sent into new manufacturing streams and as energy sources. This process also has the potential to reduce our reliance on fossil fuels.

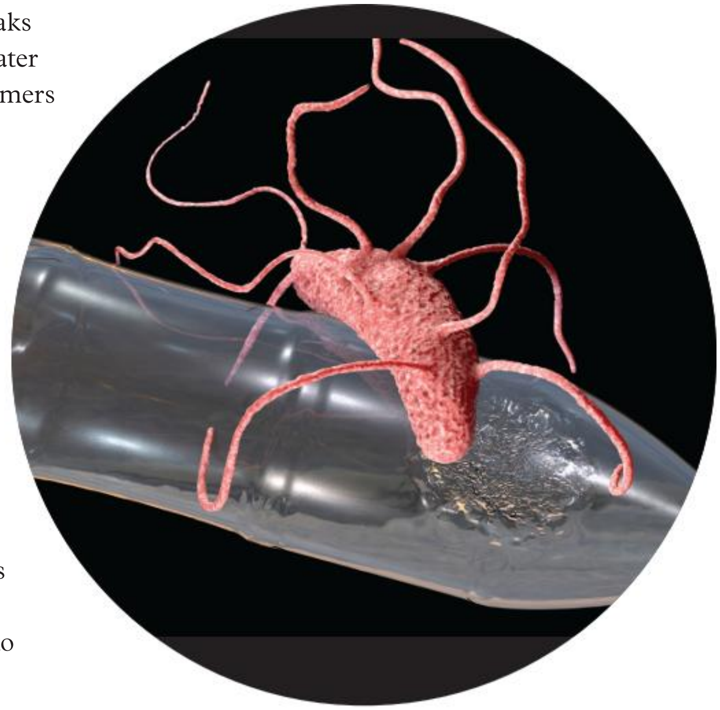


FIGURE 7 A 3D representation of the bacteria *Ideonella sakaiensis* (not an accurate size representation)

9.6 CHECK YOUR LEARNING

Describe and explain

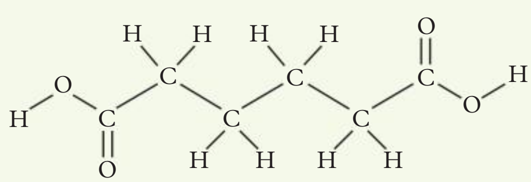
1 Describe a condensation reaction.

2 Explain how a hydrolysis reaction is the opposite of a condensation reaction.

Apply, analyse and compare

3 Compare the advantages and disadvantages of the PET recycling techniques mechanical recycling and Cat-HTR recycling.

4 Ethylene glycol is a very useful monomer in condensation reactions. Analyse the structure and draw the polymer that would be produced by the condensation reaction between it and hexanedioic acid.



Hexanedioic acid

Design and discuss

5 Discuss the properties of polyethylene terephthalate (PET) that make it such a versatile polymer.

6 Discuss how the bacteria *Ideonella sakaiensis* would contribute to a circular economy.

7 Analyse the circular economies from the PET recycling techniques – mechanical recycling and Cat-HTR recycling. Discuss if both these methods address United Nations Sustainable Development Goal 12 ‘Responsible consumption and production’.

CHAPTER 9 Review

Chapter summary

9.1
- Polymers are long chains of repeating monomer units.
- Addition polymers are made from monomers that have a double carbon–carbon bond, which breaks to form new bonds in the polymer.
- Condensation polymers are made from monomers that have functional groups that can react in condensation reactions and form new bonds in the polymer.

9.2
- Addition polymers have different properties depending on the properties of the monomers.
- Addition copolymers can be formed from monomers that each have a double carbon–carbon bond, and they have properties from each monomer.

9.3
- Thermoplastic polymers have no strong bonds between the polymer chains, only weaker intermolecular forces.
- Thermosetting polymers have strong covalent bonds between the polymer chains, known as cross-linking.
- Elastomers have an occasional cross-link between the chains, which allows the chains to stretch and move across each other.

9.4
- Copolymers and designer polymers are made from specifically selected or created monomers that bring specific properties to a polymer.
- Many polymer materials contain both crystalline regions and amorphous regions, and the properties of the polymers depends on the percentage of each region.

9.5
- There are seven different categories of fossil-fuel-based plastics, based on the polymers that form the plastics and their recyclability.
- There are three types of chemical recycling: conversion, depolymerisation and dissolution.
- Mechanical recycling involves many mechanical processes, which affect quality.

9.6
- A condensation reaction is one in which two molecules combine to form a larger molecule and produce a small molecule as a by-product.
- Some polymers that are produced by condensation reactions can be broken down by hydrolysis reactions.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can …	Confidently	Mostly	Not really	Revision link
• describe the differences between addition and condensation reactions as processes for producing natural and manufactured polymers from monomers	☐	☐	☐	Go back to Topic 9.1
• explain how addition polymers are formed by the polymerisation of alkene monomers	☐	☐	☐	Go back to Topic 9.2
• distinguish between linear (thermoplastic) and cross-linked (thermosetting) addition polymers with reference to structure and properties	☐	☐	☐	Go back to Topic 9.3
• describe the features of linear addition polymers designed for a particular purpose, including the selection of a suitable monomer (structure and properties), chain length and degree of branching	☐	☐	☐	Go back to Topic 9.4
• describe how different plastics are categorised as fossil-fuel-based (HDPE, PVC, LDPE, PP, PS) or bioplastics (PLA, Bio-PE, Bio-PP)	☐	☐	☐	Go back to Topic 9.5
• describe the recyclability of polymers (mechanical, chemical, organic recycling), compostability, circularity and renewability of raw ingredients	☐	☐	☐	Go back to Topic 9.5
• describe innovations in polymer manufacture using condensation reactions, and the breakdown of polymers using hydrolysis reactions, contributing to the transition from a linear economy towards a circular economy	☐	☐	☐	Go back to Topic 9.6

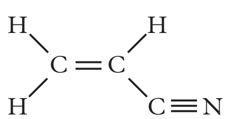
Revision questions

Multiple choice

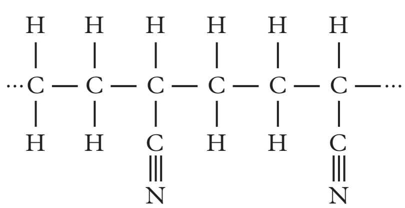
1 Which of the following is a copolymer?

A ABS
B Teflon
C PVC
D Polypropylene

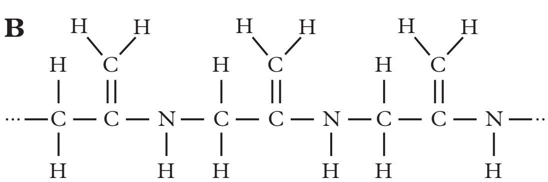
2 Identify the addition polymer that is made from the following monomer.



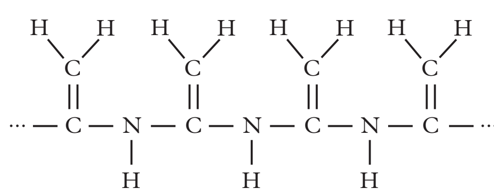
A



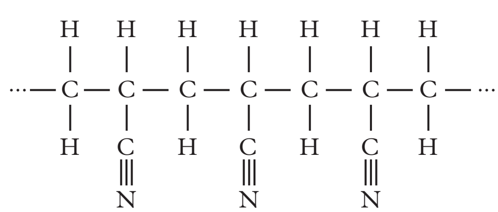
B



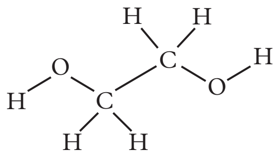
C



D



Use the following structures of monomers to answer Questions 3 and 4.

I


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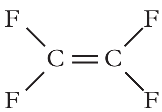
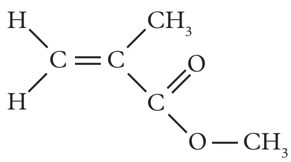
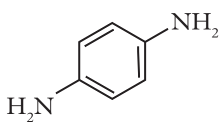
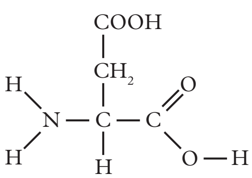


III


IV


V


3 Identify which of the monomers could be used in an addition polymerisation reaction.

A I and V

B II and III

C III and IV

D I, IV and V

4 Identify which of the monomers could be used in a condensation polymerisation reaction.

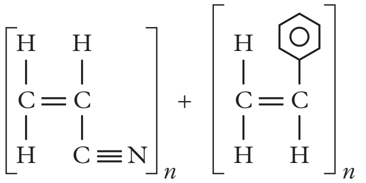
A I and V

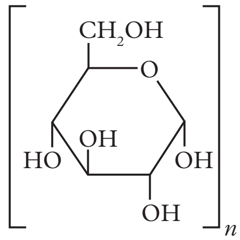
B II and III

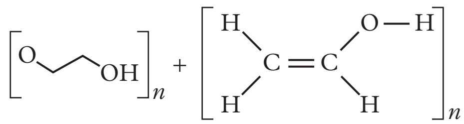
C III and IV

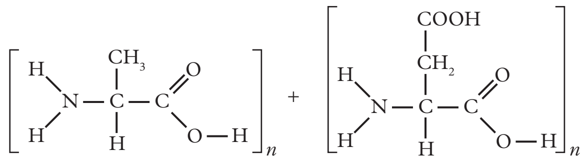
D I, IV and V

Use monomers I–IV to answer Questions 5–7.

I


II


III


IV


5 Identify the monomer(s) that would form addition polymers.

A I only

B III only

C II and IV

D None of the above

6 Identify the monomer(s) that would form condensation polymers.

A II only

B IV only

C II and IV

D III and IV

7 Identify the monomer(s) that would not react to form polymers.

A II only

B III only

C I and III

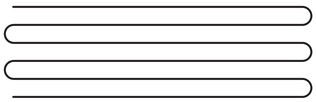
D II and IV

8 For the polymer polyethylene terephthalate (PET), select the correct answer

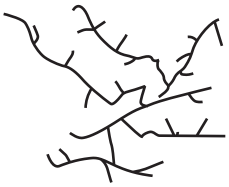
	Reaction type	Monomers
A	Condensation polymer	Homopolymer (one type of monomer)
B	Addition polymer	Homopolymer (one type of monomer)
C	Condensation polymer	Copolymer
D	Addition polymer	Copolymer

Use polymer chains I–IV to answer Questions 9 and 10.

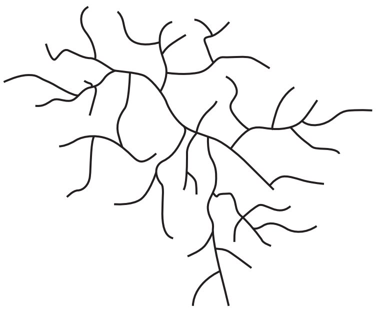
I



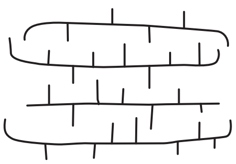
II



III



IV



9 Identify the polymer that will have the lowest melting point.

A I

B II

C III

D IV

10 Identify the polymer that will be the most crystalline.

A I

B II

C III

D IV

Short answer

Describe and explain

11 Draw the structures of the following monomers and explain if they could be used in addition or condensation polymers.

a Ethene

b Ethylene glycol

c Propene

d Phenylethene

12 Explain how a polymer chain with less branching increases the strength of a polymer.

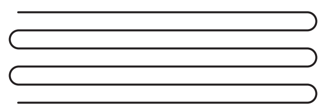
13 Explain how an addition copolymer is formed and what each monomer must have to be able to react in addition polymerisation.

14 Explain how a condensation copolymer is formed and what each monomer must have to be able to react in a condensation polymerisation reaction.

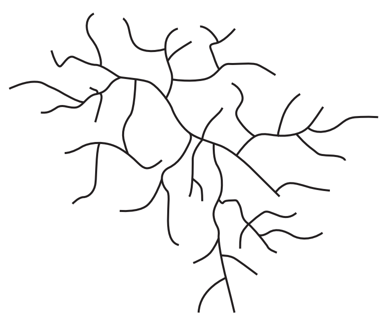
15 Explain what biodegradable plastics are and describe their impacts on the environment.

Use polymer chains I and II to answer Questions 16 and 17.

I



II

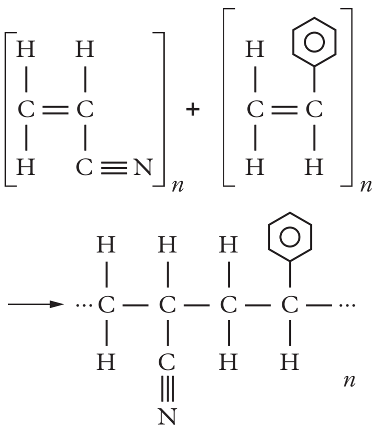


16 Explain how an increase in branching affects the crystallinity of a polymer and describe the crystallinity of polymer chains I and II.

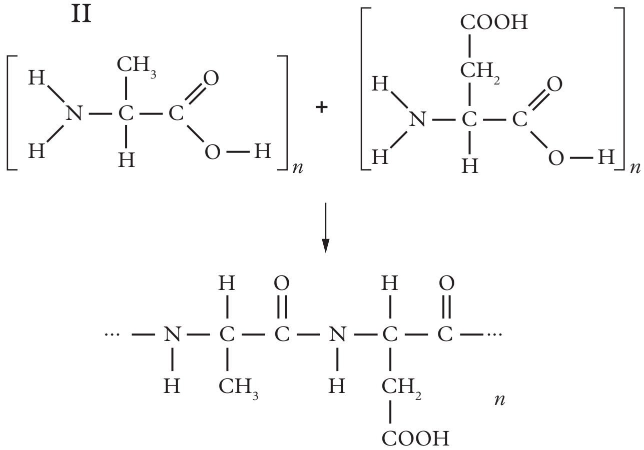
17 Describe the melting points of polymer chains I and II.

18 The following reactions show the formation of two different copolymers from their monomers.

I



II



a Identify whether they are addition or condensation polymerisation reactions.

b Identify each monomer in reaction I.

c Identify the type of natural polymer formed in reaction II.

d Explain why copolymers are manufactured.

19 Explain how using fossil fuels to make the monomers to produce polymers is an example of a linear economy.

20 Explain how recycling plastics is an example of turning a linear economy into a more circular economy.

Apply, analyse and compare

21 Compare an addition polymerisation reaction and a condensation polymerisation reaction. Outline what each reaction requires to be completed and the similarities and differences.

22 Calculate the molecular mass of:

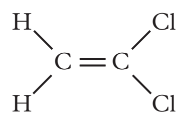
a a polymer that is made of 5000 monomer units of chloroethene

b the natural polymer cellulose that is made of 40 000 monomer units of glucose (M = 180 g mol^{-1})

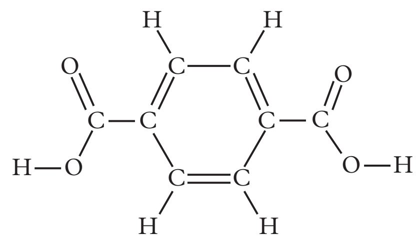
c the copolymer PET that is made of 70 500 monomer units, 55% of which are terephthalic acid (M = 166 g mol^{-1}) and 45% are ethylene glycol (M = 62 g mol^{-1}).

23 Analyse monomers I–V.

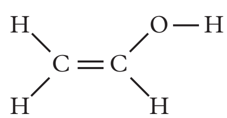
I



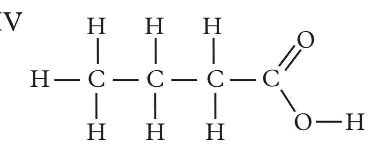
II



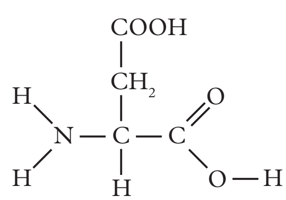
III



IV



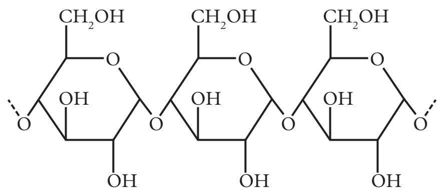
V

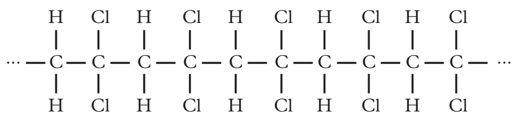


a Identify which monomer could be used in addition or condensation polymerisation reactions or neither.

b Draw a section of the polymers that could be formed from just the addition the monomers.

24 Analyse polymer segments I and II.

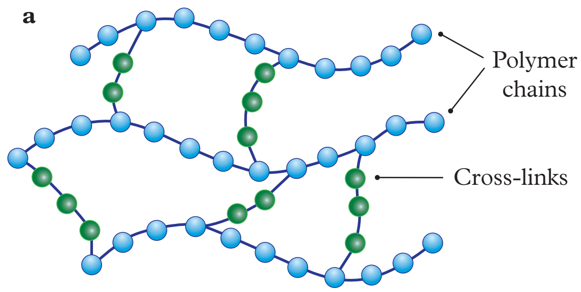
I


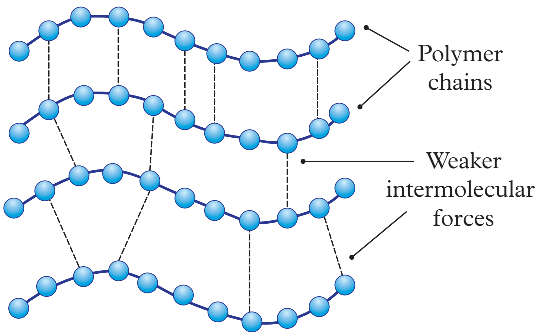
II


a Identify and draw the monomer unit(s).

b Explain whether they are condensation or addition polymers.

25 Analyse the following polymer chains.

a


b


i Identify which is a thermoplastic polymer.

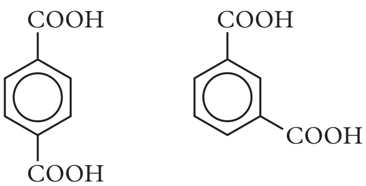
ii Identify which is a thermosetting polymer.

iii Compare the properties of the two, including melting point and recyclability.

iv Discuss the applications of each of the polymers, based on their structures.

Design and discuss

26 Terephthalic acid (shown) is one of the monomers that is usually in PET. When terephthalic acid is replaced with isophthalic acid, it creates a kink in the PET chain. This changes the properties of the polymer. Analyse the structures of terephthalic acid and isophthalic acid.



Terephthalic acid **Isophthalic acid**

a Explain why using isophthalic acid instead of terephthalic acid puts a kink in the chain. Draw the structures of both polymer chains to help explain.

b Discuss whether using the monomers of isophthalic acid instead of terephthalic acid would make the structure more crystalline or more amorphous.

c Discuss the effect on the melting point of using monomers of isophthalic acid compared with using monomers of terephthalic acid.

27 Discuss whether just recycling plastics is enough to help us achieve United Nations Sustainable Development Goal 12 'Responsible consumption and production'. Discuss other ways that we need to adapt our plastic manufacturing to achieve this sustainable development goal.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 1 AREA OF STUDY 2 Checkpoint

Multiple choice

Question 1

A compound of carbon and hydrogen has 14.3% hydrogen by mass and a molar mass of 70. The empirical formula is:

A CH_2
B CH_3
C C_2H_5
D C_4H_{10}

Question 2

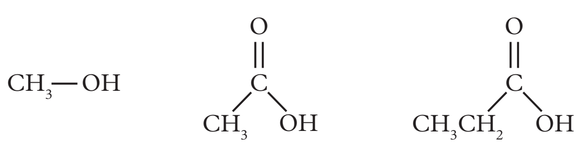
The amount (in moles) in 66 g of $(NH_4)_2SO_4$ is:

A 0.5
B 7.5
C $0.5 \times 6 \times 10^{23}$
D $7.5 \times 6 \times 10^{23}$

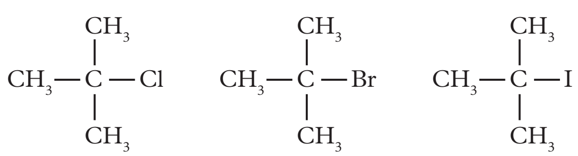
Question 3

Which one of the following groups of compounds is a homologous series?

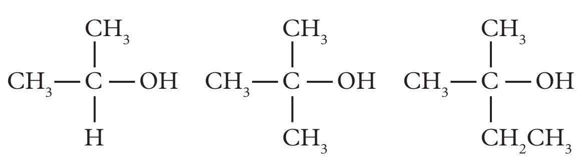
A



B



C



D CH_3CH_2Cl CH_3CHCl_2 CH_3CCl_3

Question 4

Which of the following statements applies to compounds that belong to the same homologous series?

I They contain the same functional group.
II They have the same molecular formula but different structures.
III They have similar physical properties.
IV They have similar chemical properties.

A III only
B I and IV only
C II only
D I, II, III and IV

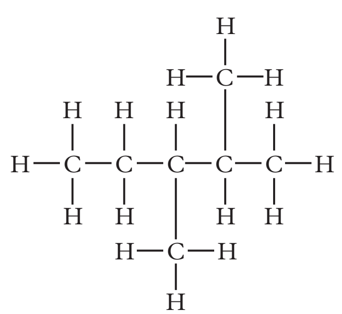
Question 5

The number of possible isomers of chloropropane is:

A 1
B 2
C 3
D 8

Question 6

The systematic name for the molecule shown is:



A 2,3-dimethylpentane.
B 3,4-dimethylpentane.
C heptane.
D 1,2,3-trimethylpentane.

Question 7

Which of the following statements are true?

I Alkenes are a series of related compounds with the general formula C_nH_{2n}.

II Alkenes have the same physical properties as other members of the homologous series, but different chemical properties.

III Alkenes differ by a $-CH_2$ group between consecutive members of the series.

IV Alkenes have the same empirical formula as alkanes and alkynes.

A I and II

B I and III

C I and IV

D I, III and IV

Question 8

Listed below are four metallic compounds with their molar masses (shown in brackets). Which of these compounds contains the metallic element with the highest percentage by mass?

A U_3O_8 (M = 842 g mol^{-1})

B Fe_2O_3 (M = 159.6 g mol^{-1})

C $Cu(NO_3)_2$ (M = 187.5 g mol^{-1})

D $CrBr_3$ (M = 291.7 g mol^{-1})

Question 9

A student calculated the relative atomic mass of boron to be 10.6 by using the following data.

Isotope	Relative abundance	Relative isotopic mass
^{10}B	40%	10.0
^{11}B	60%	11.0

The accepted RAM (B) is 10.8. The incorrect result has occurred because the relative:

A isotopic mass of ^{11}B is actually 11.2.

B abundance of ^{10}B is actually 60%.

C isotopic mass of ^{10}B is actually 10.2.

D abundance of ^{11}B is actually 80%.

Question 10

Ethene undergoes addition reactions with a wide range of reactants. The product of the addition reaction between ethene and hydrogen chloride is:

A $CHClCH_2$

B CH_3CH_2Cl

C CH_2ClCH_2Cl

D CH_2CH_2Cl

Short answer

Question 1 (14 marks)

The table provides formulas of some hydrocarbons.

Compound	Formula
A	CH_4
B	C_2H_2
C	C_2H_4
D	C_2H_6
E	C_3H_4
F	C_3H_6
G	C_3H_8
H	C_4H_6
I	C_4H_8
J	C_4H_{10}

a **i** List the letters that correspond to compounds that represent alkanes.

ii Give the general formula for alkanes.

2 marks

b **i** Write the letter that corresponds to a compound for which isomers exist.

ii Draw the structural formulas and name two isomers corresponding to this letter. Draw a table in your workbook like the one below.

Name:	Name:
Structural formula:	Structural formula:

iii Define 'isomer'.

6 marks

c **i** Write the letter that corresponds to a compound with two carbon atoms and a double bond.

ii Identify the name of the compound you have selected.

iii Draw the structural formula (with correct shape) of the selected compound.

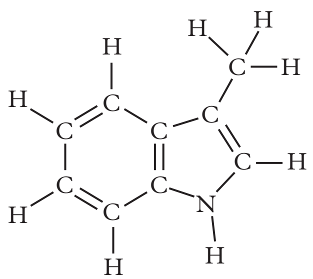
iv A standard industrial test for unsaturation is to add a few drops of red bromine water to the sample under analysis. Construct a balanced chemical equation for the addition reaction between your selected compound and bromine.

v This compound can be polymerised. Construct a balanced chemical equation to illustrate the polymerisation reaction of your selected compound.

vi Name the polymer formed. 6 marks

Question 2 (3 marks)

3-Methylindole, also known as skatole, has a strong, unpleasant odour of faeces. It is also found in cigarettes in very low concentrations. The molecular structure is shown here. The molar mass of skatole is 131 g mol^{-1}.



a Write the empirical formula for skatole. 1 mark

b Calculate the mass (in grams) of one molecule of skatole. 1 mark

c Calculate the amount (in moles) of skatole in 0.015 g of skatole molecules. 1 mark

Question 3 (7 marks)

a Calculate the mass of CO_2 (in kg) produced when 540 g of ethanol burns in air according to the equation:

$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

2 marks

b The table shows the molar masses and boiling points of ethane and ethanol.

Compound	Molar mass (g mol^{-1})	Boiling point (°C)
Ethane	30.07	–89.0
Ethanol	46.07	78.5

Compare the boiling points of ethanol and ethane and explain the difference. 2 marks

c Ethanol made from plant-based biomass is called bioethanol.

i Name the chemical process by which ethanol is produced by plants.

ii Explain how replacing fossil fuels with plant-based biomass is more sustainable.

3 marks

Question 4 (10 marks)

Copy and complete the table by filling in the missing information.

Name	Structural formula	Semi-structural formula
Hexane		
		$CH_3(CH_2)_4COOH$
		$CH_2ClCH_2CH_2Cl$
		$CH_2{=}CH(CH_2)_3CH_3$
2-Methyl-4,5-dichlorohept-2-ene		

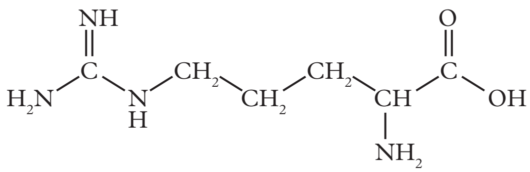
Question 5 (6 marks)

Amino acids are the molecules that make up proteins. A new amino acid has been made. A sample of 123 mg of this new amino acid is analysed and found to contain 19.2 mg of carbon, 4.8 mg of hydrogen, 22.4 mg of nitrogen and 51.1 mg of sulfur. The rest is oxygen.

a Determine the empirical formula. 3 marks

b In a separate experiment, the molar mass is found to be 154.6 g mol^{-1}. Determine the molecular formula. 1 mark

c The structure of the amino acid arginine is shown below. Calculate the percentage by mass of nitrogen in this compound.



2 marks

TOTAL MARKS

/50 marks

CHAPTER 10

Research investigation

KEY KNOWLEDGE

Scientific evidence

→ the distinction between primary and secondary data

→ the nature of evidence and information: distinction between opinion, anecdote and evidence; and scientific and non-scientific ideas

→ the quality of evidence, including validity and authority of data and sources of possible errors or bias

→ methods of organising, analysing and evaluating secondary data

→ the use of a logbook to authenticate collated data

Sustainability

→ sustainability concepts and principles: green chemistry principles, sustainable development, and the transition from a linear economy towards a circular economy

→ identification of sustainability concepts and principles relevant to the selected research question

Scientific communication

→ chemical concepts specific to the investigation: definitions of key terms; and the use of appropriate chemical terminology, conventions and representations

→ characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts; ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence; and appropriateness for purpose and audience

→ the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

→ the influence of social, economic, legal and/or political factors relevant to the selected research question

→ conventions for referencing and acknowledging sources of information

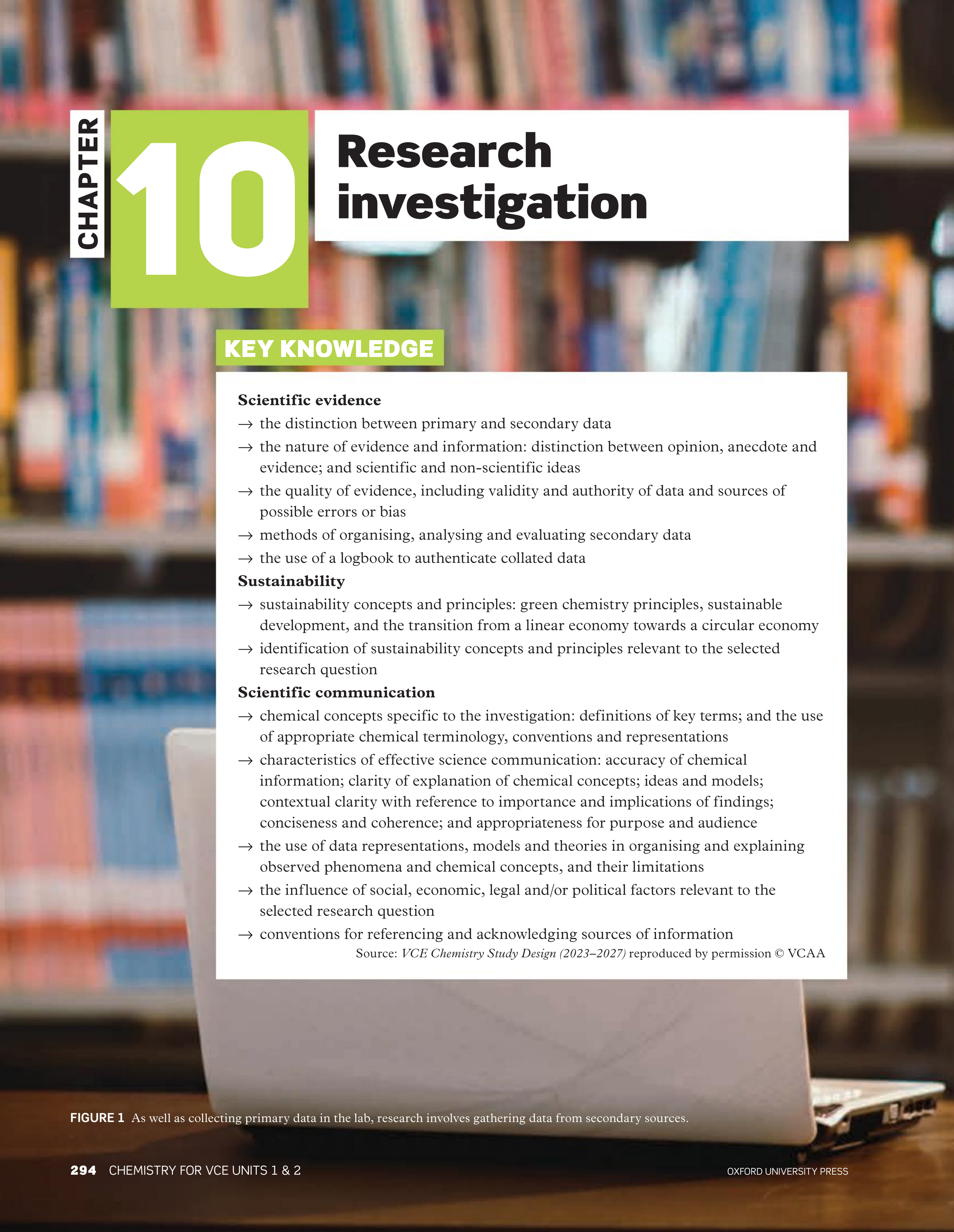


FIGURE 1 As well as collecting primary data in the lab, research involves gathering data from secondary sources.

GROUNDWORK

In Chapter 10, you will conduct your own research investigation.

This chapter will build on concepts you have already learnt in Unit 1 and will allow you to practise skills from Chapter 1. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

10A There are direct links between Chapters 2, 4, 8 and 9 and sustainability. Identify the United Nations Sustainability Development Goals and green chemistry principles that each chapter directly relates to.

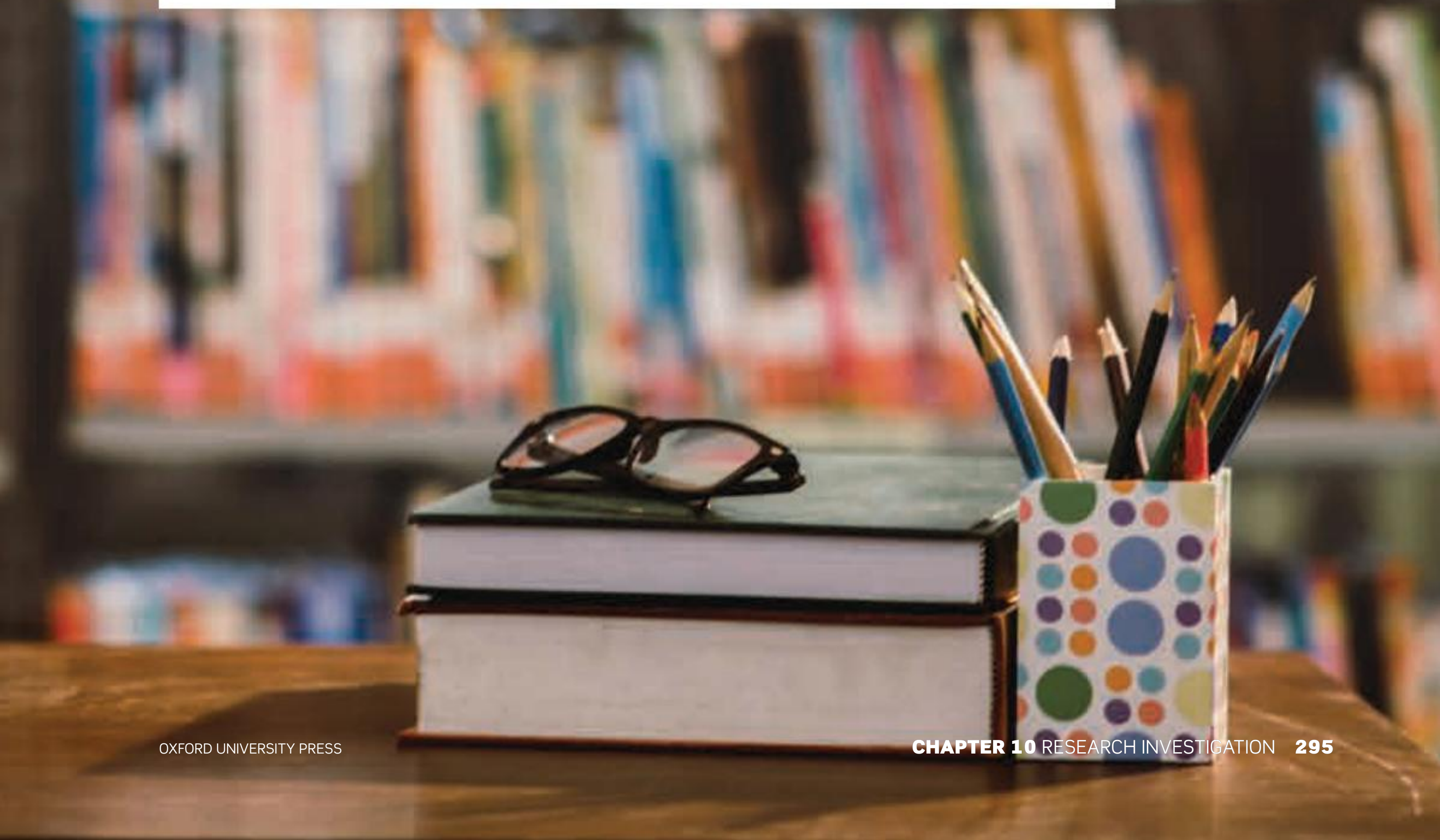
10A Groundwork resource
Sustainability perspectives

10B Comment on whether sustainability concepts explored in Chapters 2, 4, 8 and 9 involve a shift from a linear economy to a circular economy.

10B Groundwork resource
Linear and circular economies

10C This research investigation involves gathering and interpreting secondary sources of data and information. Is this considered a case study, a literature review or fieldwork? Justify your response.

10C Groundwork resource
Secondary sources



10.1 Investigation topics

KEY IDEAS

In this this topic, you will:

+ encounter some ideas of what topic you might like to investigate.

In Unit 1, you will conduct a research investigation. This investigation focuses on applying chemical principles to create a more sustainable future. You will then communicate your research, including explaining the chemical concepts that are relevant to the investigation and identifying the social, economic, legal, political and ethical impacts of the topic in terms of sustainability. To do this, you will need to:

- select and understand your investigation question
- develop a plan or outline
- research information and evaluate your sources
- analyse and interpret your information to form evidence-based arguments
- present your information.

The rest of this chapter will guide you through each stage of your investigation. Don't forget to also check what assessment information your teacher has given you!

The possible topics you might explore include:

- endangered elements in the periodic table
- producing and using 'greener' polymers
- the chemistry of Aboriginal and Torres Strait Islander Peoples' practices
- the sustainability of a commercial product or material.

Research tip

Remember, Chapter 1 Chemistry toolkit also provides information that will help you to conduct your investigation.

Endangered elements in the periodic table

You might like to further investigate endangered elements in the periodic table.

Endangered elements

In Chapter 2, you learnt about endangered elements, such as helium, indium and phosphorus. These elements are at high risk of becoming in very short supply (endangered) if we do not change the ways in which we use them.

FIGURE 1 Phosphorus is used in fertiliser for agriculture.

The depletion of critical elements is most dangerous. These are the elements used in industries that contribute significantly to our economy and standard of living.

This investigation topic asks you to explore the ways in which we use critical elements in order to:

- manage our use of critical elements in a more sustainable way
- replace critical elements with more sustainable alternatives
- avoid the extinction of elements.

Suggested research questions

- Which chemicals are used in the manufacture of fireworks, what is the environmental impact of the combustion of these chemicals to produce the colourful effects seen in fireworks displays, and what alternatives are available?
- Based on their usefulness for society, how would you compare the value of lanthanoids and actinoids with the value of other metal groups in the periodic table?
- Why is helium classified as a critical and endangered element, and how can it be saved given that its atmospheric recovery is almost impossible?
- How is indium mined and used in manufacture of products such as LCD screen televisions and computer monitors, mobile phones or photovoltaic panels, and what alternatives are available if indium becomes scarce?
- How do the properties of the metalloids (such as germanium, antimony, tellurium) differ so much to their neighbours on the periodic table, and how have these properties made them highly important for society and consequentially scarce in supply?
- How are precious metals from electronic waste (e-waste) recycled and what are the environmental and economic benefits of these recovery processes?

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Other research topics

Some other ideas and topics that you might consider exploring for your investigation into endangered elements include the following.

- **Extinction of dangerous elements:** Is it better for humans and the environment if dangerous elements are allowed to become extinct?
- **Phosphorus recovery**: How is phosphorus recovered through the incineration of sewage sludge? Is it better for the planet to look for an alternative to phosphorus?
- **Synthetic elements**: How are elements synthesised? Could we artificially create endangered elements in the laboratory?

Producing and using 'greener' polymers

You can explore the production and use of 'greener' polymers.

'Greener' polymers

As you have seen throughout Chapter 9, polymers are all around us.

Synthetic polymers are made by scientists from petroleum oil, and include, nylon, low-density polyethylene (LDPE), high-density polyethylene (HDPE) and Teflon. There are many challenges that come with using synthetic polymers made from non-renewable resources, such as their sustainability, renewability, disposal and environmental impacts. Biopolymers, also called natural polymers, are naturally occurring in plant and animal cells and include proteins and carbohydrates.

This investigation topic asks you to explore the sustainability of different polymers and how different activities or products could help us to:

- reduce our reliance on synthetic polymers
- replace synthetic polymers with more sustainable alternatives
- reuse or recycle polymers.

Suggested research questions

- What are plant-based biopolymers and what are the impacts of their production on the environment?
- How do biodegradable and degradable polymers, compostable polymers and recyclable polymers differ in structure, production and environmental impacts?
- What is the difference between micropolymers and nanopolymers, and how are used plastic materials and litter managed and repurposed?
- Is the recycling of packaging products containing aluminium more sustainable than LDPE polymer-based packaging products?
- Why is the sale of plastic water bottles and single use plastics banned in many countries?
- How do animal proteins compare with non-animal proteins for different applications, such as meat substitutes and non-animal leather?
- How do the chemical structures of elastomers differ from the structures of thermosetting and thermoplastic polymers, and what are the implications of the production of elastomers for society?
- What impact does the vulcanisation of rubber have on the environment and the communities where rubber is sourced and produced?
- What are the risks and benefits to the environment of the manufacturing, production and application of synthetic fibres for the textile industry (for example, synthetic grass, active wear, shoes and single-use plastics such as takeaway cups, containers, and electrical and electronic products such as mobile phone cords and USB flash drives)?

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FIGURE 2 Biopolymers can be made from organic materials such as bamboo.

Other research topics

Some other ideas and topics that you might consider exploring for your investigation into greener polymers include the following.

- **Plastic straws**: Why are we so against plastic drinking straws in Australia? Did getting rid of them help our plastic waste problem?
- **Recycling plastic**: Does it take more energy to recycle plastic or produce new plastic? Is it better for the planet to just incinerate plastic waste rather than recycle it?
- **Burning waste**: Can we safely burn waste to create fuel? Could this be a solution to our plastic waste (i.e. burning it and producing energy as they do in Denmark)?
- **House insulation**: Why would we use recycled clothes to insulate houses? Does it work?

The chemistry of Aboriginal and Torres Strait Islander Peoples' practices

You may be interested in researching how Aboriginal and Torres Strait Islander Peoples apply knowledge about chemistry to their practices.

Aboriginal and Torres Strait Islander Peoples' knowledges

Over thousands of years, Aboriginal and Torres Strait Islander Peoples have accumulated and refined their traditional knowledge. Chemists around the world can learn from these practices.

This investigation topic asks you to explore how the practices that Aboriginal and Torres Strait Islander Peoples have developed over thousands of years can help us to sustainably modify and process raw materials.

FIGURE 3 The Yandruwandha people know how to prepare nardoo (an aquatic fern) by heating it to destroy toxins.

Suggested research questions

- Which plants are important to Aboriginal and Torres Strait Islander Peoples for their medicinal properties, how are the plants processed before they are used, and what are the active ingredients (for example, the terpineols, cineoles and pinenes as the active constituents of tea trees and eucalyptus resin)?
- What are the chemical processes that occur when Aboriginal and Torres Strait Islander Peoples detoxify poisonous food items: for example, the preparation of nardoo as a food source by heating, and the detoxification of cycad seeds through the removal of cycasins?
- How do Aboriginal and Torres Strait Islander Peoples utilise animal fats, calcination and plant pigments to vary the properties of the paints they make, and how does this compare to Western paint production processes and materials?
- How do binders and fixatives work to allow Aboriginal and Torres Strait Islander Peoples' paintings to be preserved for thousands of years?
- How do Aboriginal and Torres Strait Islander Peoples' glue formulations parallel the use of modern epoxy resins, and how sustainable are the chemical processes involved in producing these materials?
- How are plant-based toxins such as saponins used in Aboriginal and Torres Strait Islander Peoples' fishing practices, and how is this similar to other First Nation Peoples' fishing practices around the world?
- Kakadu plums have long been a component of Aboriginal and Torres Strait Islander Peoples' diets. What active ingredients do they contain that make them a 'super food'?

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Other research topics

Some other ideas and topics that you might consider exploring for your investigation into the chemistry knowledge of Aboriginal and Torres Strait Islander Peoples include the following.

- **Energy content from plant-based materials:** Which resins and gums derived from a range of trees and shrubs have the highest energy content per gram?
- **Production of ash:** Pyrolysis of which plant produces the most alkaline ash?

The sustainability of a commercial product or material

Your investigation might relate to the sustainability of existing commercial products or materials.

Towards a circular economy

To continue meeting the demands of consumers while also managing resources more sustainably, many manufacturers are moving from a **linear economy** to a **circular economy**. By being more responsible with how we produce and consume products and materials, we can reduce the negative impacts on the environment.

This investigation topic asks you to explore how understanding the life cycles and environmental impacts of commercial products and materials could help us to:

- use natural resources in a more productive and sustainable way
- generate less waste
- reduce environmental impacts.

linear economy
a way of managing resources that operates on a take–make–dispose model; new resources are used and disposed of after use

circular economy
a model of production and consumption that involves sharing, leasing, re-using, repairing, refurbishing and recycling existing materials and products for as long as possible

FIGURE 4 Could 'green steel' be the answer to cleaner steelmaking?

Suggested research questions

- What is a 'green steel' and what are the implications of its production for human health and the environment?
- Research a metal mined in Australia: for example, gold, copper or lithium. How is the metal processed and what are its useful properties? To what extent has the metal production and use moved towards a circular economy over the last decade? What innovations have led to the production of the metal being more sustainable over time?
- Select a commercial product that is available in different formulations: for example, vinegar (fermented, synthetic); salt (river salt, sea sold, iodised salt, Himalayan salt); cleaning products (soaps and detergents); oil (fish oil, coconut oil, olive oil); or milk (whole milk, skim milk, low-fat milk, A2 milk, plant milks such as almond, soy and coconut). What ingredients are in the product? How do the ingredients compare in the different product formulations? How is the product made? To what extend does the production of the product involve a linear economy or a circular economy? How does the production and use of the product impact human health and the environment?
- Select a product whose composition has changed over time: for example, hair comb (tortoiseshell to polymer); dental fillings (from silver amalgam and gold to porcelain and composite resin fillings); contact lenses (glass to polymers); paints (lead-based to oil-based and water-based); and tennis racquet strings (from cat gut to nylon and polyester). How have the properties and efficacies of the products changed over time? To what extend have the manufacturing processes become 'greener'?
- Examine the life cycle of a new product or material: for example, unbreakable glass inspired by seashells; new nanomaterials for the treatment of skin infections; and ultra-thin self-healing polymers to make water-resistant coatings. What is the relationship between the properties, structure and the nature and strength of the chemical bonding in the product or material? What are the raw materials used to make the product or material? How is the product or material manufactured? How are the by-products of production treated and managed? Is the product recyclable? Can any wastes during production or at the end of the product's use be repurposed into a useful product or material?

Other research topics

Some other ideas and topics that you might consider exploring for your investigation into the sustainability of commercial products or materials include the following.

- **Biodegradable plastic**: Is this a real thing, or are we now just consuming more plastics in our foods?
- **Recycling tyres**: Why is it so hard to recycle tyres? What are tyres made of that make them so difficult to recycle?
- **Denim recycling**: Why does it take so much energy and resources to produce new denim? Does it take the same amount to produce recycled denim?
- **Vegan leather**: One company in the Netherlands is producing leather from mangoes. How does this compare with leather produced from animals?

10.2 Selecting and understanding your investigation question

KEY IDEAS

In this topic, you will:

+ learn how to write your investigation question
+ gain a better understanding of your investigation question by breaking it down.

Step 1. Select a topic and write an investigation question

This task is a great opportunity to research and understand a key topic that interests you. You can select a question from the suggestions in Topic 10.1 or use them as a starting point to write your own.

Research questions could start with a **guiding word** (Table 1). They should direct you to explore one or more **key terms**. Your question should be specific, clear and complex (i.e. it should not be a yes/no question).

Let's use this investigation question as an example:

What are plant-based biopolymers and what are the impacts of their production on the environment?

Research tip

If you need help writing a research question, you can refer to Topic 1.3 for guidance.

guiding word
a word used to ask a question, e.g. how, what, when, why, where, who

key term
a word that relates to the key topic you are investigating

Step 2. Understand the question

To properly answer your investigation question, you need to make sure that you understand all the key words and identify exactly what the question is asking you to do. First, you might find it helpful to identify the guiding word(s) and key term(s).

In the following example, a number of guiding words and key terms have been identified.

Guiding word — Key term — Guiding word

What are plant-based biopolymers and what are the impacts of their production on the environment?

Key term — Key term

You can then break the question down further into smaller questions. It is also helpful to generate a list of **expanding questions**. These are questions you will need to answer first in order to answer the bigger investigation question. Writing them using more specific **command terms** will help you.

Research tip

The question may change as you start researching. That is OK as long as your investigation is answering your question!

expanding question
a smaller and more specific question that you can derive from the investigation question to help you answer it

command term
a word that provides direction on what you need to do, e.g. identify, explain, analyse, evaluate

In the example below, the question has been broken down into two parts. Expanding questions have been generated for each of them with command terms bolded.

1 What are plant-based biopolymers?
   - **Understand** and **explain** what a polymer is.
   - **Explain** what a plant-based biopolymer is.
   - **Identify** and **describe** specific examples of plant-based biopolymers.
2 What are the impacts of their production on the environment?
   - **Identify** and **describe** environmental impacts of polymers.
   - **Understand** and **explain** how plant-based polymers are produced.
   - **Evaluate** and **assess** the impacts of plant-based polymers on the environment.

This step is important because it will help you to understand what chemical concepts you will need to explain.

TABLE 1 Some helpful guiding words and command terms

Guiding words	Command terms
how, what, when, why, where, who	identify, define, explain, describe, summarise, compare, analyse, evaluate, assess, discuss, justify

TO-DO LIST

- Select a topic and write an investigation question.
- Make sure your question is posed as a question and not a statement.
- Identify the key terms and guiding words in your question.
- Generate a list of expanding questions to guide your research.

FIGURE 1 Which investigation question will you investigate?

10.3 Preparing a research outline

KEY IDEAS

In this topic, you will:

+ learn how to plan your investigation using a research outline
+ identify some ways in which your investigation topic links to sustainability.

Step 3. Prepare a research outline

research outline a tool to help you organise your ideas; can be a skeleton for presenting information later

A **research outline** is a useful planning tool to keep you on track as you do your research. It is used to organise and draw links between ideas. You can use it to work out the order in which you want to write about the information, which provides a handy skeleton for presenting your information later.

You can start with the expanding questions and give yourself more specific tasks to find the information you need. You can continuously add ideas and information to it as you work through the research task. As you can see in the example below, it doesn't have to be neat!

Research tip

Go back to Topic 1.11 if you need some more guidance on linking to sustainability.

– Define a polymer USE STUDENT BOOK
– Define a biopolymer

- **Understand** and **explain** what a polymer is.

  The difference? Link them

- **Explain** what a plant-based biopolymer is. Introduce with a paragraph on biopolymers?
- **Identify** and **describe** specific examples of plant-based biopolymers.

– Get examples USE INTERNET
– Sourced from where? Cellulose? Sugarcane? etc

FIGURE 1 This research task focuses on linking chemistry to sustainability.

Linking to sustainability

Now is a good time to start thinking about how you could link your investigation to sustainability, including the United Nations Sustainable Development Goals, the green chemistry principles and transitioning from a linear economy to a circular economy. An example is shown below.

- **Identify** and **describe** environmental impacts of polymers. Linear economy?
- **Understand** and **explain** how plant-based polymers are produced.

  Link to green chem principle 'feedstock'

- **Evaluate** and **assess** the impacts of plant-based polymers on the environment.

  Link to circular economy?

Once you have outlined where you are headed and linked your ideas, you can start your research.

TO-DO LIST

- Prepare a research outline to help organise and draw links between ideas.
- Identify possible links to the sustainability perspectives.

10.4 Analysing and evaluating your sources

KEY IDEAS

In this topic, you will:

+ find out how to source and evaluate information from different resources.

Research tip

Don't be swayed by bias. Many websites you may have an agenda, a vested interest (where having a particular opinion benefits them) or be funded by a particular organisation. Use the CRAAP method to help you evaluate your sources.

Step 4. Source your information

It is time to start your research! There is an abundance of resources for you to access. The obvious resources are the internet and books, but you could also try scientific magazines, videos, podcasts, or even getting in touch with people working in the field and asking questions. It is important to use a variety of resources so that the information is not biased or limited by one perspective.

When you access a new piece of information, remember to record the details of it in an organised way. You may choose to bookmark all the pages you accessed or use a specific website or program that will record these for you. You could include:

- the title of the source
- who it was written by
- when it was written
- page numbers or URLs
- the date you accessed it.

FIGURE 1 Scientific magazines are a useful source of information.

You'll need this information later to format your reference list. It's a lot easier to record this information as you go than to find it again later.

Step 5. Analyse and evaluate your chosen secondary sources

Not all information out there is reliable or written by experts. It is important to be able to differentiate between **opinion**, **anecdote** and **evidence**. You need to be careful that you only refer to data and research that is credible.

One way to evaluate the reliability of a source is to use the CRAAP method:

Currency: When was the information published? Is it out of date?

Relevance: How does the information fit in with your research question?

Authority: Who published or wrote this information? Are they qualified?

Accuracy: Is the information valid, reproduceable, repeatable and accurate?

Purpose: Why does this information exist? Does the author have an agenda or bias?

If a source does not pass the CRAAP test – for example, if it is out of date, or written by a biased source – do not use it to support your research investigation.

opinion
a person or organisation's judgment or view on a topic; may not necessarily be based on evidence or fact

anecdote
evidence from an individual's personal experience or observations that can be used to support a judgment or view

evidence
information or data collected on a topic that can help form a conclusion

TO-DO LIST

- Source information from different resources.
- Use the CRAAP method to evaluate the reliability of each source.
- Keep track of the details of the sources you have used.

10.5 Organising and interpreting information

KEY IDEAS

In this topic, you will:

+ discover how to organise your information in a way that works for you
+ find out how to link your information and ideas to sustainability.

Research tip

Topic 1.4 provides some useful guidance on how to organise information in a logbook.

Research tip

If multiple sources present the same idea, this could be an indicator that the information is credible. However, it's important to make sure that evidence to support the information has been provided.

Step 6. Organise your information

There are many ways you can organise the information you have collected. One example is using dot points to summarise the information. For each dot point, you could make a note of how the information is relevant to your investigation question.

Drawing links between multiple sources helps you to build a stronger argument in your research investigation. It also acts as a way to critically examine the information and data you have collected. Some important questions to ask yourself are:

- do the sources present the same or similar ideas?
- is there conflicting information?
- has any information been extended or built from previously established information?

There is no one, best way to bring together your information. This depends heavily on your own preferences, and your way of organising ideas may be very different from that of other students in your class. You might like to try:

- dot point summaries of your key findings
- mind maps
- using tables
- Venn diagrams or other graphic organisers.

Step 7. Linking to sustainability

FIGURE 1 Nine of the 17 Sustainable Development Goals are relevant to VCE Chemistry.

You must link the chemical principles relevant to your research investigation to sustainability. Keep in mind the three key perspectives: sustainable development (in reference to the United Nations Sustainable Development Goals), green chemistry principles and moving from a linear economy towards a circular economy. Go back to Topic 10.1 if you need a refresher.

Some example questions to get you thinking are:

- sociocultural – is the practice respectful of different cultures? Is it inclusive?
- economic – what are the costs involved? Is it more expensive or less expensive to use more sustainable practices?
- political – is this a priority area for the government?
- legal – are there any laws preventing the practice from being used?
- ethical – does the practice have a positive or a negative effect on people and/or animals?

TO-DO LIST

- Decide on a way to organise your information.
- Link the chemical principles relevant to your investigation to the sustainability perspectives.

10.6 Communicating your findings

KEY IDEAS

In this topic, you will:

+ learn how to communicate and present your research findings.

Step 8. Present your information

You will have started to put your information together in Step 6. Once you have collected everything you need, you can begin to create a more coherent answer to your investigation question. This means turning the information into a form that can be understood and assessed.

FIGURE 1 One way that you can communicate the answer to your research investigation is through an oral presentation.

Choosing a presentation style

Your teacher may have outlined a specific type of presentation style that they want you to do; otherwise, some options are:

- written report
- poster
- slide show
- video
- oral presentation
- animation.

Putting together your presentation

When you put together your presentation, make sure you:

- include **relevant** information only
  - define all key chemistry terms relevant to the investigation
  - explain relevant chemical concepts
  - use the correct terminology, conventions and representations
- provide **accurate** information
- are **clear** and **concise**
  - use subheadings and dot points where appropriate
  - use diagrams, data representations and models to help explain concepts
- explain the **importance** and **significance** of your ideas
  - link to sustainability
  - highlight any limitations
  - conclude with a final answer to your investigation question
- use the **correct language**
  - remember who your audience is
  - use formal language and avoid writing in the first person
- **practise** and **prepare** (if you are presenting an oral report or video).

FIGURE 2 Data and diagrams are useful pieces of evidence to include in your presentation.

Research tip

Understand what you are writing. If you don't understand a particular word or phrase, look it up. Don't include information you don't understand.

Research tip

You can find more information on scientific communication, including how to write a report, in Topic 1.10.

TO-DO LIST

- Select a style in which to present your information.
- Summarise your findings, making sure to:
  - start with an introduction and end with a conclusion
  - answer your research question.
  - draw links to sustainability.
- Proofread your work.

10.7 Acknowledging your sources

KEY IDEAS

In this topic, you will:

+ develop an understanding of how to acknowledge your sources by referencing them.

Research tip

It isn't enough to just copy, paste and change a few words. This is your research investigation, and you are answering your own question. The material you present should be yours.

in-text reference an acknowledgement of the source immediately after the research or information is referred to

bibliography a full list of all the resources used in some research

Step 9. Reference your sources and compile a bibliography

When communicating the answer to your question, you will need to include references. Acknowledging your sources of data or information is very important. If you do not do this, it can seem like you are trying to pass these ideas off as your own work. This is called plagiarism and is taken very seriously.

There are many ways of referencing sources in your presentation. Your teacher might have a particular style of referencing that they would like you to use. If not, a common style is APA (American Psychological Association) referencing.

In-text references are used each time you write about, specifically reference or quote the work of other people. The reference appears immediately after you have referred to the research or information (e.g. using quotes or paraphrasing).

A final reference list or **bibliography** is added to the end of your report or submitted along with an oral presentation. It is the full list of all the resources you have used in your research.

Table 1 shows a few examples of how to cite references in-text and in a bibliography.

TABLE 1 Formats and examples of in-text and bibliography referencing for different types of sources using APA style

Source type	In-text reference	Bibliography reference
Book	(Author(s) last name, year of publication) Example: 'Biopolymers are becoming more widely used in medicine (Smith & Jones, 2020).'	Author(s) last name, initial. (year of publication). *Title of book* (edition). Publisher. Example: Smith, J. & Jones, B. (2020). *The extensive guide to biopolymers* (1st ed.). Oxford University Press.
Scientific article		Author(s) last name, initial. (year of publication). Title of article. *Journal Title, volume number*(issue number), page–page. doi:xxxx Example: Smith, J. & Jones, B. (2020). The extensive guide to biopolymers. *Journal of Biopolymers, 2*(1), 320–324. doi:10.2020/jbiopolymers.2.1
Website	(Author or organisation, year of publication) Example: 'Biopolymers are becoming more widely used in medicine (Oxford University Press, 2020).'	Author(s) last name, initial. (year of publication). *Title of work*. Retrieved from URL. Example: Oxford University press. (2020). *Biopolymers in medicine*. Retrieved from http://www.oxforduniversitypress.com/biopolymers-in-medicine
Video	(Author(s) last name or screen name, year of publication, timestamp) Example: 'Biopolymers are becoming more widely used in medicine (Smith, 2020, 35:02).'	Author(s) last name or screen name. (year, date of posting). *Title of video* [Format]. Retrieved from 'website address' Example: Smith, J. (2020, 31 March). *Biopolymers in medicine* [Video file]. Retrieved from http://www.youtube.com/watch?v=biopolymersinmedicine

TO-DO LIST

- Include in-text references where required.
- Format your bibliography.

CHAPTER 10 Review

Chapter summary

- As part of Unit 1 Area of Study 3, you will conduct a research investigation related to chemical principles explored in Area of Study 1 and Area of Study 2.
- There are many different topics you could investigate. You must choose an investigation question that involves applying chemical principles to create a more sustainable future.
- To answer your investigation question, you can break down the question into smaller parts and generate expanding questions to guide your research.
- Planning your investigation using a research outline is helpful to keep you on track.
- Not all sources of information are reliable. You can use the CRAAP method to evaluate sources.
- Information can be organised in different ways. You must back link to sustainability.
- You can communicate your research findings in different ways. Include information that is relevant, accurate, significant and important. You must be clear and concise and use the correct language.
- You must acknowledge your sources by in-text referencing and in a bibliography.

Research investigation checklist

Use the following checklist to make sure you have completed the research investigation.

I have …	
understood the question/topic	☐
used my own words	☐
defined and explained any scientific terms	☐
selected my main sources of information	☐
clearly communicated information	☐
linked ideas back to the main question	☐
linked ideas back to sustainability	☐
referred to relevant scientific/chemical theories and models	☐
started with an introduction and finished with a conclusion	☐
analysed all my sources and information using the CRAAP method	☐
included references	☐
proofread my work	☐

UNIT

1 Review

PART A – Revisit and revise

Part A of the Unit Review will help you revisit and revise all of the key concepts and terms from Unit 1 and test your understanding to identify the strengths and weaknesses in your knowledge.

Unit 1 overview

The chart below shows all of the Areas of Study for Unit 1 and the relevant chapters in your student book. Go to the pages shown to review the key concepts for each chapter.

Unit 1

- Area of Study 1
  - **Chapter 2** Elements and the periodic table (page 70)
  - **Chapter 3** Covalent substances (page 112)
  - **Chapter 4** Reactions of metals (page 136)
  - **Chapter 5** Reactions of ionic compounds (page 158)
  - **Chapter 6** Separation and identification of the components of mixtures (page 180)
- Area of Study 2
  - **Chapter 7** Quantifying atoms and compounds (page 208)
  - **Chapter 8** Families of organic compounds (page 244)
  - **Chapter 9** Polymers and society (page 284)
- Area of Study 3
  - **Chapter 10** Research investigation (go to your Student obook pro)

Test your understanding

Use the following table to guide your revision:

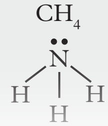
Step 1 – Read the Key knowledge for this Unit.

Step 2 – Test your understanding of the Key knowledge by answering the question(s).

Step 3 – Rate your understanding of each Key knowledge from low to high.

Step 4 – Use the topic and page numbers to revise the concepts you've identified as needing practice.

Only the first few Key knowledge dot points are shown. Access the rest of the Test your understanding questions in your **obook pro**.

Key knowledge	Test yourself	Rate yourself	Focus your revision
• the definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons	**1** Draw and label an aluminium atom, including the number of protons, neutrons and electrons.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.1 Pages 52–55
	2 Consider ${}^{23}_{11}Na^+$, ${}^{24}_{11}Na$ and ${}^{23}_{11}Na$ and identify the: **a** atom **b** ion.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.1 Pages 52–55
• the periodic table (including shell and subshell electron configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements	**3** Write the electron shell and subshell configuration for chlorine. **4** Arrange the elements Ca, F, Si, Li, Br and K in increasing order of: **a** electronegativity **b** atomic radius **c** ionisation energy.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.2 Pages 56–65
• critical elements (e.g. helium, phosphorus, rare-earth elements and post-transition metals and metalloids) and the importance of recycling processes for element recovery	**5** Discuss the importance of moving towards a circular economy with respect to the critical elements.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 2.3 Pages 66–69
• the use of Lewis (electron dot) structures, structural formulas and molecular formulas to model the following molecules: hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene	**6** Analyse the list of molecules: Chlorine Cl Oxygen O::O Carbon dioxide :Ö::C::Ö: Methane CH_4 Ammonia 	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 3.1 Pages 76–82

PART B – Exam essentials

Now that you've completed your revision for Unit 1, it's time to learn and practise some of the skills you'll need to answer exam questions like a pro! To help you, our expert authors have created the following advice and tips maximise your results on the end-of-year examination.

Exam tip 1: Always include states in chemical equations

- States are very important in chemical equations. One example is in a precipitation reaction, to identify the precipitate (s) and the ions that are in solution (aq).
- Consider the hints in the question:
  If it says 'a solution of …' or gives a concentration (e.g. '1 M of …'), then the state is aqueous (aq).
  If it says 'molten', then the state is liquid (l).
  If it says 'a precipitate', then the state is solid (s).
- Remember your intermolecular forces for covalent molecules. Small covalent molecules, such as N_2 and O_2, are gases (g) at room temperature.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 7

Student A decided to precipitate the lead(II) ions in the hair dye as lead(II) iodide. She added an excess of 0.1 M KI solution to 20.0 mL of hair dye.

a. i. Write a balanced equation for the precipitation of lead(II) iodide.

Source: *VCE 2012 Chemistry Exam 1* reproduced by permission © VCAA

A high-scoring response

a. i. Write a balanced equation for the precipitation of lead(II) iodide.

$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$ No spectator ions shown

1 mark

States shown

Is a precipitate

A low-scoring response

a. i. Write a balanced equation for the precipitation of lead(II) iodide.

$Pb^{2+} + 2KI \rightarrow PbI_2 + 2K^+$ No states in equation

1 mark

Spectator ions have not been removed

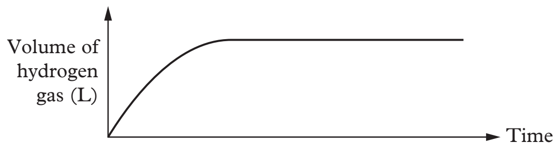
Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guide below to mark the response.

Question 1

A 2.0 g piece of magnesium ribbon was added to a known volume of 2.0 M hydrochloric acid.
The volume of hydrogen gas produced during the reaction was measured and recorded.
The graph below shows the result of this experiment.



a. Write an equation for the reaction between magnesium and hydrochloric acid.

$2Mg(s) + 2HCl(l) \rightarrow 2MgCl(aq) + H_2(g)$

2 marks

Source: *VCE 2008 Chemistry Exam 2* reproduced by permission © VCAA

Marking guide

Question 1a	1 mark for the correctly balanced reaction between HCl and magnesium metal
	1 mark for the correct states of each of the reactants and products

Exam tip 2: Read the question properly

- Look for the task words (e.g. explain, identify, compare, calculate, describe). Make sure you know what they mean. They will tell you how to answer the question.
- Look for any 'tip' in the question, e.g. 'in terms of bonding'.

See it in action

Read the real exam question on the next page and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 7

b. Alkanes with short carbon chains boil at lower temperatures than long chain alkanes. Explain, in terms of bonding, why this is so.

Source: *VCE 2006 Chemistry Exam 1* reproduced by permission © VCAA

A high-scoring response

Highlighted task words and hints

b. Alkanes with short carbon chains boil at lower temperatures than long chain alkanes. Explain, in terms of bonding, why this is so.

✓

- Short chain alkanes have weaker/fewer dispersion forces
- Long chain alkanes have more/stronger dispersion forces
- Fewer/weaker dispersion forces = lower boiling points

1 mark

Clear answer to the question

Key ideas are easier to read as dot points.

Link to bonding in the question

A low-scoring response

b. Alkanes with short carbon chains boil at lower temperatures than long chain alkanes. Explain, in terms of bonding, why this is so.

✗

Alkanes have covalent bonding, so smaller chains will have fewer covalent bonds, meaning they will have lower boiling temperatures

1 mark

Made the connection to the question
Smaller = lower boiling point

Link to bonding but the wrong type of bonding

Boiling point is all about intermolecular forces, like dispersion forces

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guidance below to mark the response.

Question 1

b. Methane can be obtained from natural gas deposits **or** as a biochemical fuel from biomass.

i. Which of these sources of methane would be considered more sustainable? Explain your answer.

Methane from biomass is more sustainable because it isn't a fossil fuel.

2 marks

Source: *VCE 2010 Exam 2* reproduced by permission © VCAA

Marking guide

Question 1a	1 mark for correctly selecting the more sustainable source of methane
	1 mark for explaining why that fuel source is more sustainable

Exam tip 3: Use correct chemical language

When naming hydrocarbons, it is important to recall and use the IUPAC rules with the correct terminology:

- Use hyphens to separate numbers from words when naming organic compounds; for example, '3-bromohexane' not '3 bromohexane'.
- Use commas to separate numbers when naming organic compounds; for example, '1,3-dibromohexane' not '13-dibromohexane'.

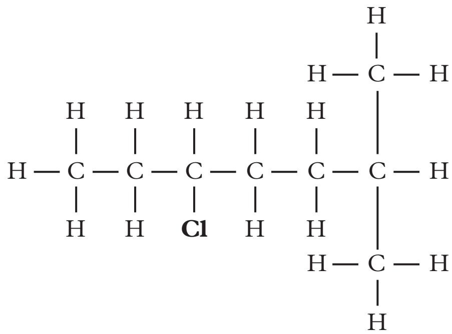
See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 6

a. Give a systematic name for each of the following compounds.

ii



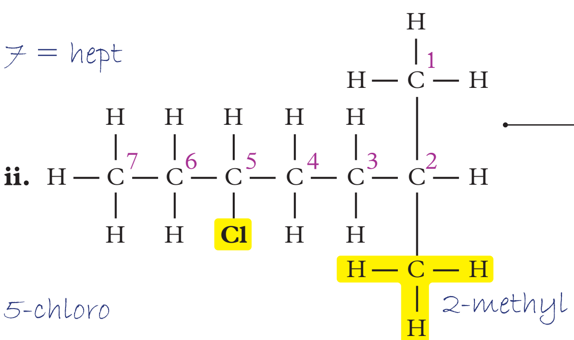
2 + 2 = 4 marks

Source: *VCE 2008 Chemistry Exam 1* reproduced by permission © VCAA

A high-scoring response

a. Give a systematic name for each of the following compounds.

ii.



Identified functional groups and numbered carbons from the end closest to a functional group

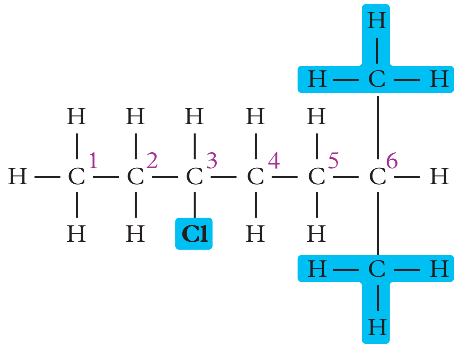
5-chloro-2-methylheptane

2 + 2 = 4 marks

Hyphens between numbers and words ONLY

A low-scoring response

a. Give a systematic name for each of the following compounds.

ii. 

3-chloro 6,6-di-methylhexane

Missed hyphen here

There should be no hyphen here

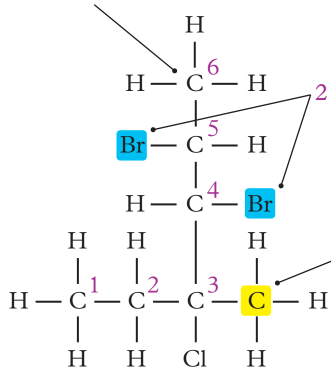
2 + 2 = 4 marks

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guide below to mark the response.

Question 1



a. iii. Give the IUPAC name of the compound that has the structural formula shown above. 1 mark

3-methyl, 4-5-bromo hexane

Source: *VCE 2018 Chemistry Exam 1* reproduced by permission © VCAA

Marking guide

Question 1a	1 mark for the correct systematic name of the structure, remembering the naming rules and hyphens between names and numbers and commas between numbers

Practice makes perfect

Now that you know all these tips, it's time for you to move on to Part C – Exam practice to put them into practice.

PART C – Exam practice

Multiple choice

Question 1

Which of the following lists of elements are in order of increasing electronegativity?

A Potassium, magnesium, boron, oxygen
B Calcium, beryllium, sulfur, fluorine
C Chlorine, phosphorus, calcium, sodium
D Fluorine, silicon, magnesium, potassium

Question 2

Identify the atom with the greatest number of neutrons.

A $^{74}_{32}Ge$
B $^{75}_{33}As$
C $^{76}_{34}Se$
D $^{81}_{35}Br$

Question 3

Which of the following molecular shapes is correctly named?

	V-shaped	Linear
A	CO_2	H_2O
B	CN	CO_2
C	H_2O	H_2S
D	H_2S	CN

Question 4

Which property of metals can *not* be explained by the metallic bonding model?

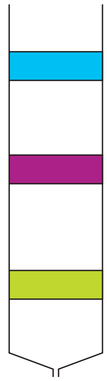
A Magnetism
B Conductivity
C Ductility
D Lustre

Question 5

When solutions of silver nitrate and sodium carbonate are mixed, a white precipitate is formed. Identify the formula of the white precipitate.

A $AgCO_3$
B Ag_2CO_3
C Na_2NO_3
D $NaNO_3$

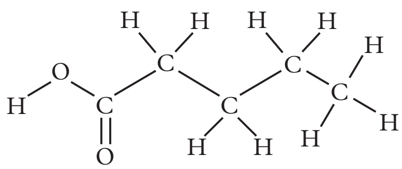
Question 6



Column chromatography was undertaken to identify the pigments in a plant sample. What conclusions could be made about the yellow pigment?

A It has a high R_t and a high affinity for the stationary phase.
B It has a low R_t and a high affinity for the stationary phase.
C It has a high R_t and a high affinity for the mobile phase.
D It has a low R_t and a high affinity for the mobile phase.

Questions 7 and 8 refer to the structure shown.



Question 7

Calculate the number of moles of atoms present in 8.30 g of the compound.

A 0.0814
B 1.38
C 102
D 4.9×10^{22}

Question 8

How many structural isomers does the compound have that are carboxylic acids?

A 1
B 2
C 3
D 4

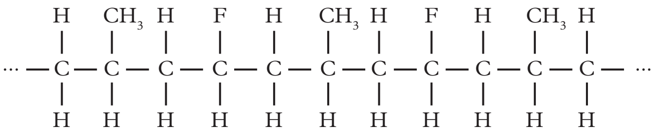
Question 9

Which of the following hydrocarbons has a higher boiling point than butane?

A Ethane
B Ethanol
C Methylpropane
D Butene

Question 10

Copolymers form when two or more different monomers react. Part of a copolymer is shown.



The two alkenes that react together to form this polymer are:

A ethene and fluoroethene.

B propene and fluoroethene.

C ethene and 1-fluoroprop-2-ene.

D ethene and 1-fluoroprop-1-ene.

Short answer

Question 1 (7 marks)

Period 4 of the periodic table contains metals, starting from potassium to non-metals, such as bromine and krypton.

a Write the subshell electron configuration for the group 2 element in this period. (1 mark)

b Outline the following trends in this group as you go from left to right.

i Electronegativity (1 mark)

ii Atomic radius (1 mark)

c Some of the properties of the metals in this group are listed in the table.

Metal	Melting point (°C)	Boiling point (°C)	Reactivity with oxygen	Reactivity with dilute HCl
Potassium	63.5	758.8	Rapid reaction	Violent reaction
Chromium	1907	2672	Very slow reaction	Reaction
Iron	1538	2862	Very slow reaction	Slow reaction
Zinc	420	907	Slow reaction	X

i Describe a physical property of metals that isn't shown in the table above. (1 mark)

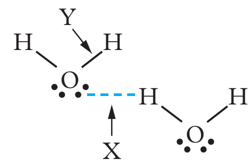
ii Using the information in the table, identify one difference in properties between main group metals and transition metals. (1 mark)

iii Write the balanced chemical equation for the reaction between HCl and zinc metal. (1 mark)

iv Describe how you could determine whether a reaction between zinc and HCl will take place. (1 mark)

Question 2 (7 marks)

Water is a covalent molecule and can interact with other water molecules.



a Describe the molecular shape of water molecules. (1 mark)

b Identify the types of bonding that are indicated by labels X and Y. (2 marks)

c Discuss why pure deionised water will not conduct electricity, but tap water does. (4 marks)

Question 3 (8 marks)

Ionic substances have high melting points. For example, sodium chloride (NaCl) has a melting point of 801°C and magnesium sulfate ($MgSO_4$) has a melting point of 1124°C.

a Using the ionic bonding model, explain why ionic compounds have high melting points. (1 mark)

b Describe how the state of an ionic substance can affect its ability to conduct electricity. (2 marks)

c Two students completed a practical to determine if precipitation reactions would occur between different aqueous solutions. The reactions were carried out and their results are shown in the table below.

Reaction	Reactant 1	Reactant 2	Precipitate formed
A	Lead nitrate	Sodium carbonate	Yes
B	Sodium carbonate	Copper(II) sulfate	Yes
C	Magnesium chloride	Sodium hydroxide	X
D	Sodium nitrate	Potassium iodide	X

i Write a fully balanced chemical equation for precipitation reaction A. (2 marks)

ii Identify and give the molecular formulas of the spectator ions in reaction B. (1 mark) iii Determine whether there were precipitates formed in reactions C and D. (2 marks)

Question 4 (10 marks)

Consider the alcohol butan-1-ol.

a Write the semi-structural formula of butan-1-ol. (1 mark)

b Draw the structural formulas of two isomers of butan-1-ol. (2 marks)

c An alcohol in the same family as butan-1-ol has the following composition by mass: 52% carbon, 13% hydrogen and 35% oxygen.

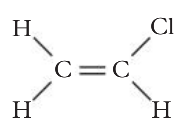
i Determine the empirical formula of the alcohol. (2 marks)

ii For this alcohol, the empirical formula is the same as the molecular formula. Use this information to calculate the number of carbon atoms in a 32.0 g sample of the alcohol. (3 marks)

d Compare the boiling point of butan-1-ol to those of the other members of its homologous series, with reference to bonding. (2 marks)

Question 5 (8 marks)

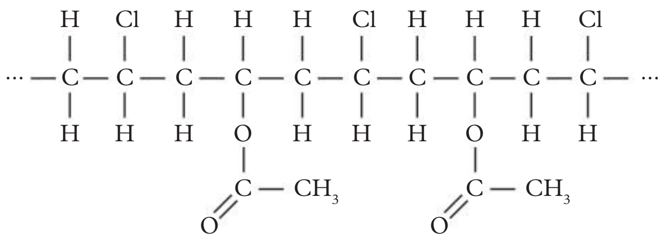
Polyvinyl chloride is a polymer made from the monomer vinyl chloride (chloroethene) (shown here).



a Draw a section of PVC that contains at least five repeating units. (1 mark)

b PVC is a thermoplastic polymer. Explain how this is different from a thermosetting polymer, with reference to properties, bonding and structure. (3 marks)

c PVC can be used in copolymers, such as polyvinyl chloride acetate (PVCA), which is used in credit cards and vinyl records. PVCA has the structure:



i Draw the monomer that reacts with PVC to form the copolymer PVCA. (1 mark)

ii Determine the molecular formula of the monomer you drew for part **c i**. (1 mark)

d Another alkene monomer, ethene, can be derived from biomass or from fossil fuels. Both form the polymer polyethylene which can be recycled. Explain how the recycling of ethene from biomass is more of a circular economy compared to recycling ethene from fossil-fuel-based sources. (2 marks)

TOTAL MARKS

________ /50 marks

You can find the following resources for this section in your obook pro:

Unit 1 Review Part A
Test your understanding questions

Unit 1 Review Part C
Exam practice questions

Weblink
Past examinations and examiners' reports

UNIT 2

HOW DO CHEMICAL REACTIONS SHAPE THE NATURAL WORLD?

FIGURE 1 Chemists use a range of equipment to test and analyse substances.

Unit 2 Overview

In Unit 2 of VCE Chemistry, you will learn how chemists analyse the materials and products we use every day, and explore the properties and applications of water, acid–base and redox reactions.

Unit 2 Areas of Study

The learning for this unit has been divided into three Areas of Study. The table below shows how each Area of Study aligns with the chapters in this book and lists the page numbers for each chapter.

Unit 2 Outcomes

In this unit you will:

→ explain the properties of water in terms of structure and bonding, and experimentally investigate and analyse applications of acid–base and redox reactions in society

→ calculate solution concentrations and predict solubilities, use volumetric analysis and instrumental techniques to analyse for acids, bases and salts, and apply stoichiometry to calculate chemical quantities

→ draw an evidence-based conclusion from primary data generated from a student-adapted or student-designed scientific investigation related to the production of gases, acid–base or redox reactions or the analysis of substances in water.

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

CHAPTER 11

Water as a unique chemical

KEY KNOWLEDGE

→ existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water

→ explanation of the anomalous properties of H_2O (ice and water), with reference to hydrogen bonding:
- trends in the boiling points of Group 16 hydrides
- the density of solid ice compared with liquid water at low temperatures
- specific heat capacity of water including units and symbols

→ the relatively high latent heat of vaporisation of water and its impact on the regulation of the temperature of the oceans and aquatic life

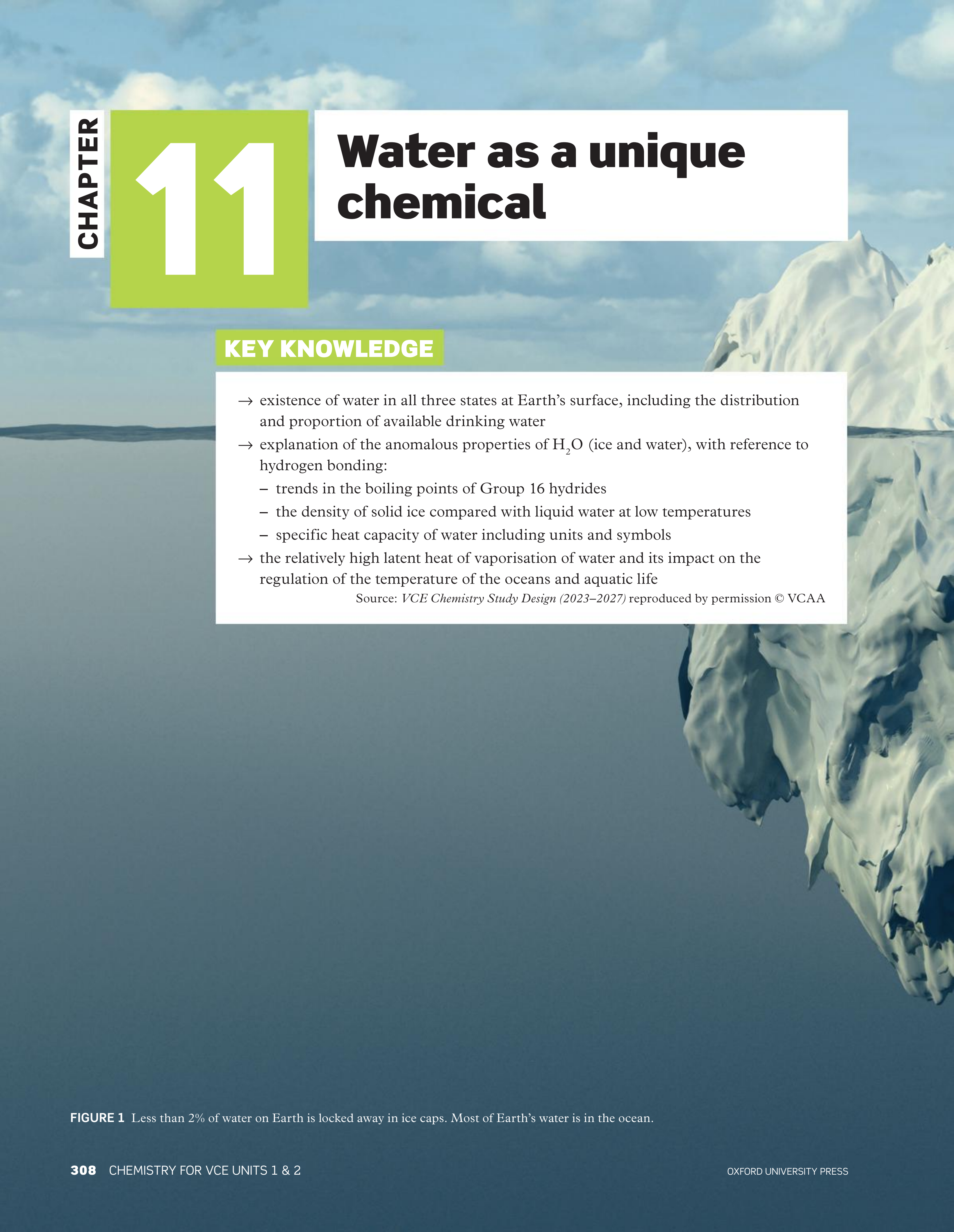


FIGURE 1 Less than 2% of water on Earth is locked away in ice caps. Most of Earth's water is in the ocean.

GROUNDWORK

In Chapter 11, you will learn about how the structure of water gives it its unique properties, and how these properties are important for the survival of living things.

This chapter will build on concepts you have already learnt in Chapter 3. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

11A What are the three types of intermolecular forces? Rank them from strongest to weakest.

11A Groundwork resource
Intermolecular forces

11B What features must a molecule have so it can form hydrogen bonds?

11B Groundwork resource
Hydrogen bonding

11C How do the intermolecular forces between molecules affect their melting and boiling points?

11C Groundwork resource
Intermolecular forces and physical properties

PRACTICALS

11.1A PRACTICAL: FIELDWORK	What is the quality of water sourced from different locations?	Page 515
11.1B PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT	How can contaminated water be purified for drinking?	pro
11.2 PRACTICAL: CONTROLLED EXPERIMENT	How can you create a heating curve for liquid water?	pro

11.1 The three states of water

KEY IDEAS

In this topic, you will learn that:

- water on Earth exists in all three states of matter: solid, liquid and gas
- most water on Earth is too salty or too dirty to drink, or not accessible
- salty water can be distilled or desalinated, and dirty water can be purified to produce clean and safe drinking water.

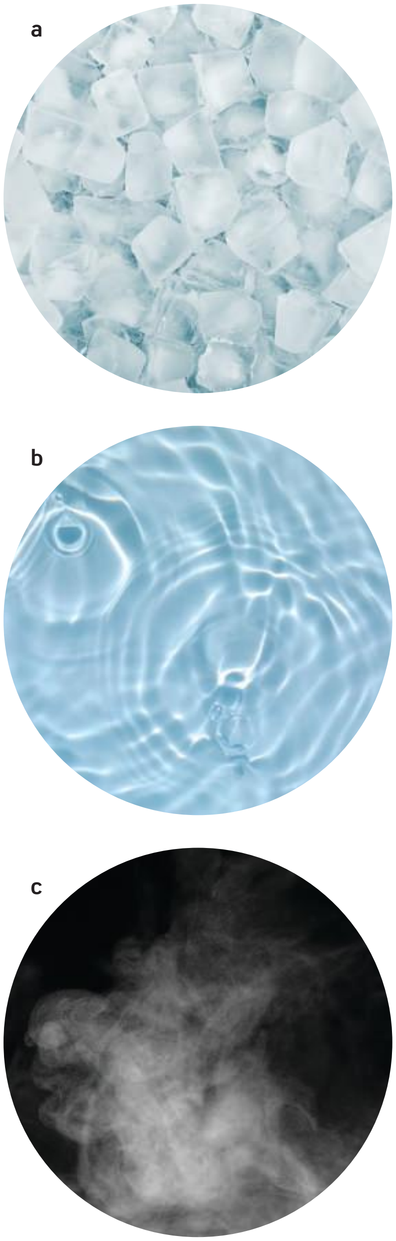


FIGURE 1 Water exists naturally on Earth in three different states: **a** solid, **b** liquid and **c** gas.

Water is the only substance that occurs naturally in three physical states of matter. Solid water can be found as ice cubes in your lemonade or in Antarctic glaciers. Liquid water flows through the pipes in your house, in your local river and the ocean. Gaseous water or water vapour can be observed from a boiling kettle or in the bathroom after you have had a hot shower.

The ability to exist in three different states is due to the structure of water.

Structure of water

Water is a polar, covalently bonded substance with the chemical formula H_2O. The two hydrogen atoms are bonded around the central oxygen atom. This results in a bent shape with a bond angle of 104.5° (Figure 2).

The difference in electronegativity between hydrogen and oxygen is 1.22, which makes the bond between them a polar covalent bond. The electrons spend more time around the more electronegative oxygen atom. This results in an uneven distribution of electrons or a dipole. You might remember from Chapter 3 that a partial negative charge (δ–) is assigned to the more electronegative oxygen atom and a partial positive charge (δ+) is assigned to the hydrogen atom.

This structure is what gives water its unique properties, which you will explore in Topic 11.2.

Water on Earth

Earth is known as the ‘blue planet’ because water is in very high abundance. Water is critical to support life on Earth. In fact, wherever there is water on Earth, life has also been found. More than 70% of Earth’s surface is covered in water. Water also exists below Earth’s surface and in the atmosphere.

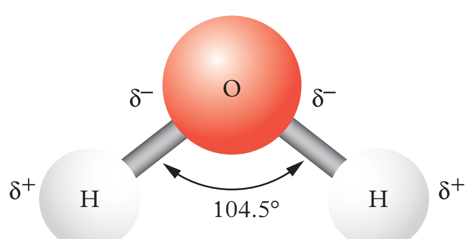


FIGURE 2 Water is a bent molecule.

11.1A
PRACTICAL: FIELDWORK

What is the quality of water sourced from different locations? Go to page **515**.

11.1B
PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can contaminated water be purified for drinking? Go to your **Student obook pro**.

Distribution of water on Earth

Almost all of Earth's water is in the oceans. This is approximately 1 338 000 000 km^3. Only a very small proportion (around 3.5%) of Earth's water occurs in other areas, such as ice caps, glaciers, lakes and rivers. The distribution of water on Earth is shown in Figure 3.

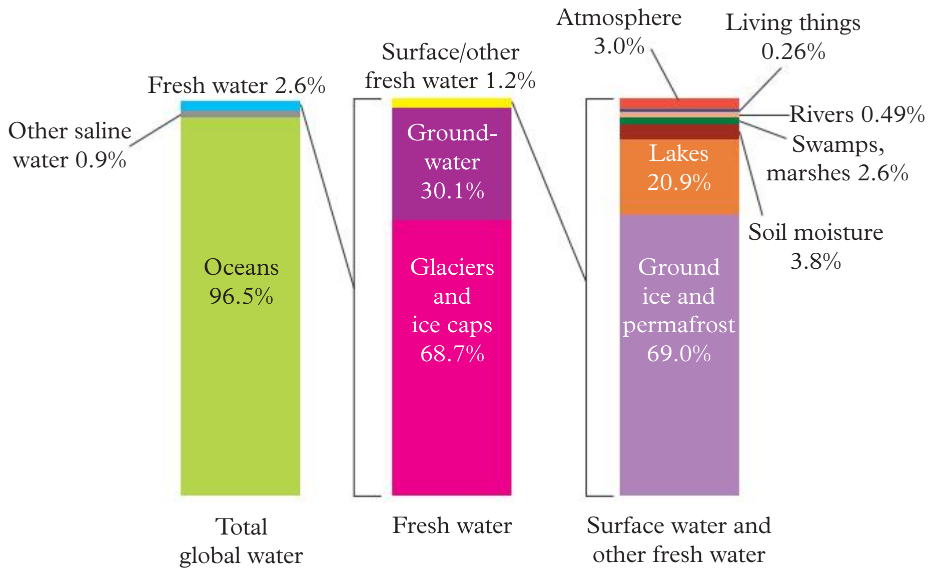


FIGURE 3 Water on Earth is distributed across many different sources.

Access to drinkable water on Earth

Only about 0.3% of Earth's water is **potable**, or safe to drink. The rest (99.7%) is either too salty or too dirty to drink. Of the 0.3% potable water, only a small proportion is accessible to humans. Much of it is frozen in glaciers or is very deep underground.

The most accessible potable water is called surface water. This includes rivers, streams and reservoirs. Most of Victoria's mains water supply comes from reservoirs that collect river water. In fact, cities are often built along rivers or at river estuaries partly because rivers provide are such excellent sources of accessible, potable water.

Drinking water can be sourced from the sea. However, this water must be purified before it is considered potable. Explore the processes involved in purification of water in Real-world chemistry 11.1.

potable
safe to drink

11.1 Real-world chemistry
Find me in your obook pro

11.1 CHECK YOUR LEARNING

Describe and explain

1 Describe where most of Victoria's mains water supply is sourced.

2 Explain why seawater is not potable.

3 Describe a situation in which water from living things or 'biological water' could be a useful supply of drinking water.

Apply, analyse and compare

4 Compare the availability of solid water and liquid water on Earth.

Design and discuss

5 Discuss the various methods that exist to obtain clean, potable water from ocean water.

11.2 Anomalous properties of water

KEY IDEAS

In this topic, you will learn that:

+ water has many unique properties because of its ability to form hydrogen bonds between its molecules
+ water has a much higher boiling point than the other group 16 hydrides because of the strong hydrogen bonding between water molecules
+ ice has a lower density than liquid water because ice has four hydrogen bonds around each H_2O molecule, which pushes the H_2O molecules far apart
+ the specific heat capacity of water is very high (4.18 J g^{-1} K^{-1}) because hydrogen bonds are stronger than most other types of intermolecular forces and can absorb more energy before being broken.

anomalous
unique; does not necessarily follow trends

Water is a substance with many unique or **anomalous** properties. This is due to its structure and ability to form strong hydrogen bonds with other water molecules. In this topic, we will explore how this gives water its properties, including:

- melting and boiling points
- density
- specific heat capacity.

Hydrogen bonding between water molecules

Study tip

Remember that hydrogen bonding only occurs when a hydrogen is attached to the atoms that are FON: fluorine, oxygen and nitrogen.

As you learnt in Chapter 3, water forms strong hydrogen bonds with other water molecules. These intermolecular forces form between the oxygen atom of one water molecule and a hydrogen atom of a neighbouring water molecule. Hydrogen bonding is the strongest type of intermolecular bonding.

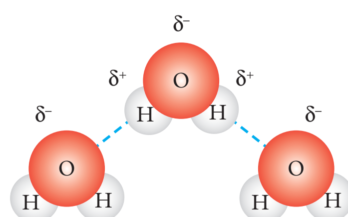


FIGURE 1 Water molecules form hydrogen bonds with other water molecules.

Melting and boiling points of water

The melting and boiling points of a compound are determined by the strength of the intermolecular forces present. Stronger intermolecular forces result in higher melting points because more heat energy is required to weaken the forces between molecules. Similarly, stronger intermolecular forces also result in a higher boiling point because more heat energy is required to completely break the forces between the molecules.

Trends in boiling points of the group 16 hydrides

The group 16 hydrides consist of two hydrogen atoms bonded to a group 16 atom. Examples are water (H_2O), hydrogen sulfide (H_2S), hydrogen selenide (H_2Se), hydrogen telluride (H_2Te) and polonium hydride (H_2Po). All these molecules have a bent shape.

11.2
PRACTICAL: CONTROLLED EXPERIMENT

How can you create a heating curve for liquid water? your **Student obook pro.**

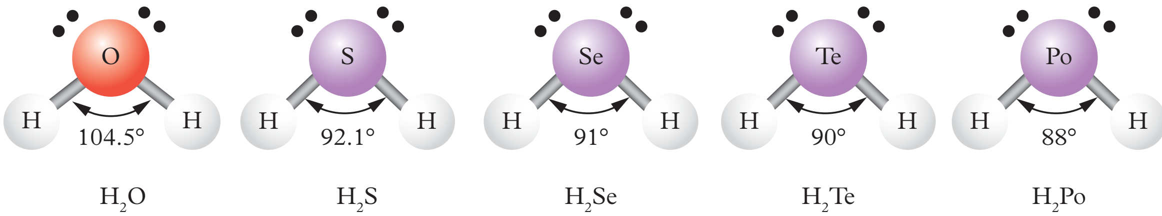


FIGURE 2 The group 16 hydrides are bent molecules.

Two interesting trends can be observed in the boiling points of the group 16 hydrides:

1. a general increase in boiling point going down group 16 in the periodic table
2. an anomalous result for water due to hydrogen bonding.

The number of electrons increases down group 16 in the periodic table. This means there are stronger dispersion forces. More heat energy is required to overcome the intermolecular bonding and turn the liquid hydride into a gas. Dipole–dipole attractions are also present.

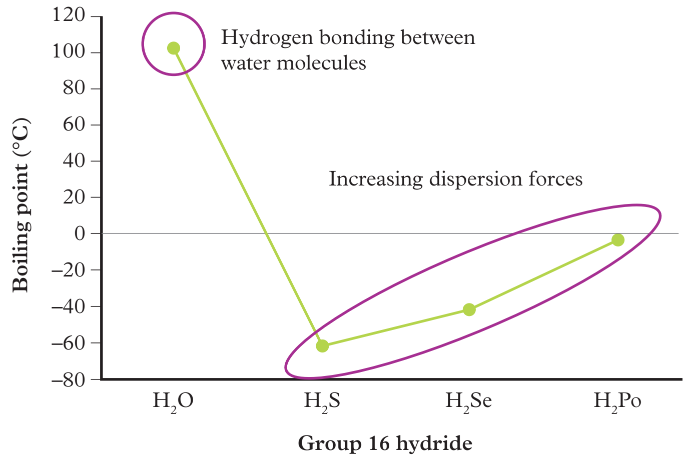


FIGURE 3 The boiling points of the group 16 hydrides

Water has a much higher boiling point than the other group 16 hydrides because of the presence of polar O–H covalent bonds. This allows water to form hydrogen bonds with other water molecules. Because water can form strong hydrogen bonds, more heat energy is required to overcome the intermolecular bonding and turn liquid water into water vapour.

11.2 SKILL DRILL

Considering safety and ethics when planning experiments

Key science skill: Comply with safety and ethical guidelines

Your Chemistry teacher wants you and your lab partner to design an experiment to identify the boiling points of three different substances: water (H_2O), hydrochloric acid (HCl) and ethanol (CH_3OH). Once the two of you have finished writing your hypothesis, aim, materials and method, and have set up a table to input your results, you show your teacher. However, they send you back to your desk and tell you that you have forgotten some very important things – safety and ethics.

Practise your skills

1. Identify three hazards you will need to include in your risk assessment.
2. For one of the hazards listed in your answer to Question 1, describe one control you can implement to prevent accidents, injuries or harm to you and others.
3. Identify two ethical considerations that you will need to make before you commence the experiment.

Need help conducting your experiment safely and ethically? See Topics 1.5 and 1.6 (pages 16–18).

FIGURE 4 Ice has a lower density than water, allowing it to float on top of the water.

Density of ice and water

Ice floats on water because solid ice has a lower **density** than liquid water (0.917 g cm^{-3} compared with 1.00 g cm^{-3}). This means that the water molecules in ice are further apart than they are in liquid water.

Ice has a lower density than water because, at temperatures below 0°C, the water molecules vibrate slowly enough to allow each water molecule to form four stable hydrogen bonds with other water molecules. The length of each hydrogen bond spaces the water molecules far apart in a three-dimensional hexagonal lattice.

In liquid water, only two or three hydrogen bonds are formed around each water molecule. The bonds are less stable because the water molecules are moving more quickly. Because there are fewer intermolecular forces, the distance between molecules is shorter and the density is higher. The effect of temperature on density of water is shown in Figure 5.

density
the amount of matter contained in a specific volume; the compactness of a substance

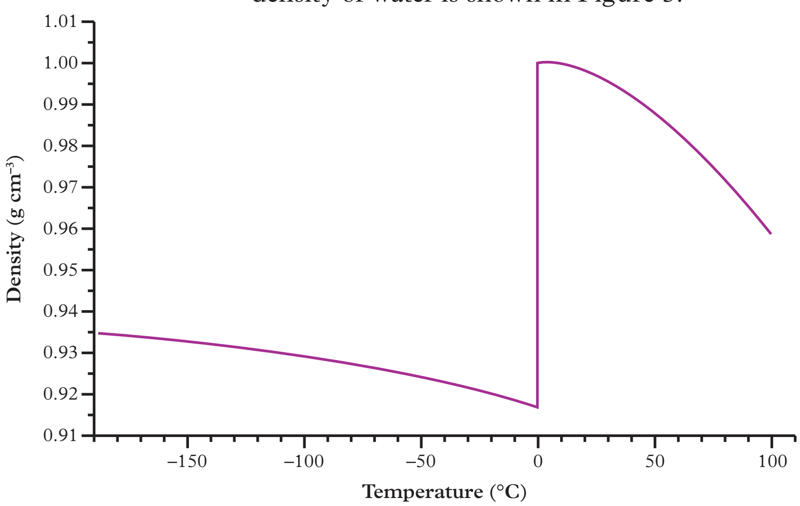


FIGURE 5 The density of water at different temperatures

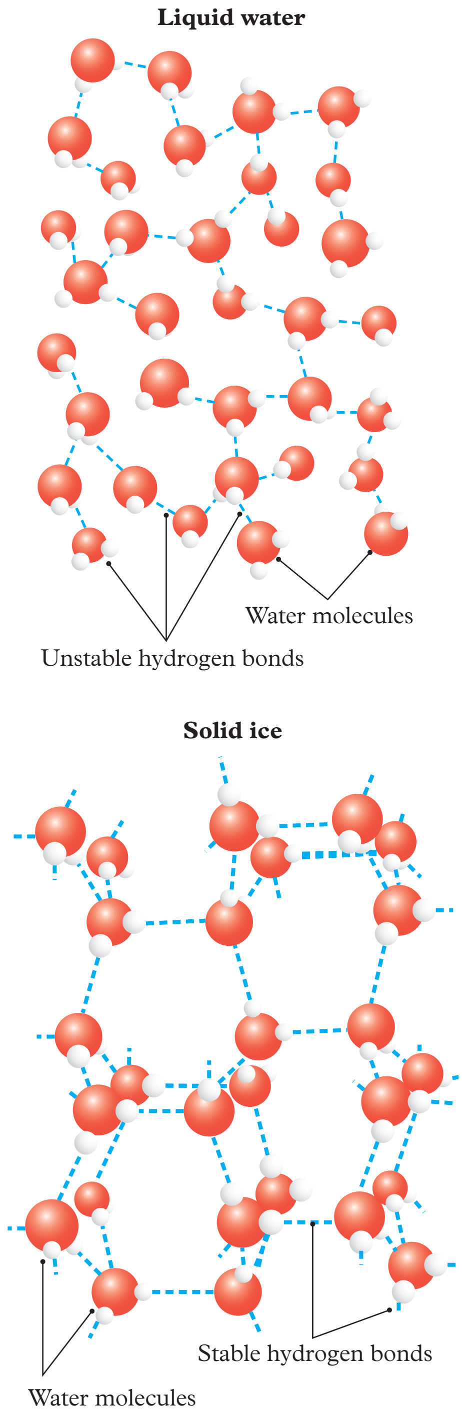


FIGURE 6 There are fewer and less stable hydrogen bonds between water molecules in liquid water compared to solid ice.

The fact that ice floats on water is crucial for the survival of aquatic life. When the water in lakes begins to freeze, any ice formed stays at the surface. This acts as an insulating barrier that prevents the water underneath from freezing. Even under extremely cold conditions, although a thick layer of ice forms, most of the water underneath the ice stays in liquid form. This allows aquatic life to survive underneath the ice.

Specific heat capacity of water

A very large amount of heat energy is required to increase the temperature of water. This property is called **specific heat capacity** and is represented by the symbol *c*. Specific heat capacity is measured in J g^{-1} K^{-1}. It is the amount of energy required to raise 1 g of substance by 1°C (or 1 K, which is the same). The specific heat capacities of some substances are shown in Table 1.

Water's high heat capacity is also due to water's ability to hydrogen bond. The hydrogen bonds between water molecules are stronger than the intermolecular forces between other substances of similar molar mass (e.g. methane). Hydrogen bonds can absorb more heat energy before they break. Try Challenge 11.2 to see if you can calculate the change in temperature of water when it is heated.

TABLE 1 The specific heat capacities of some substances

Material	Specific heat capacity (c) (J g^{-1} K^{-1})
Water ($H_2O(l)$)	4.18
Ethanol ($C_2H_5OH(l)$)	2.44
Octane ($C_8H_{18}(l)$)	2.22
Ice ($H_2O(s)$)	2.05
Magnesium ($Mg(s)$)	1.02
Air (mixture)	1.01
Sand	0.84
Zinc ($Zn(s)$)	0.39
Silver ($Ag(s)$)	0.23
Gold ($Au(s)$)	0.13

FIGURE 7 Animals can survive in the water underneath a frozen surface.

FIGURE 8 A large amount of heat energy is required to raise the temperature of water by 1°C.

specific heat capacity
the energy in joules required to increase the temperature of 1 g of a substance by 1°C

11.2 Challenge
Find me in your obook pro

11.2 CHECK YOUR LEARNING

Describe and explain

1 Explain why water has a higher boiling point than hydrogen sulfide.

2 Explain why the boiling point of hydrogen telluride is greater than the boiling point of hydrogen selenide.

3 Describe the difference between melting point and boiling point.

Apply, analyse and compare

4 Compare the intermolecular forces between water molecules in liquid water and solid ice to explain why water expands in volume when it freezes.

5 Consider the following covalent molecules:

I H_2O **II** H_2 **III** HCl **IV** $CHCl_2$

a Draw the structures of the molecules and label any dipoles with δ+ and δ–.

b Identify the intermolecular forces that exist between molecules.

6 Identify the molecules from Question 5 that you would expect to have the highest boiling point and the lowest boiling point.

Design and discuss

7 On a hot summer's day, the sand on the beach is very hot, but the ocean water is very cool. Discuss why this is the case, referring to the specific heat capacity of sand and water in your response.

11.3 Vaporisation of water

KEY IDEAS

In this topic, you will learn that:

- latent heat of vaporisation (L_{vap}) is the energy required to break intermolecular forces to change a liquid into a gas
- the relatively high latent heat of vaporisation of water makes it a very effective coolant.

In Topic 11.2, you learnt about the specific heat capacity of water, which is the heat energy required to change the temperature of 1 g of water by 1°C. In this topic, you will learn about a related concept – the latent heat of vaporisation.

Latent heat

latent heat
the amount of energy, in kJ, required to change the state of one mole of a substance

Latent heat is the energy required to change the physical state of one mole of a substance. It is measured in kJ mol^{-1}. For water, it is the energy required to change from ice to liquid water, and from liquid water to water vapour.

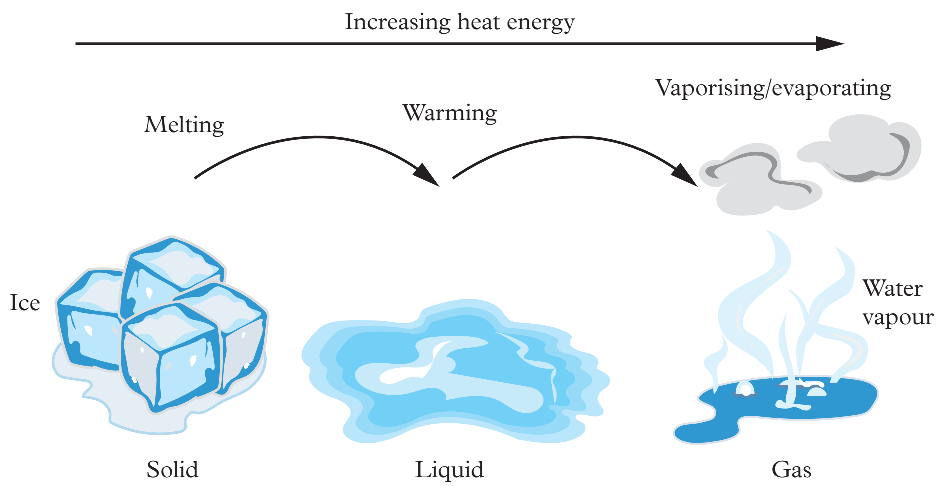


FIGURE 1 When heat energy is applied, ice melts into liquid water, which can then boil and vaporise into water vapour.

latent heat of fusion
the amount of energy, in kJ, required to melt one mole of substance from solid to liquid state at its melting point

Latent heat of fusion (L_{fus})

The heat energy required to completely melt one mole of a substance from the solid state into the liquid state is called the **latent heat of fusion** (L_{fus}).

For water, it is the amount of energy needed to weaken intermolecular bonds and disrupt the lattice structure of ice. L_{fus} is 6.0 kJ mol^{-1}. In other words, 6.0 kJ of energy is required to change one mole of ice (solid) into water (liquid). During this process, the temperature of H_2O remains the same at 0°C.

latent heat of vaporisation
the amount of energy, in kJ, required to evaporate one mole of substance from liquid to gaseous state at its boiling point

Latent heat of vaporisation (L_{vap})

Latent heat of vaporisation (L_{vap}) is defined as the energy required to change one mole of substance from the liquid state into the gaseous state. L_{vap} represents the amount of energy required to break all the intermolecular forces between molecules in the liquid state as they are vaporised (or evaporated) into the gaseous state.

Water has a relatively high L_{vap} (40.7 kJ mol^{-1}). This means that more heat energy can be absorbed before liquid water changes to water vapour (gas). During this process, the temperature of water remains the same at 100°C.

TABLE 1 The latent heats of vaporisation (L_{vap}) of some substances

Substance	Latent heat of vaporisation (kJ mol^{-1})
Water (H_2O)	40.7
Ethanol (C_2H_5OH)	38.6
Octane (C_8H_{18})	34.0
Methane (CH_4)	8.18
Hydrogen (H_2)	0.45
Helium (He)	0.084

The heating curve for water

Let's explore the latent heat of fusion and vaporisation by looking at the heating curve for water. As shown in Figure 2, this heating curve can be considered in five sections:

1. **Heating ice**: As heat energy is absorbed, the H_2O molecules in ice vibrate faster in their fixed positions and gain kinetic energy (and thus increase in temperature).
2. **Melting ice:** 0°C is the melting point when ice begins to melt into liquid water. Some of the hydrogen bonds between the H_2O molecules are broken and the molecules start to flow over each other. The melting ice remains at the same temperature (0°C) as heat continues to be absorbed, so the curve plateaus. The total amount of heat absorbed in the melting process is called the latent heat of fusion.
3. **Heating water:** Once all the ice has melted, more heat energy is absorbed. The H_2O molecules continue to flow over each other. They gain kinetic energy and temperature increases to 100°C.
4. **Vaporising water:** 100°C is the boiling point when liquid water begins to vaporise. All the remaining hydrogen bonds between H_2O molecules break, allowing the molecules to separate. The boiling water stays at the same temperature (100°C) as heat continues to be absorbed, so the curve plateaus. The total amount of heat absorbed in the boiling or vaporising process is called the latent heat of vaporisation.
5. **Heating water vapour:** Once all the water has vaporised, more heat energy is absorbed. the H_2O molecules move faster. They gain more kinetic energy and temperature continues to rise beyond 100°C.

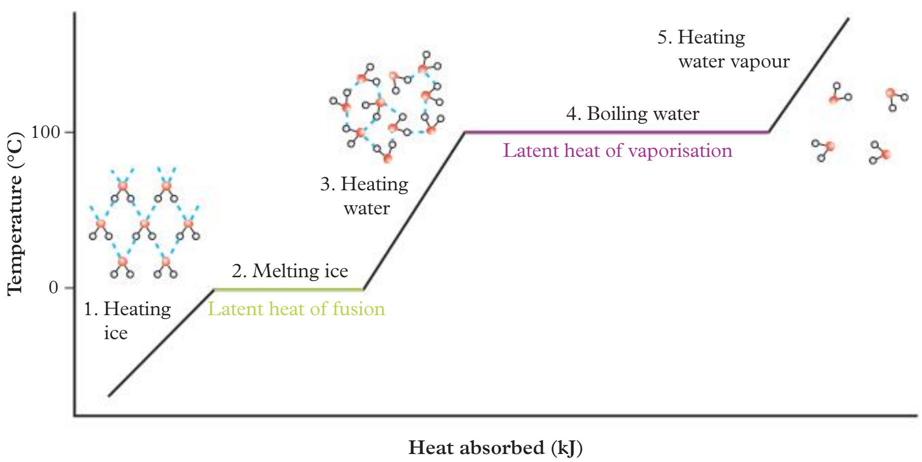


FIGURE 2 The heating curve for water

11.3 Worked example
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11.3 Worked example
Video demonstration

Worked example 11.3 shows you how to read a heating curve for benzene.

Water as a natural coolant

The high latent heat of vaporisation allows water to absorb large amounts of heat energy as it heats up and evaporates. This makes it useful for maintaining temperatures of natural bodies of water. It also helps to remove heat from a system, such as a living organism or a body of water, without dehydration. This ability of water to act as a natural **coolant** is very important for life on Earth.

coolant
a substance that removes heat

Maintaining the ocean temperature

Evaporation is a process that occurs at the surface of water. Because of its high L_{vap}, water can absorb more energy before it evaporates. Combined with water's high specific heat capacity, this helps to maintain a relatively constant ocean temperature.

This is important when temperatures increase during summer because the ocean and other bodies of water will absorb the heat. Therefore, there will not be rapid changes in temperature. This is critical for the survival of aquatic life that live in the oceans and rivers. In fact, over its entire lifetime, an organism that lives in the deep sea may only ever experience a 0.5°C change in temperature. This helps to keep its internal body temperature as constant as possible.

FIGURE 3 Aquatic animals such as turtles rely on the high specific heat capacity of water to maintain their body temperature.

FIGURE 4 The high latent heat of vaporisation of water prevents lakes from drying up.

Maintaining ocean temperatures within a narrow range is also important because the solubilities of gases in the ocean are sensitive to changes in temperature. You will explore this in more detail in Chapter 16. When concentrations of dissolved gases such as carbon dioxide (CO_2) increase in the ocean, this can result in the formation of carbonic acid (H_2CO_3), which makes it difficult for coral reefs to survive.

More CO_2 in the ocean also increases the acidity of the ocean. This can affect the ability of shellfish to produce their protective shells. You will learn more about this in Chapter 12.

Preventing water loss by evaporation from oceans, rivers, lakes and ponds

The high L_{vap} of water reduces the rate of evaporation from the surface of lakes and ponds. This is because a large amount of energy is required to break the hydrogen bonds between H_2O molecules in the liquid state during evaporation.

If the L_{vap} of water was lower than 4.18 J g^{-1} K^{-1}, then water would evaporate faster from lakes, ponds and reservoirs, and possibly even the ocean. This would mean that aquatic life would have a smaller volume of water in which to live.

Cooling the body

Unlike aquatic animals, terrestrial organisms experience much greater and more rapid changes in temperature. You might have noticed that the weather in Victoria fluctuates a lot – some days, the maximum temperature can be 20°C higher than the minimum. Fortunately, we have adaptations that allow us to cope with these changes.

Plants and animals make use of the high L_{vap} of water to cool themselves in hot weather. Water can absorb heat energy from the organism's body. This is then released as different forms of moisture: dogs pant, humans sweat, and plants open their stomata (pores on the surface of a leaf) to increase transpiration (the evaporation of water from the leaves of a plant).

Evaporating water decreases the temperature of the plant or animal. Water's high L_{vap} means that a large amount of heat energy is absorbed per mole of water. In other words, more heat can be absorbed per drop of sweat. This is important because it prevents the organism from losing too much fluid (becoming dehydrated) as it cools down.

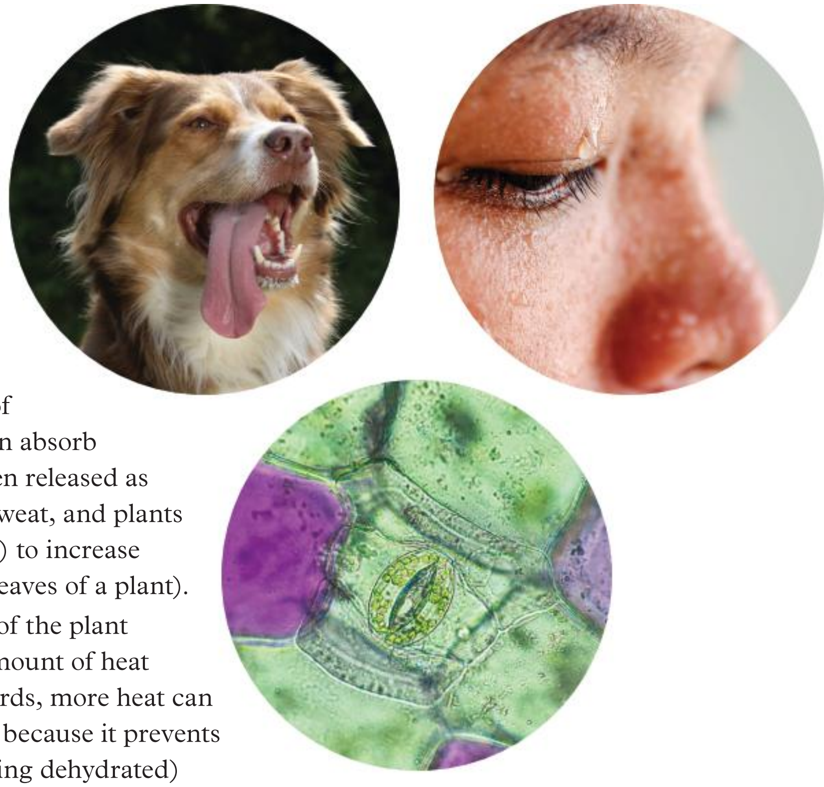


FIGURE 5 Dogs pant, humans sweat and plants release water from their stomata to cool down.

11.3 CHECK YOUR LEARNING

Describe and explain

1 Define 'latent heat'.

2 Identify the latent heat of vaporisation of water (in kJ mol^{-1}).

3 Explain why water has a relatively high latent heat of vaporisation.

Apply, analyse and compare

4 Compare latent heat of fusion and latent heat of vaporisation.

5 A heating curve for ethanol is shown in Figure 6.

  a Identify the physical state of ethanol at sections:

   i A–B   ii C–D   iii E–F

  b Identify the processes occurring at sections:

   i B–C   ii D–E

  c Describe what is happening to the ethanol molecules at the following sections of the graph. Refer to kinetic energy and intermolecular forces in your response.

   i A–B   ii C–D   iii E–F

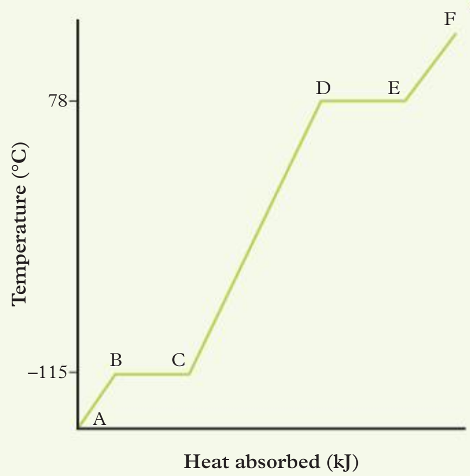


FIGURE 6 A heating curve for ethanol

  d Redraw the graph in your work book and label the latent heat of fusion and latent heat of vaporisation.

Design and discuss

6 Design an experiment to obtain a heating curve such as the one in Figure 6.

CHAPTER

11 Review

Chapter summary

11.1

- Water is found on Earth in all three of the states of matter: solid, liquid and gas.
- Only 0.3% of the water on Earth is potable (drinkable) and most potable water is not easily accessible.
- Salt can be removed from seawater by desalination or distillation. Dirty water can be purified by filtration, boiling, irradiation with ultraviolet light, chemical treatment or a combination of these technologies. These methods can be used to produce clean, safe drinking water.

11.2

- Water molecules have strong intermolecular hydrogen bonds between them.
- The boiling points of the group 16 hydrides increase down the group. However, water is an exception to this trend.
- The boiling point of water (100°C) is higher than for the other group 16 hydrides because water molecules form hydrogen bonds. These intermolecular forces are stronger than the dispersion forces and dipole–dipole forces between the other group 16 hydrides.
- Below 0°C, water molecules form four stable hydrogen bonds with other water molecules. This pushes the water molecules apart in a hexagonal lattice, making ice less dense than water.
- Water has a very high specific heat capacity of 4.18 J g^{-1} K^{-1}. The hydrogen bonds between water molecules are stronger than most other types of intermolecular forces. Therefore, they absorb more energy before they break.

11.3

- Water has a very high latent heat of fusion (L_{fus}), which means that a large amount of energy is required to break the hydrogen bonds and change solid ice into liquid water.
- Water has a very high latent heat of vaporisation (L_{vap}), which means that a large amount of energy is required to break the hydrogen bonds and change liquid water into a gas.
- The relatively high latent heat of vaporisation (L_{vap}) of water makes it a very effective coolant.
- Together, the high specific heat capacity and latent heat of vaporisation keep the ocean temperature stable, prevent bodies of water from drying up, and help living organisms regulate their internal body temperatures.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• explain that water exists in all three states at Earth's surface	☐	☐	☐	Go back to Topic 11.1
• describe the distribution and proportion of available drinking water	☐	☐	☐	Go back to Topic 11.1
• use hydrogen bonding to explain the anomalous properties of H_2O (ice and water), including trends in the boiling points of group 16 hydrides, the density of solid ice compared with that of liquid water at low temperatures, and the specific heat capacity of water, including units and symbols	☐	☐	☐	Go back to Topic 11.2
• describe latent heat of vaporisation and explain why water has a relatively high latent heat of vaporisation	☐	☐	☐	Go back to Topic 11.3
• explain how the properties of water affect the regulation of the temperature of the oceans and aquatic life	☐	☐	☐	Go back to Topic 11.3

Revision questions

Multiple choice

1 During the evaporation of water, heat energy is:

A absorbed to break hydrogen bonds.
B released to break hydrogen bonds.
C absorbed to form hydrogen bonds.
D released to form hydrogen bonds.

2 Most of Earth's water is:

A in the atmosphere.
B in glaciers.
C underground.
D in the oceans.

3 Ocean water can be desalinated by:

A chlorination.
B boiling.
C distillation.
D filtration.

4 The boiling point of H_2O is much higher than that of methane mostly because water has:

A more dispersion forces than methane.
B more electrons than methane.
C weaker hydrogen bonds than methane.
D stronger intermolecular bonds than methane.

5 The density of solid ice is lower than that of liquid water because:

A dispersion forces between water molecules stabilise below 0°C and hold the water molecules further apart in a hexagonal lattice.
B dispersion forces between water molecules stabilise below 0°C and hold the water molecules closer together in a hexagonal lattice.
C hydrogen bonds between water molecules stabilise below 0°C and hold the water molecules closer together in a hexagonal lattice.
D hydrogen bonds between water molecules stabilise below 0°C and hold the water molecules further apart in a hexagonal lattice.

6 55 kJ of heat energy is supplied to 450 g water. If the initial temperature of the water was 11°C, what is the final temperature of the water?

A 11°C

B 29°C

C 40°C

D 302°C

7 Water is an effective coolant because its hydrogen bonds:

A release large amounts of energy when broken.

B absorb a large amount of energy before they break.

C are stronger than covalent bonds.

D are longer than covalent bonds.

8 The latent heat of vaporisation is the:

A temperature at which a solid will sublime into a gas.

B energy required for a liquid to vaporise into a gas.

C temperature at which a liquid will vaporise into a gas.

D energy required for a liquid to condense into a gas.

9 Which of the following statements is incorrect?

A The energy required for one mole of ice to melt into liquid water is called the latent heat of fusion.

B Beyond 100°C, the water molecules absorb more heat energy and their kinetic energy increases.

C By the time the temperature has increased to 30°C, there will only be water in liquid form.

D As water is boiling, the water molecules temporarily stop absorbing heat energy.

10 The anomalous properties of water are primarily due to its:

A ability to form hydrogen bonds with itself.

B bent shape.

C small size.

D electronegative oxygen.

Short answer

Describe and explain

11 Explain why water has a higher boiling point than hydrogen selenide.

12 Explain why water has a relatively high latent heat of vaporisation.

13 Explain why humans cannot drink seawater.

14 Describe the behaviour of H_2O molecules when ice melts into liquid water.

15 Describe the behaviour of H_2O molecules when liquid water vaporises into water vapour.

16 Explain how the high specific heat capacity of water is important for the survival of aquatic animals.

Apply, analyse and compare

17 Compare specific heat capacity and latent heat.

18 Explain, by referring to the types of intermolecular forces in each substance, the difference in the latent heats of vaporisation (L_{vap}) for hydrogen fluoride gas (25.2 kJ mol^{-1}) and fluorine gas (3.27 kJ mol^{-1}).

19 Explain why water is an effective coolant used in gaming computers.

FIGURE 1 A cooling system for a gaming computer

20 Three substances, all initially at 20.0°C, are being heated at a constant rate in the experimental set-up shown in Figure 2. The specific heat capacities for the substances, from left to right, are 4.18 J g^{-1} mol^{-1}, 2.46 J g^{-1} mol^{-1} and 1.67 J g^{-1} mol^{-1}.

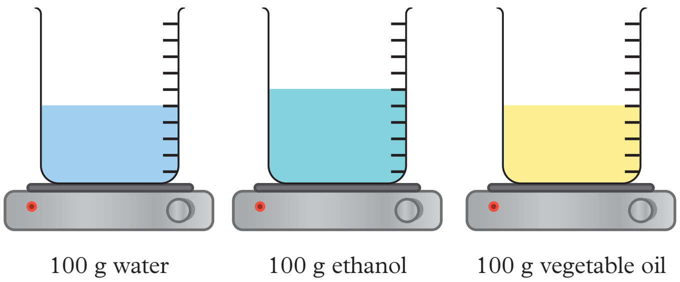


FIGURE 2 An experimental set-up

- **a** Which substance would you expect to reach 32.0°C first? Explain your choice.
- **b** List any safety precautions you would need to consider when conducting this experiment.
- **c** Use Figure 2 to determine which of the three liquids in the experiment has the lowest density.

21 Ice floats on water.

- **a** Explain, in terms of bonding, why ice floats on water.
- **b** Explain how this is advantageous to aquatic life.

22 A heating curve for 100.0 g of an unknown substance, initially a solid at –4°C, is shown in Figure 3.

- **a** Use the heating curve to find the:
  - **i** melting point of the substance
  - **ii** boiling point of the substance.
- **b** Redraw the graph in your workbook and label the latent heat of fusion and latent heat of vaporisation.

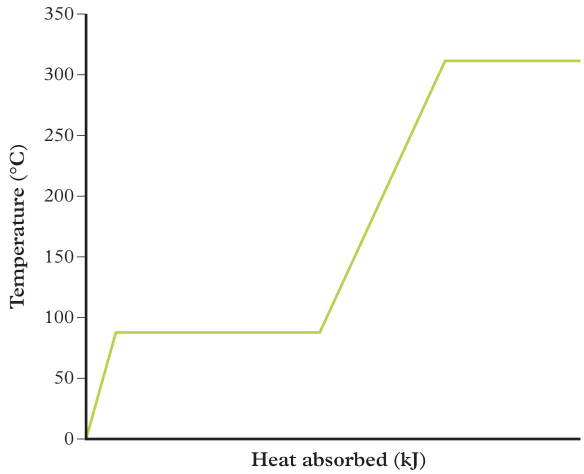


FIGURE 3 A heating curve for an unknown substance

23 Suggest why the latent heat of vaporisation of xenon is higher than that of argon.

$L_{vap}(Ar) = 6.44$ kJ mol^{-1}

$L_{vap}(Xe) = 12.6$ kJ mol^{-1}

Design and discuss

24 Design an experiment that could determine the specific heat capacity of an unknown substance.

25 Design an experiment that could determine the latent heat of vaporisation of an unknown substance. The substance has a boiling point of around 80°C.

26 Discuss the following statement: 'The existence of life is only possible because of hydrogen bonding'.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 12

Acid–base reactions

KEY KNOWLEDGE

→ the Brønsted–Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and the writing of balanced ionic and full equations, with states, for their reactions in water

→ the distinction between strong and weak acids and strong and weak bases, and between concentrated and dilute acids and bases, including common examples

→ neutralisation reactions to produce salts:

  → reactions of acids with metal carbonates and hydroxides, including balanced full and ionic equations, with states

  → types of antacids and their use in the neutralisation of stomach acid

→ use of the logarithmic pH scale to rank solutions from most acidic to most basic; calculation of pH for strong acid and strong base solutions of known concentration using the ionic product of water (K_w at a given temperature)

→ accuracy and precision in measurement as illustrated by the comparison of natural indicators, commercial indicators, and pH meters to determine the relative strengths of acidic and basic solutions

→ applications of acid–base reactions in society: for example, natural acidity of rain due to dissolved CO_2 and the distinction between the natural acidity of rain and acid rain, or the action of CO_2 forming a weak acid in oceans and the consequences for shell growth in marine invertebrates

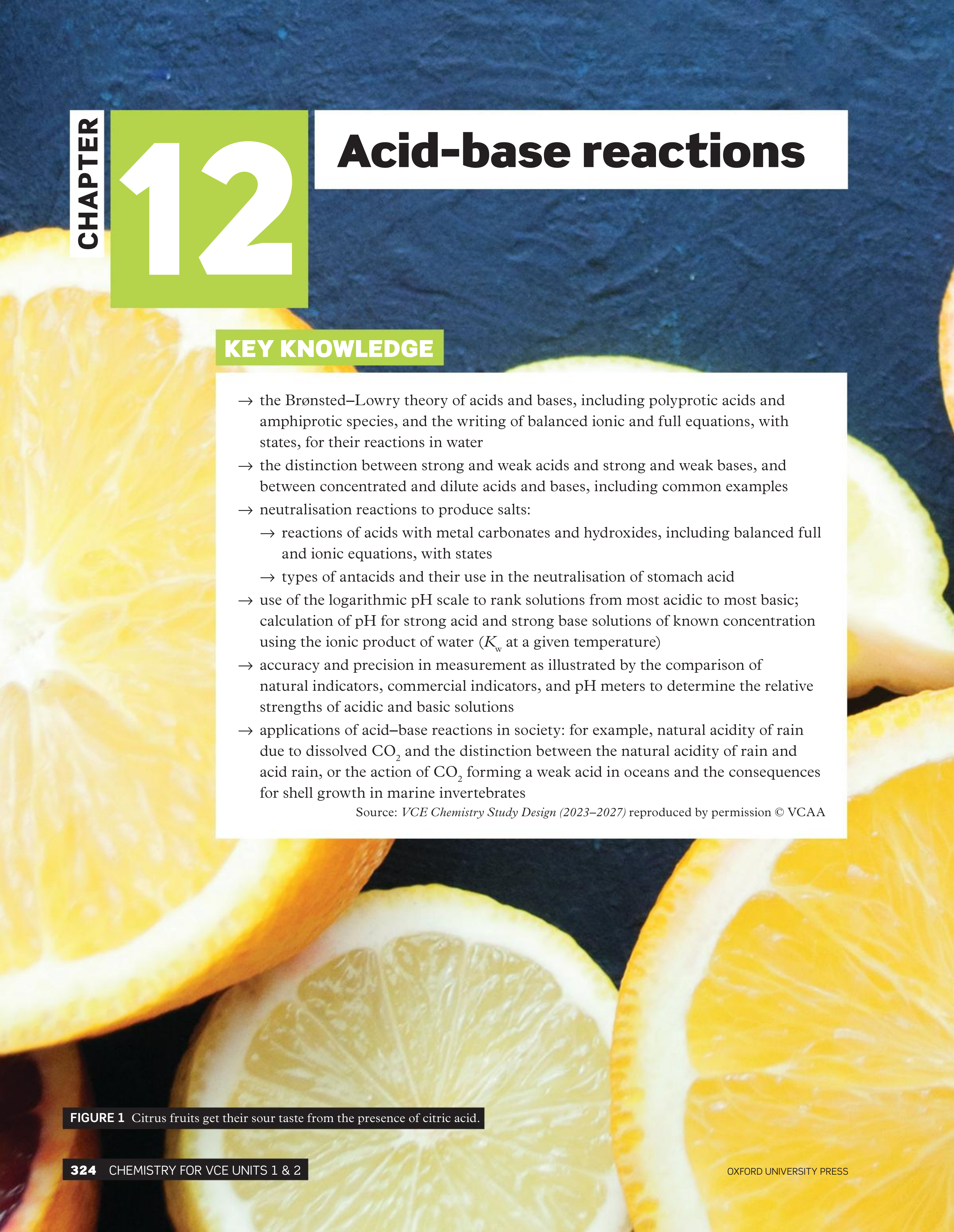
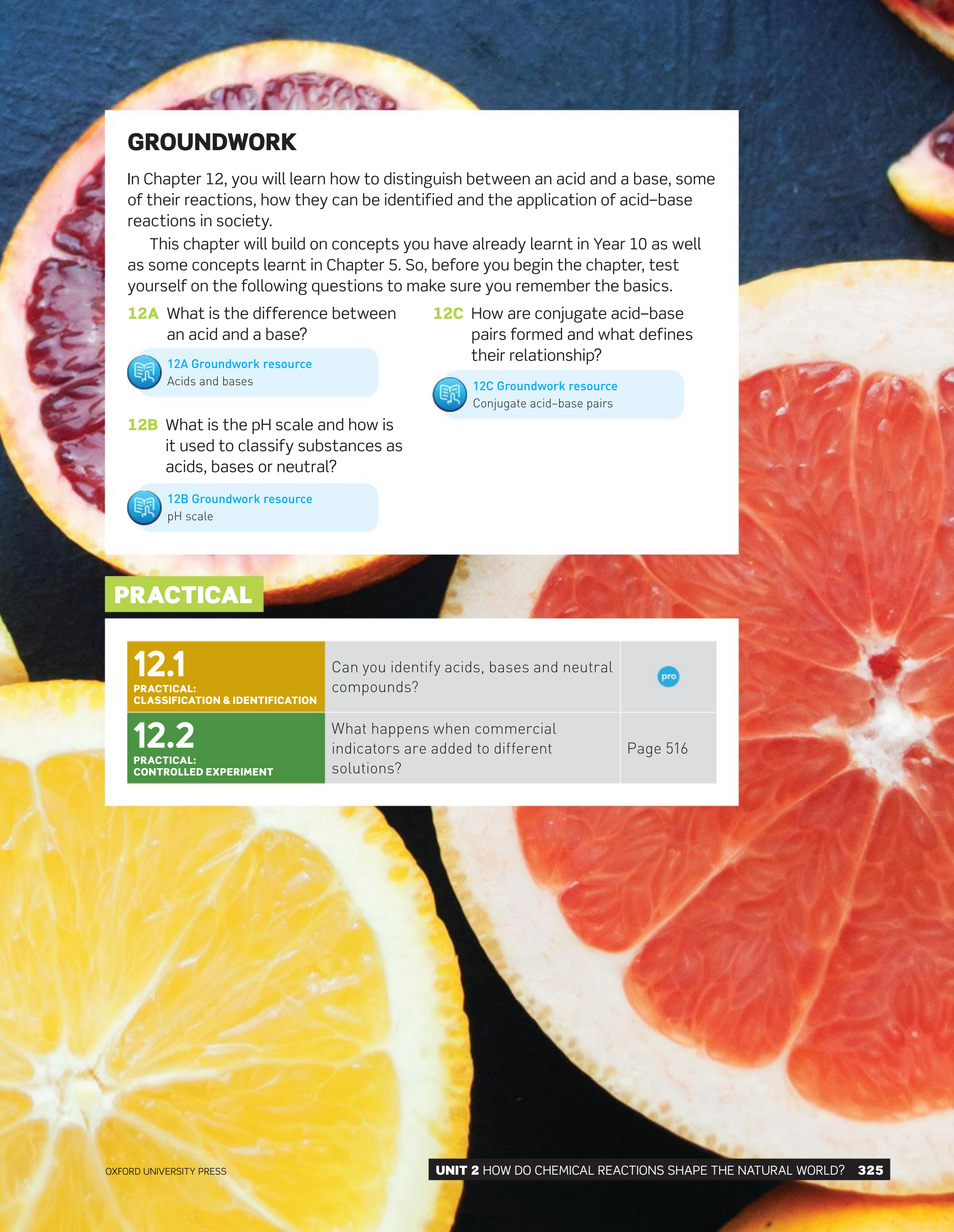


FIGURE 1 Citrus fruits get their sour taste from the presence of citric acid.



GROUNDWORK

In Chapter 12, you will learn how to distinguish between an acid and a base, some of their reactions, how they can be identified and the application of acid–base reactions in society.

This chapter will build on concepts you have already learnt in Year 10 as well as some concepts learnt in Chapter 5. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

12A What is the difference between an acid and a base?

12A Groundwork resource
Acids and bases

12B What is the pH scale and how is it used to classify substances as acids, bases or neutral?

12B Groundwork resource
pH scale

12C How are conjugate acid–base pairs formed and what defines their relationship?

12C Groundwork resource
Conjugate acid–base pairs

PRACTICAL

12.1 PRACTICAL: CLASSIFICATION & IDENTIFICATION	Can you identify acids, bases and neutral compounds?	pro
12.2 PRACTICAL: CONTROLLED EXPERIMENT	What happens when commercial indicators are added to different solutions?	Page 516

12.1 The Brønsted–Lowry theory of acids and bases

KEY IDEAS

In this topic, you will learn that:

- the Brønsted–Lowry theory defines acids as hydrogen ion (H^+) or proton donors and bases as hydrogen ion (H^+) or proton acceptors, forming conjugate acid–base pairs
- polyprotic acids can donate more than one proton from each molecule
- amphiprotic substances can act as an acid or a base in a reaction
- acid–base reactions in water can be represented by balanced ionic and full chemical equations, with states included.

Acids and bases

acid
a molecule that can donate a proton to a base

base
a molecule that can accept a proton from an acid

Acids and **bases** are groups of chemical compounds with opposite properties; they are often found in the laboratory as well in everyday life. Acids often taste sour and can be corrosive to substances or skin. Examples include hydrochloric acid in the stomach, sulfuric acid used in car batteries, vinegar, citrus fruit juices, aspirin, the amino acids in proteins and the nucleic acids in living organisms.

a  b 

FIGURE 1 **a** Lemons and other citrus fruits contain citric acid, which makes them taste sour. **b** Soap bars contain bases, which make the bar slippery when wet.

Bases often taste bitter, feel slippery and may irritate or burn skin. Examples include baking soda, soap, ammonia in cleaning products, caustic soda and antacids used to ease stomach reflux.

In this chapter, you will learn about the properties and reactions of acids and bases. You will also learn about how you can identify acids and bases and their uses in everyday applications.

The Brønsted–Lowry theory

In 1923, chemists Johannes Brønsted and Thomas Lowry independently defined acids as hydrogen ion (H^+) or proton donors and bases as hydrogen ion (H^+) or proton acceptors. Acid–base reactions involve a proton transfer in which a hydrogen ion is transferred from the acid to the base. This enables acids and bases to react together readily in aqueous solutions.

acid–base reaction
the reaction that occurs between an acid and a base

The Brønsted–Lowry theory states that:

- acids donate protons
- bases accept protons
- an **acid–base reaction** involves a proton being donated from an acid and accepted by a base.

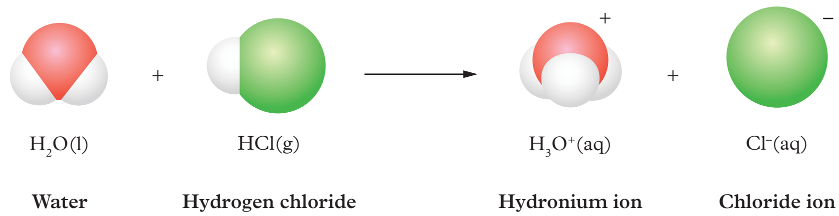


FIGURE 2 The reaction of hydrogen chloride with water is an example of an acid–base reaction.

Figure 2 shows an example of an acid–base reaction in which each HCl molecule donates a proton (H^+) to a water molecule. This means HCl is the acid. Each water molecule accepts a proton (H^+) from HCl, so water acts as the base.

Conjugate pairs

The acid–base reaction in Figure 2 shows that HCl and Cl^- can be formed from each other by either the loss or gain of a proton. The same applies to H_2O and H_3O^+. When an acid and base pair differ by one proton such as the HCl and Cl^- pair or the H_2O and H_3O^+ pair, they are referred to as an acid–base **conjugate pair** (Figure 3).

conjugate pair
two compounds that differ by the presence or absence of an H^+ or a proton in their formulas

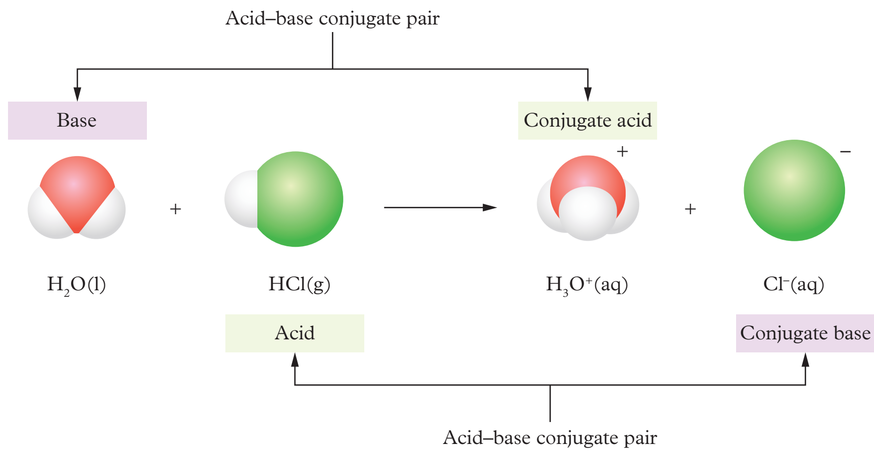


FIGURE 3 Conjugate acid–base pairs consist of an acid and a base that differ by one proton.

In Figure 3, the H_3O^+ is an example of a **conjugate acid** because it has accepted a proton into its structure and now has the potential to donate a proton to another base. The Cl^- is an example of a **conjugate base** because it formed when HCl lost a proton and now has the potential to accept a proton from another acid.

You can determine the conjugate base of an acid by subtracting one H^+. The conjugate acid of a base can be determined by adding an H^+. This relationship is shown in Figure 4.

conjugate acid
a molecule formed when a base accepts a proton

conjugate base
a molecule formed when an acid loses a proton

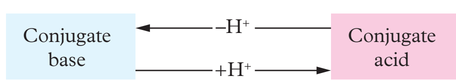


FIGURE 4 Conjugate acid–base pairs are made up of an acid and a base that differ by one proton.

Ionisation

The reaction of hydrochloric acid with water shown in Figure 2 is also considered an **ionisation reaction** because uncharged molecules have reacted in such a way that charged ions have been produced.

ionisation reaction
a reaction that results in molecules losing or gaining an electrostatic charge

Polyprotic acids

monoprotic
can donate one proton per molecule

polyprotic
can donate more than one proton per molecule

Acids can be classified according to their ability to donate protons. Some acids such as hydrochloric acid can only donate one proton to a base. These are called **monoprotic** acids ('mono' meaning 'one'). A **polyprotic** acid is an acid that can donate more than one proton to a base.

diprotic
can donate two protons per molecule

Diprotic acids can donate two protons ('di' meaning 'two'). For example, when sulfuric acid (H_2SO_4) reacts in water, the aqueous H_2SO_4 loses one H^+ during its first ionisation to become aqueous HSO_4^-. HSO_4^- then loses another H^+ during the second ionisation to become SO_4^{2-}.

$$H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq) \qquad \text{First ionisation}$$

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq) \qquad \text{Second ionisation}$$

Double arrows $\rightleftharpoons$ indicate that a reaction is reversible.

triprotic
can donate three protons per molecule

Triprotic acids can donate three protons. For example, when phosphoric acid (H_3PO_4) reacts in water, it loses one H^+ during its first ionisation to become $H_2PO_4^-$. The resulting species can lose two more H^+ to form HPO_4^{2-} and PO_4^{3-}.

$$H_3PO_4(aq) + H_2O(l) \rightarrow H_2PO_4^-(aq) + H_3O^+(aq) \qquad \text{First ionisation}$$

$$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq) \qquad \text{Second ionisation}$$

$$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq) \qquad \text{Third ionisation}$$

Amphiprotic species

amphiprotic
can both donate and accept a proton, allowing it to act as either an acid or a base

Amphiprotic substances can act as either acids or bases depending on the acid–base characteristics of the other reactant involved. (Note that the prefix 'amphi' means 'both' or 'on both sides'). Water is one example of an amphiprotic substances.

When aqueous HNO_3 reacts with water, water acts as a base and accepts a proton to form H_3O^+.

$$HNO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_3^-(aq)$$

When aqueous NH_3 reacts with water, water acts as an acid and donates a proton to form OH^-.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Figure 5 shows a variety of amphiprotic species and what they can form during acid–base reactions.

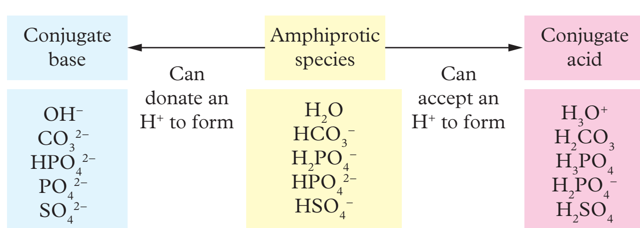


FIGURE 5 A list of amphiprotic species and what they can form during acid–base reactions

Constructing acid–base balanced equations

Like ionic compounds, acids and bases also dissociate to form ions in the presence of water. Acid–base reactions in water can be represented by balanced full and ionic equations, with states included. For example, the full balanced equation for the dissociation of aqueous sodium hydrogen carbonate in the presence of water is:

$$NaHCO_3(aq) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq) + H_2CO_3(aq)$$

Recall that when constructing an ionic equation, only the species taking part in the reaction are listed and not the spectator ions. Refer to the solubility table (Table 1 in Topic 5.4) to recall which type of ions are the spectator ions. (Hint: they are always soluble.)

For the reaction of aqueous sodium hydrogen carbonate in the presence of water, the ionic equation would be written as:

$$HCO_3^-(aq) + H_2O(l) \rightarrow OH^-(aq) + H_2CO_3(aq)$$

If unsure of which species to include in an acid–base ionic equation, make sure that all ionic species and the acids are dissociated, then remove the spectator ions from both sides of the equation arrow (see Worked example 12.1).

Study tip

Remember that an acid is a proton (H^+) donor, a base is a proton acceptor and acid–base conjugate pairs are formed when an acid reacts with a base.

12.1 WORKED EXAMPLE

CONSTRUCTING BALANCED ACID-BASE REACTION EQUATIONS

Construct the balanced ionic and full equations for the reaction of hydroiodic acid (HI) in water. Identify the conjugate acid–base pairs in the reaction.

Solution

Think	Do
Step 1: Write an equation that shows the full reaction of hydroiodic acid and water then check for the balance of atoms and neutral charges.	$HI(aq) + H_2O(l) \rightarrow H_3O^+(aq) + I^-(aq)$ Ratio of atoms: *Atom* / *Left side of arrow* / *Right side of arrow* H / 3 / 3 I / 1 / 1 O / 1 / 1 Atoms are balanced. Ratio of charges: *Left side of arrow* / *Right side of arrow* Neutral / +1 and –1 = Neutral Charges are balanced.
Step 2: Use a solubility table to identify the spectator ion in the equation.	According to the solubility table in Chapter 5, page 155, the spectator ion is $I^-(aq)$.
Step 3: Remove the spectator ion to write the balanced ionic equation.	The net ionic equation is: $H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$
Step 4: Check to see if all atoms and charges are equal on both sides of the ionic equation.	In this case, all atoms are balanced and the charge on both sides of the equation arrow is +1.
Step 5: Highlight the acid–base conjugate pairs from the reaction in Step 1.	The acid–base conjugate pairs are: • $HI(aq)$ and $I^-(aq)$, where $I^-(aq)$ is the conjugate base • $H_2O(l)$ and $H_3O^+(aq)$, where $H_3O^+(aq)$ is the conjugate acid.

12.1 CHALLENGE

Constructing balanced ionic equations

Construct the balanced ionic equation for the following reactions in water:

1 $H_2PO_4^-(aq) + NH_3(aq) \rightleftharpoons$

2 $HCO_3^-(aq) + H_2S(aq) \rightleftharpoons$

3 $HSO_4^-(aq) + HCO_3^-(aq) \rightleftharpoons$

12.1 CHECK YOUR LEARNING

Describe and explain

1 Explain the difference between an acid and a base according to the Brønsted–Lowry definition.

2 Describe the relationship between a conjugate acid and conjugate base pair.

Apply, analyse and compare

3 Write the conjugate base for the following acids.

a HNO_2

b HSO_3^-

4 Write the conjugate acid for the following bases.

a $HCOO^-$

b IO_3^-

Design and discuss

5 Using balanced chemical equations, explain how an amphiprotic species behaves.

6 Construct the balanced full and ionic equations for the following reactions.

a $RbOH(aq) + H_2O(l) \rightarrow$

b $H_2CO_3(aq) + H_2O(l) \rightleftharpoons$

FIGURE 6 Water is an amphiprotic species because it can either donate a proton or accept a proton.

12.2 Strong and weak acids and bases

KEY IDEAS

In this topic, you will learn that:

+ the strength of an acid or a base depends on the amount of ionisation in solution, or the tendency to donate (or accept) hydrogen ions (H^+)
+ acid or base strength should not be confused with 'concentrated' or 'dilute'.

Acid and base strength

Acids and bases are often categorised by their strength and classified as either 'weak' or 'strong'. The strength of an acid or base depends on the:

1 extent to which ions dissociate in the presence of water

2 ability of the acid or base to donate or accept hydrogen ions (H^+).

A summary of the relative strengths of common acids and bases is shown in Figure 1.

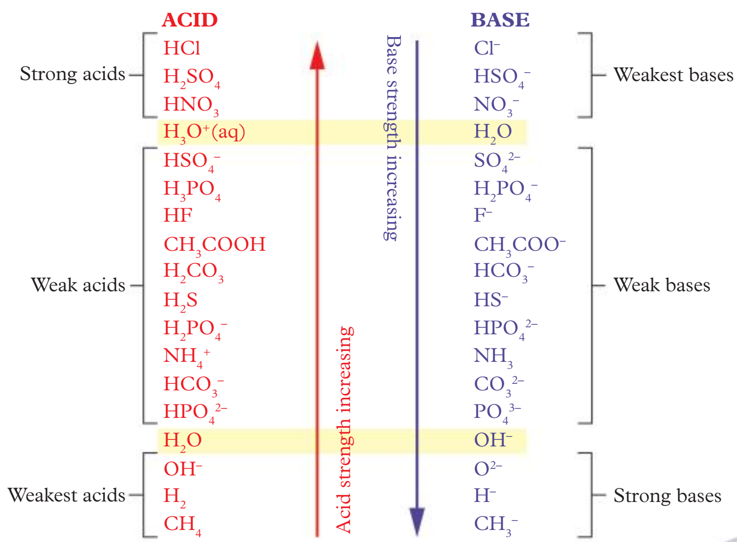


FIGURE 1 A summary of relative acid and base strengths

Strong and weak acids

strong acid
an acid that completely dissociates in water

A **strong acid** readily donates a proton(s) to a base and completely dissociates in water. For example, hydrochloric acid (HCl) is a strong acid. The ionisation of HCl is represented by a single arrow in the equation below; it can also be seen in Figure 6 on page 333.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Other strong acids that behave this way are nitric acid (HNO_3), sulfuric acid (H_2SO_4) and $HClO_4$ (perchloric acid).

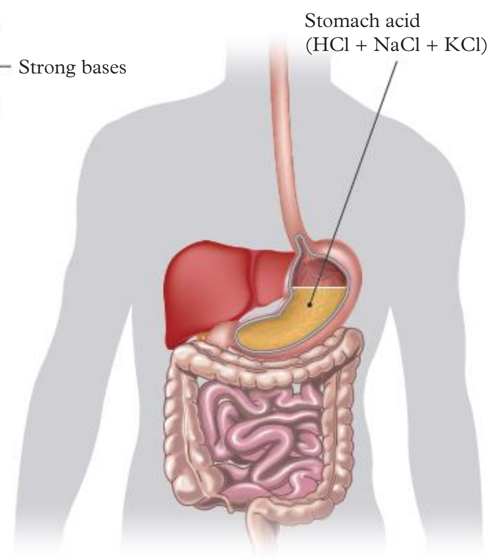


FIGURE 2 Stomach acid contains the strong acid HCl, which assists with the breakdown of food.

12.2
PRACTICAL: CONTROLLED EXPERIMENT

What happens when commercial indicators are added to different solutions? Go to page **516**.

weak acid
an acid that does not completely dissociate in water

A **weak acid** does not readily donate a proton(s) to a base and only partially dissociates in water. For example, hydrocyanic acid (HCN) is a weak acid. The incomplete ionisation of HCN in water is represented by the reversible equation shown below; it can also be seen in Figure 6.

$$HCN(aq) + H_2O(l) \rightleftharpoons CN^-(aq) + H_3O^+(aq)$$

Other weak acids are ethanoic (acetic) acid (CH_3COOH) and formic acid (HCOOH).

FIGURE 3 Vinegar contains the weak acid ethanoic acid CH_3CHOOH.

Strong and weak bases

strong base
a base that completely dissociates in water

A **strong base** readily accepts a proton(s) from an acid and completely dissociates in water. For example, sodium hydroxide (NaOH) is a strong base. The dissociation of NaOH is represented in the following equation; it can also be seen in Figure 6.

$$NaOH(aq) + H_2O\ (l) \rightarrow Na^+(aq) + OH^-(aq)$$

Other strong bases that behave in this way are the oxide ion (O^{2-}), and the ionic compounds sodium oxide (Na_2O) and potassium hydroxide (KOH).

FIGURE 4 The strong base sodium hydroxide (caustic soda) is frequently used in drain cleaners.

FIGURE 5 The weak base ammonia (NH_3) is a key ingredient in many fertilisers.

A **weak base** will not readily accept a proton(s) from an acid and only partially dissociates in water. For example, ammonia (NH_3) is a weak base. The incomplete dissociation of NH_3 is represented by the following reversible equation; it can also be seen in Figure 6.

weak base
a base that does not completely dissociate in water

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

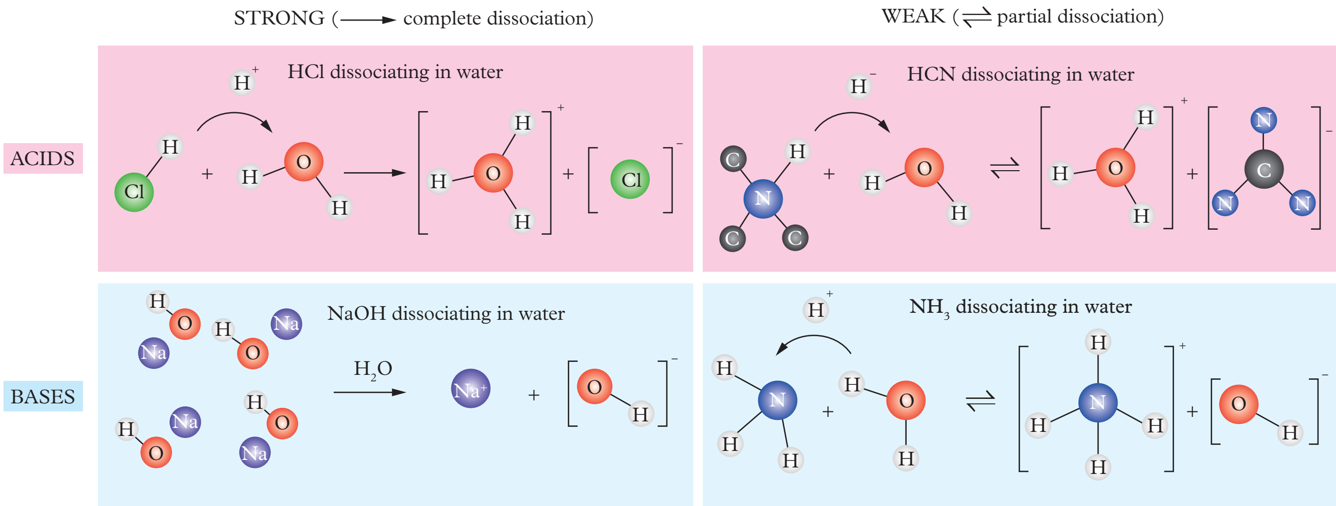


FIGURE 6 The differences between strong and weak acids and bases in solution: HCl is a strong acid (top left) and HCN is a weak acid (top right); NaOH is a strong base (bottom left) and NH_3 is a weak base (bottom right).

Acids and base concentrations

The **concentration** of a solution refers to the amount of a substance in a volume of solution. The formula to calculate the concentration of a substance is:

concentration
the amount of solute in a specific volume of a solution

$$\text{concentration} = \frac{\text{mass}}{\text{volume}}$$

You will learn more about calculating solution concentrations in Chapter 14.

Solutions that contain a large amount of a substance are referred to as **concentrated**. Solutions that contain a small amount of a substance are referred to as **dilute**. The difference between dilute and concentrated solutions is represented in Figure 7.

concentrated
a solution that contains many particles per unit of volume

dilute
a solution that contains few particles per unit of volume

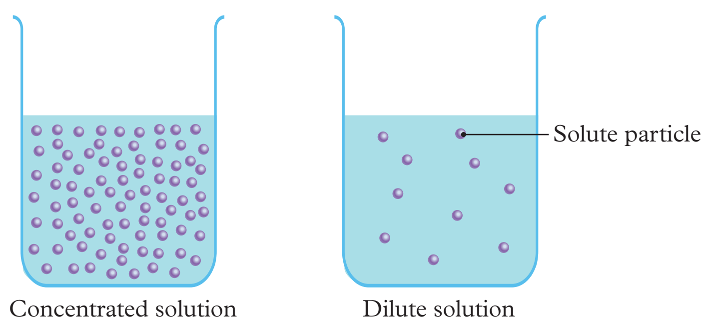


FIGURE 7 Concentrated solutions contain large amounts of a substance; dilute solutions contain small amounts of a substance.

Acid or base strengths should not be confused with the terms 'concentrated' and 'dilute'. Strength is a qualitative (descriptive) term, whereas concentration and dilution are quantitative terms that refer to the amount of acid or base in a given volume of solution. Concentrated, dilute, strong and weak acids are represented in Figure 8 on the next page.

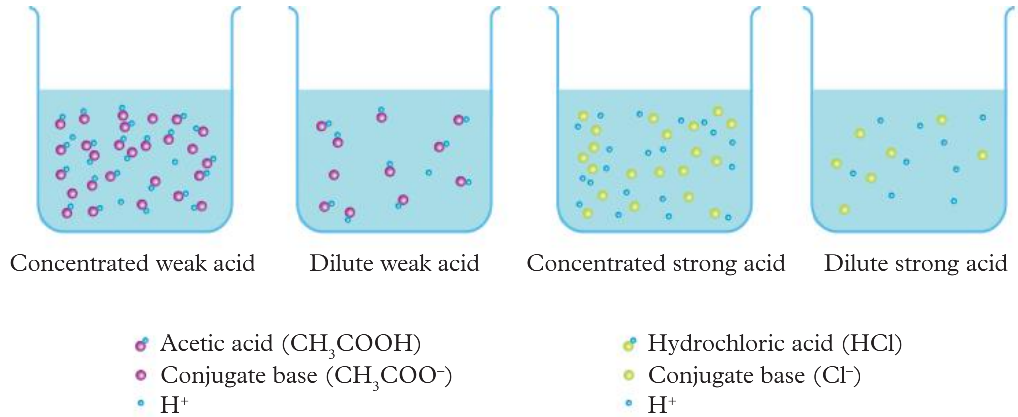


FIGURE 8 A representation of concentrated and dilute with the strong acid hydrochloric acid (HCl) and weak acid ethanoic acid (acetic acid) (CH_3COOH)

12.2 CHALLENGE

Calculating concentration

Figure 9 shows a solution in which the mass of the solute (sodium hydroxide) is 4.0 g and the volume of the solution is 1.0 dm^3. Calculate the concentration of sodium hydroxide in the solution.

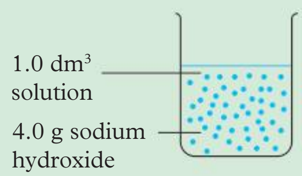


FIGURE 9 Sodium hydroxide in solution

Study tip

Strength is a qualitative or descriptive term. Concentration and dilution are quantitative terms that involve specific formulas (concentration $= \frac{\text{mass}}{\text{volume}}$ or $\frac{\text{number of moles}}{\text{volume}}$).

12.2 CHECK YOUR LEARNING

Describe and explain

1 Using Figure 1, identify a molecule that is considered both a weak acid and a weak base.

2 Explain the difference between a strong and a weak acid, using balanced equations to support your answer.

3 Explain the difference between a strong and a weak base, using balanced equations to support your answer.

Apply, analyse and compare

4 Contrast concentrated and dilute.

5 Contrast strength and concentration when describing acids and bases.

Design and discuss

6 Draw labelled diagrams to show the difference between:

a a concentrated strong base and a dilute strong base in aqueous solution

b a concentrated weak base and a dilute weak base in aqueous solution.

12.3 Neutralisation reactions to produce salts

KEY IDEAS

In this topic, you will learn that:

- acids react with hydroxides to produce salts and water
- acids react with carbonates to produce salts, water and carbon dioxide gas
- ingredients in antacids assist in the neutralisation of stomach acid.

Neutralisation reactions

When an acid reacts with water, it will produce hydrogen ions (H^+), but when a base reacts with water it produces hydroxide ions (OH^-). If an acid and base are mixed, H^+ formed from the acid will react with OH^- formed from the base. The H^+ and OH^- combine to produce water (H_2O) and dissociated ions combine to produce a metal salt as follows:

$$\text{acid} + \text{base} \rightarrow \text{water} + \text{metal salt}$$

This type of reaction, where an acid reacts with a base to produce water and a metal salt, is called **neutralisation**. One example is the reaction of HCl with NaOH in solution (Figure 1).

neutralisation the reaction of an acid with a base that produces a metal salt and water

$$\underbrace{HCl}_{\text{Acid}} + \underbrace{NaOH}_{\text{Base}} \longrightarrow \underbrace{H_2O}_{\text{Water}} + \underbrace{NaCl}_{\text{Metal salt}}$$

Dissociation: ionic equation

$$H^+ + Cl^- + Na^+ + OH^- \longrightarrow H_2O + Na^+ + Cl^-$$

FIGURE 1 In neutralisation reactions, the hydrogen ion and hydroxide ion combine to form water.

Acid reactions with metal hydroxides

When an acid reacts with a metal hydroxide, it will produce a metal salt and water. This reaction is considered a neutralisation because the products of the acid–base reaction are water and a metal salt.

$$\text{acid} + \text{metal hydroxide} \rightarrow \text{water} + \text{metal salt}$$

For example, when nitric acid reacts with sodium hydroxide, it produces the metal salt sodium nitrate and water:

$$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

The ionic equation for this reaction is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Acid reactions with metal carbonates

When an acid reacts with a carbonate, the products are a metal salt, water and carbon dioxide. Because an acid and a base react to form water and a metal salt, neutralisation occurs:

$$\text{acid} + \text{metal carbonate} \rightarrow \text{water} + \text{metal salt} + \text{carbon dioxide}$$

For example, when hydrochloric acid reacts with calcium carbonate, it produces the metal salt calcium chloride, water and carbon dioxide gas:

$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

The ionic equation for this reaction is:

$$2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$$

limewater
a saturated aqueous solution of calcium hydroxide

During this type of reaction, carbon dioxide gas can be collected and bubbled through **limewater**, as shown in Figure 2. The formation of a milky white precipitate indicates the presence of carbon dioxide gas. A second test is to hold a lit match or candle near the mouth of the reaction vessel. The match will extinguish in the presence of carbon dioxide gas.

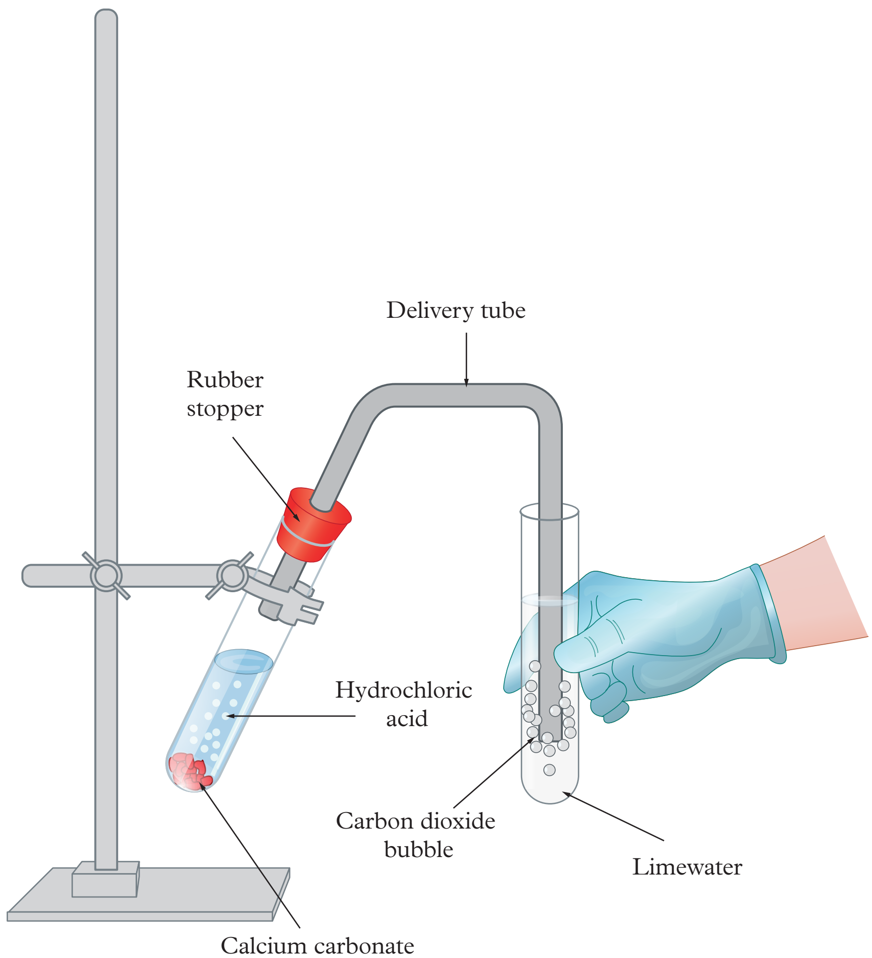


FIGURE 2 The experimental set-up to test for a metal carbonate with limewater

Antacid neutralisation of stomach acid

stomach acid
a digestive fluid consisting of salts and HCl that aids in the breakdown and digestion of food

Stomach acid is a digestive fluid in the stomach and is made up of salts such as NaCl and KCl as well as the strong acid HCl. The acidity of stomach acid aids in the breakdown of food, which is essential for nutrient absorption. However, if stomach acid leaches into the oesophagus, it can cause severe discomfort.

Between the stomach and oesophagus is a valve or sphincter that prevents the backflow of stomach contents and stomach acid. When the sphincter is weak or if pressure increases in the stomach, stomach acid can flow into the oesophagus and damage its sensitive lining and nerves. This causes heartburn or **acid reflux**, a digestive condition in which people experience a painful burning sensation in their stomach or chest. Acid reflux can be triggered by consuming spicy or acidic food or drink or by overeating.

acid reflux
a digestive problem caused by the liquid content of the stomach flowing up into the oesophagus

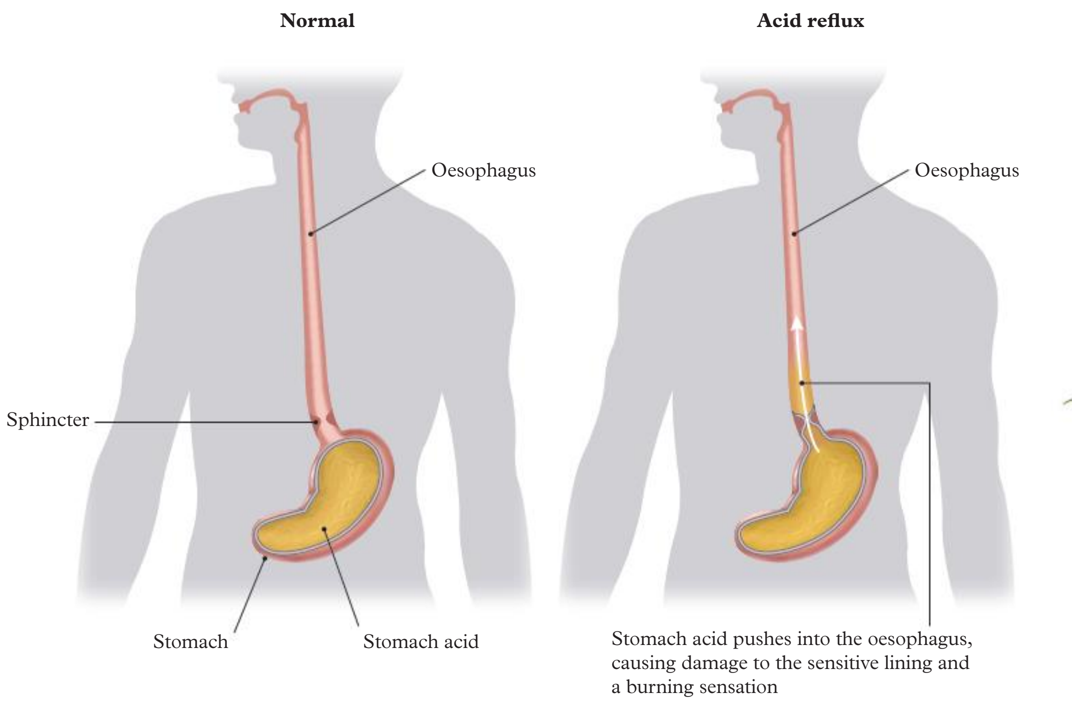


FIGURE 3 If stomach acid leaches into the oesophagus, it can cause severe discomfort known as acid reflux.

FIGURE 4 Eating spicy and acidic foods can trigger acid reflux.

If acid is detected by the sensory nerves in the lining of the oesophagus, a message is sent to the brain to produce saliva to the neutralise the stomach acid and protect the oesophagus.

The pain associated with heartburn can often be so intense that people seek medication for immediate relief. **Antacids** are substances taken to neutralise stomach acid. They are able to neutralise stomach acid because their active ingredient is a base. The neutralisation of stomach acid helps to reduce the symptoms of heartburn and relieves pain. This also helps the digestion of foods and beverages.

antacid
a substance that works to neutralise an acid

FIGURE 5 Various brands of antacids: **a** Gaviscon, **b** Eno, **c** Mylanta

Certain antacid brands will coat the lining of the oesophagus to protect the lining from stomach acid. Others act as a gel on the stomach's surface, which helps prevent acid leaking into the oesophagus and causing acid reflux.

Antacids are available in liquid form or as chewable tablets. They're sold under various brand names but contain common ingredients, including:

- magnesium carbonate ($MgCO_3$)
- magnesium trisilicate ($Mg_2O_8Si_3$)
- magnesium hydroxide $Mg(OH)_2$
- aluminium hydroxide ($Al(OH)_3$)
- calcium carbonate ($CaCO_3$)
- sodium bicarbonate ($NaHCO_3$).

An example of an antacid solution neutralising stomach acid is the reaction of aluminium hydroxide with hydrochloric acid, shown in the following equation:

$$Al(OH)_3(aq) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$$

The ionic equation for this reaction is:

$$Al(OH)_3(aq) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$$

Study tip

The general reactions of acids can be summarised as follows:
Acid + metal hydroxide → metal salt + water
Acid + metal carbonate → metal salt + water + carbon dioxide gas

12.3 SKILL DRILL

Evaluating the use of natural antacids

Key science skill: Analyse, evaluate and communicate scientific ideas

It is common to come across articles or ideas in the media that suggest the use of natural remedies to treat medical conditions.

Practise your skills

1 Research a media article that presents ideas on natural antacids to relieve heartburn. Select one suggested natural remedy and evaluate whether the ideas presented in the media article (on its effectiveness to relieve heartburn) are supported by scientific literature.

Need help analysing, evaluating and communicating scientific ideas? See Topics 1.9 and 1.10 (pages 28–31).

12.3 CHECK YOUR LEARNING

Describe and explain

1 Identify the products formed when an acid reacts with a:
  a metal hydroxide
  b metal carbonate.

2 Describe two ways to determine whether carbon dioxide has been produced in an acid reaction with a metal carbonate.

Apply, analyse and compare

3 Compare acid reactions with metal hydroxides to acid reactions with metal carbonates.

Design and discuss

4 Write the balanced chemical and ionic equations for the reaction of:
  a phosphoric acid with calcium hydroxide
  b magnesium carbonate with ethanoic acid
  c lithium hydrogen carbonate with sulfuric acid
  d ammonia with hydrochloric acid
  e nitric acid with barium hydroxide.

5 Write a balanced chemical equation for the reaction of hydrated magnesium trisilicate with stomach acid.

12.4 The pH scale

KEY IDEAS

In this topic, you will learn that:

+ the pH scale is used to compare the acidic, basic or neutral properties of substances in solution
+ the pH scale is a negative logarithmic scale and is a measure of the concentration of hydrogen ions (H^+) in solution: $pH = -\log_{10}[H^+]$
+ the ionic product of water ($K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 at 25°C) can be used to calculate the pH of strong acid and base solutions.

The pH scale

pH
a qualitive measure of H^+ concentration

logarithmic scale
a non-linear scale used to display numerical data over a wide range of values in a neat way

neutral
a solution which is neither acidic nor basic

alkaline
having a pH greater than 7

The **pH** scale is a negative **logarithmic scale** that provides a measure of the concentration of the hydrogen (H^+) or hydronium (H_3O^+) ions in solution. The pH scale can be used to compare the acidic, basic or **neutral** (neither basic nor acidic) properties of different substances in solution.

- If the pH is low, the solution is more acidic and the concentration of H^+ is high in solution.
- If the pH is high, the solution is more basic or '**alkaline**' and the concentration of H^+ is low in solution.

Figure 1 shows the scale of pH values with some everyday examples. When reading a pH scale, keep in mind that:

- acids have a low pH (<7)
- pure water has a neutral pH (7)
- bases have a high pH (>7).

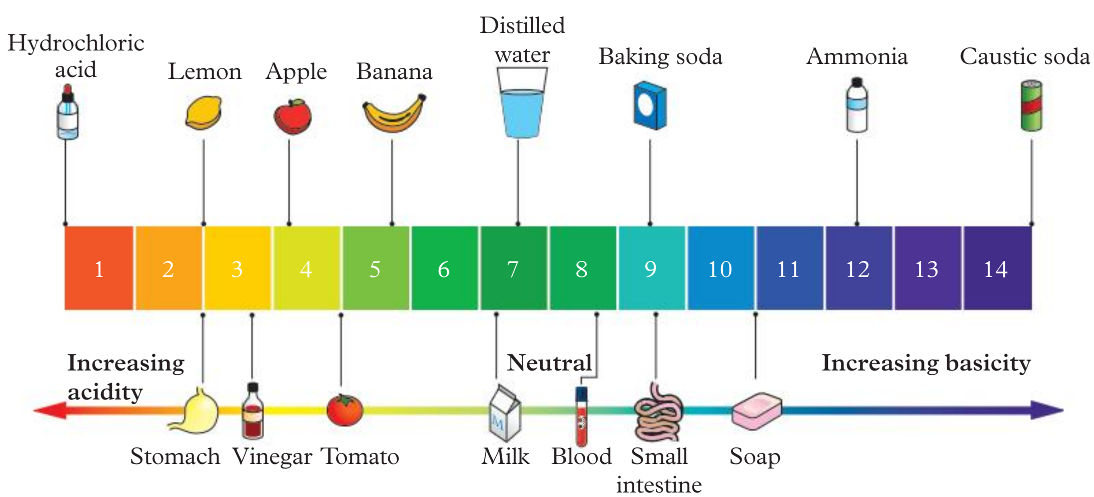


FIGURE 1 The pH scale with common examples

Calculating pH

The pH of a solution can be calculated by using the following mathematical expressions:

$$pH = -\log_{10}[H^+]$$

or

$$[H^+] = 10^{-pH}$$

In the formula above, $[H^+]$ is the molarity (in mol L^{-1}) of hydrogen or hydronium ions. Figure 2 shows the scale of pH values for acid and alkaline solutions, including the H^+ (and OH^- ion) concentrations for each value.

Calculating pOH

pOH
a qualitive measure of OH^- concentration

The scale that measures hydroxide ions is called **pOH**, and the values of pH and pOH are opposites. The pOH value is useful for changing alkalis and/or bases to a pH value, as pOH looks at the perspective from a base, and is defined by:

$$pOH = -\log_{10}[OH^-]$$

or

$$[OH^-] = 10^{-pOH}$$

Because pH and pOH are values on opposite ends of the scale, the following mathematical expression applies:

$$pH + pOH = 14$$

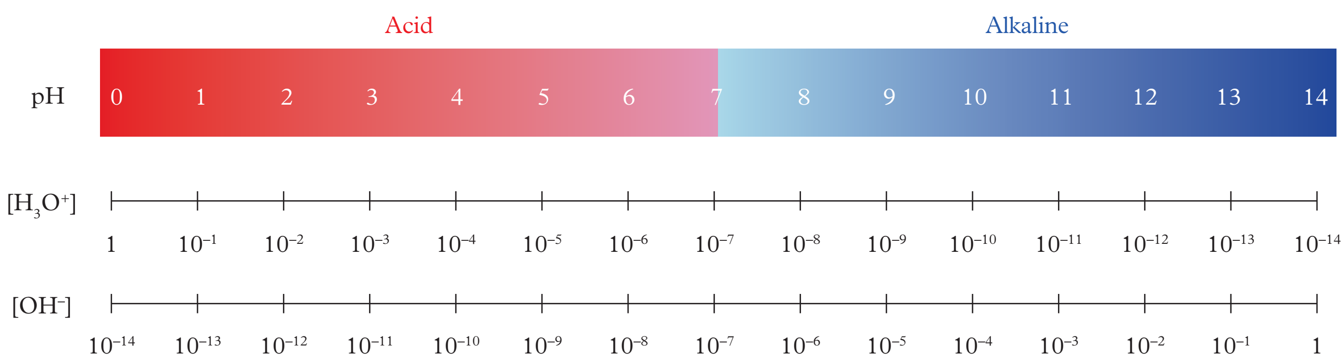


FIGURE 2 The scale of pH values for acid and alkaline solutions

The ionic product of water (K_w)

In Topic 12.1, you learnt that water is an amphiprotic substance and can function as both an acid and a base. Sometimes, a small number of water molecules in solution will dissociate into a hydrogen ion (H^+) and a hydroxide ion (OH^-). The hydrogen ion immediately attaches to a standard water molecule (H_2O) and forms the hydronium ion (H_3O^+). This process is called **self-ionisation** and is represented by the following equation:

self-ionisation
when water molecules interact with each other to form positive and negative ions

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

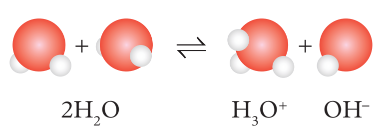


FIGURE 3 The self-ionisation of water

In distilled water, the number of hydronium ions is always equal to the number of hydroxide ions and the concentration of each is always 10^{-7} M at 25°C. The product of their molar concentrations is called the **ionic product of water** and is given the symbol K_w.

ionic product of water
the product of H_3O^+ and OH^- molar concentrations

This constant is equal to the product (or multiplication) of the concentrations of the hydronium ions and hydroxide ions in solution and is represented by the following equation:

$$K_w = [H_3O^+]\,[OH^-] = 1.00 \times 10^{-14}\ M^2 \text{ at } 25^\circ C$$

where M = molarity (mol L^{-1}) or ion concentration.

Acidic and basic solutions

In acidic solutions, H_3O^+ form from the self-ionisation of water or from acids reacting with water. This means that acidic solutions have a higher concentration of H_3O^+ than neutral and basic solutions. Because the concentrations of H_3O^+ and OH^- in solution must multiply to give K_w, in acidic solutions:

- $[H_3O^+] > 10^{-7}$ M
- $[OH^-] < 10^{-7}$ M

In basic solutions, the opposite occurs: OH^- form from the self-ionisation of water or from bases reacting with water. The means that basic solutions have a higher concentration of OH^- than neutral and acidic solutions. Because the concentrations of $[H_3O^+]$ and $[OH^-]$ in solution must multiply to give K_w, in basic solutions:

- $[OH^-] > 10^{-7}$ M
- $[H_3O^+] < 10^{-7}$ M.

12.4 CHALLENGE

Calculating pH and pOH

The OH^- ion concentration of a blood sample is 2.5×10^{-7} M.

1 Calculate the pH of the blood.
2 Explain whether your answer supports the approximate pH level for blood suggested in Figure 1.

FIGURE 4 Acidic solutions such as lemon juice have a higher concentration of H_3O+ ions than neutral and basic solutions.

Study tip

Make sure that you are familiar with the 'log' and 'inverse log' buttons on your scientific calculator in order to perform pH calculations. Remember to give your answer in the same number of significant figures as the numerical information stated in the question.

12.4 WORKED EXAMPLE

CALCULATING $[H^+]$, $[OH^-]$, pH AND pOH FOR A SOLUTION

Calculate $[H^+]$, $[OH^-]$, pH and pOH for a 0.030 M HCl solution.

Solution

Think	Do
Step 1: Write the equation for the ionisation of the acid solution and label the amount of H^+ present. In this case, 0.030 M hydrochloric acid produces 0.030 M H^+ because the molar ratio is 1:1 according to the ionisation equation coefficients.	$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$ $[HCl] = 0.030\ M = [H^+]$
Step 2: To work out $[OH^-]$, use the equation for the ionic product of water. Remember to rearrange the equation to make $[OH^-]$ the subject.	$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$ $= 0.030 \times [OH^-] = 1.00 \times 10^{-14}$ $[OH^-] = \frac{1.00 \times 10^{-14}}{0.03} = 3.3 \times 10^{-13}\ M$ (2 sig fig)
Step 3: Use the expression $pH = -\log_{10}[H^+]$ to get the pH of the solution.	$pH = -\log_{10}[0.030] = 1.5$ (2 sig fig)
Step 4: Use the expression $pOH = -\log_{10}[OH^-]$ or $pH + pOH = 14$ to get the pOH of the solution.	$pOH = -\log_{10}[OH^-]$ $= -\log_{10}[3.33 \times 10^{-13}]$ $= 12.5$ (2 sig fig) or $pOH = 14 - 1.52 = 12.5$ (2 sig fig)

12.4 CHECK YOUR LEARNING

Describe and explain

1 Explain the self-ionisation of pure water, using a balanced chemical equation to support your answer.

2 Using the pH scale in Figure 1, list two substances that are:

a acidic

b basic

c neutral or almost neutral.

Apply, analyse and compare

3 Compare pH and pOH.

4 The concentration of OH^- in a cleaning product was 10^{-3} mol L^{-1}. M. Find the pH of the solution.

5 A student used a pH meter to measure the pH of Coca-Cola® as 3.12. What is the concentration of the H^+ ions?

Design and discuss

6 Calculate the pH and pOH for a 0.20 M $Ba(OH)_2$ solution.

7 Calculate the pH and pOH for a 3.00×10^{-7} M solution of nitric acid.

12.5 Indicators and measuring pH

KEY IDEAS

In this topic, you will learn that:

+ indicators are chemicals that give a visible sign, such as a colour change, to determine whether a substance is acidic, basic or neutral
+ indicators can be manufactured synthetically or derived from natural ingredients
+ the accuracy and precision of natural and commercial indicators can vary
+ an electronic pH meter is the most accurate and precise way to measure pH.

indicator
a substance that undergoes a distinct observable change (often in colour) when pH conditions change

natural indicator
an indicator from natural sources such as plants

anthocyanin
a water-soluble pigment that changes colour in response to pH

Indicators are chemicals that give a visible sign such as a colour change when they react with specific substances. They are often used to determine whether substances are acidic, basic or neutral. Indicators can be derived from natural sources or manufactured synthetically in the laboratory.

Natural indicators

A **natural indicator** is an indicator that is produced naturally in plants and fruits. Many plants and fruits contain chemical **anthocyanins**, which are coloured pigments. These pigments are sensitive to pH and can change colour in response to different pH conditions.

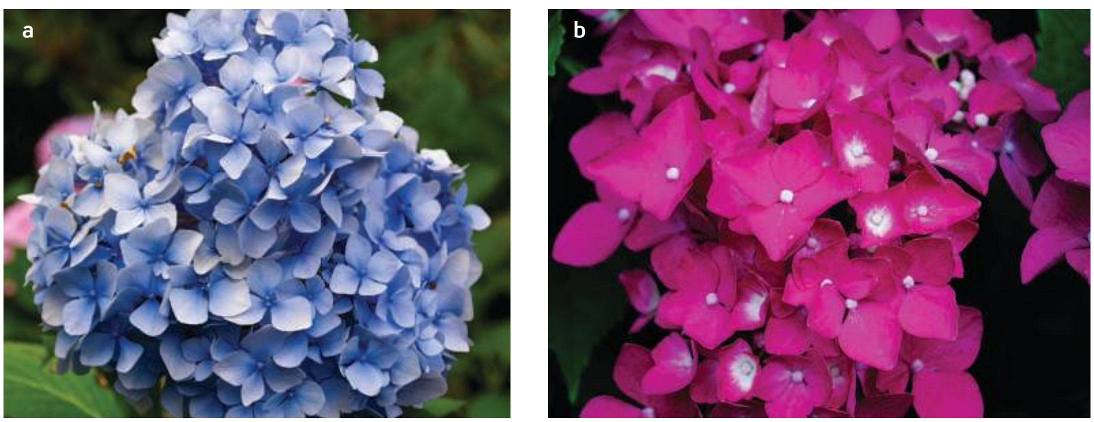


FIGURE 1 **a** Hydrangeas grown in acidic soil are blue. **b** Hydrangeas grown in neutral or basic soil are pink.

FIGURE 2 Vegetables and plants that contain the anthocyanin pigment

Anthocyanins are present in many vegetables and plants, including beetroots, red/purple cabbage, blueberries, red kidney beans and red apple skins. These pigments are used to produce dyes that can be used to make natural indicators for other qualitative measures of pH change.

Commercial indicators

There are also several synthesised commercial indicators available. These indicators are easy to source, store and use.

Litmus paper

litmus paper
treated paper that changes colour in acidic or basic solutions

One of the most common methods for testing whether a substance is acidic, basic or neutral involves using **litmus paper**.

- Acids turn blue litmus paper red, but do not change the colour of red litmus paper.
- Bases turn red litmus paper blue, but do not change the colour of blue litmus paper.
- Neutral solutions do not alter the colour of red or blue litmus paper.

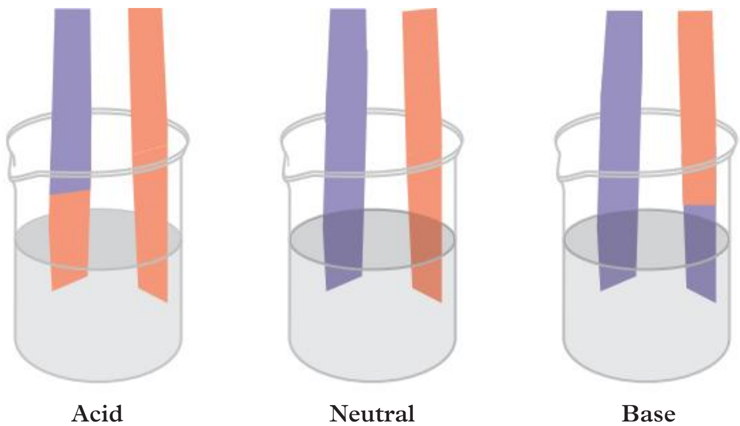


FIGURE 3 The litmus paper test in acidic, basic and neutral solutions

Litmus paper provides a qualitative measure of whether a substance is acidic, basic or neutral. These tests cannot determine the acidic or basic strength of the substance involved and have limited precision.

Universal indicator

universal indicator
a pH indicator consisting of several compounds that can present a range of colours in response to acidic, basic or neutral solutions

Universal indicator in solution or paper can be used to determine acid and base strength on the basis of colour change. Because its chemical composition includes a mixture of different dyes, universal indicator can undergo many colour changes (Figure 4).

FIGURE 4 The colours of universal indicator in basic, neutral and acidic solutions

Colour changes range from red (strong acid) to pink (weak acid), to yellow (neutral), to green (weak base), to blue (alkaline) to violet (strong base). Universal indicator is a more accurate way of measuring acid and base strength than using litmus paper, and, unlike litmus paper, it provides an indication of whether a substance is neutral. However, universal indicator still only provides a qualitative measure of acidity, basicity and neutrality.

FIGURE 5 The colours of universal indicator in paper form

Other commercial indicators

Other commercial indicators used for qualitative measures of pH are shown in Figure 6, along with the approximate ranges for each indicator and their respective colour changes.

Indicator	pH range
Methyl violet	0.15–3.2
Thymol blue	1.2–2.8 and 8.0–9.6
Methyl orange	3.2–4.4
Methyl red	4.8–6.0
Bromothymol blue	6.0–7.6
Phenolphthalein	8.2–10.0
Alizarin yellow	10.2–12.0

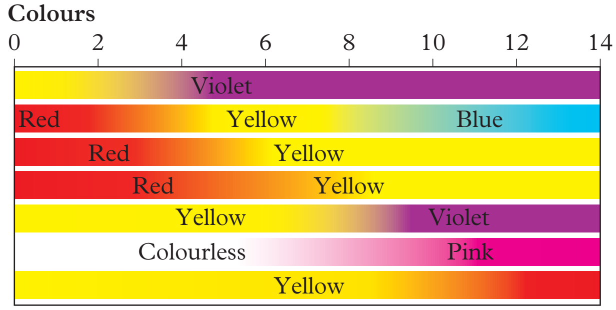


FIGURE 6 The pH ranges of some common commercial indicators

pH meter

A **pH meter** is an electronic device used to measure pH. The probe on the pH meter measures the voltage of the electrolyte (the extent of ionisation or H^+ activity) and converts the voltage to a quantitative pH reading to two decimal places (Figure 7). Using a pH meter is the most accurate and precise way to measure the pH of a substance in solution.

pH meter
an electronic device that measures hydrogen ion activity (acidity or alkalinity) in solutions

FIGURE 7 An electronic pH meter with buffer solutions for calibration

Study tip

Observation of pH through indicator colour change is a qualitative measure. A pH meter, which converts voltage of H^+ activity in an electrolyte to pH, provides a quantitative value that is more accurate and precise.

The accuracy and precision of the pH meter is maintained by regular calibration with buffer solutions. A wide variety of industries use pH meters for quality control in the manufacture and supply of goods, including the food and beverage, pharmaceutical, fuel, agriculture and water treatment industries.

FIGURE 8 An electronic pH meter being used to measure the pH of soil

12.5 REAL-WORLD CHEMISTRY

pH analysis in wine

Measuring pH accurately is essential in wine production. Most wines have a pH of between 3 and 4. Generally, winemakers want the initial grape juice used to make wine to have pH < 3.5. This is because tartrate precipitation (which starts at pH 3.8) will further increase the pH of wine, which can be detrimental to the wine.

FIGURE 9 Testing the pH of wine

Wine production relies on microbes for fermentation. Many microbes cannot survive at pH < 3.5. However, as pH rises to 4.0, many problematic microbes can survive and possibly ruin a wine.

Apply your understanding

1 The pH of most wines ranges between 3 and 4. Does this statement imply that wine is considered to be acidic, basic or neutral?

2 Give reasons why the pH of wine during production needs to be 3.5–4.0.

3 Explain why measuring the pH of wine with litmus paper or universal indicator would be unreliable.

12.5 CHECK YOUR LEARNING

Describe and explain

1 Figure 10 shows a strip of red litmus paper placed in a solution. Identify whether the solution is acidic, basic or neutral.

FIGURE 10 Red litmus paper in a solution

2 Describe how anthocyanins work as a natural indicator.

Apply, analyse and compare

3 Compare the accuracy and precision of a pH meter and natural or commercial indicators in the measurement of pH values.

4 Using Figure 6 on the previous page, which commercial indicator(s) would be suitable to use for pH changes between:

a pH 1 and 4

b pH 7 and 9

c pH 10 and 13.

Design and discuss

5 Design an experiment that uses indicators to compare the pH levels of the apple juice, orange juice, green tea, black tea, tap water and sports drink.

12.6 Acid–base reactions in society

KEY IDEAS

In this topic, you will learn that:

+ dissolved carbon dioxide contributes to the natural acidity of rain
+ acid rain contains high levels of dissolved sulfur dioxide and nitrous oxides, from the burning of fossil fuels
+ carbon dioxide dissolved in the oceans forms a weak acid that is harmful for shell growth in marine invertebrates.

carbon–oxygen cycle
the cycle in which atmospheric oxygen is converted to carbon dioxide in animal respiration and regenerated by plants during photosynthesis

Carbon dioxide and oxygen are cycled in nature through the **carbon–oxygen cycle** by the processes of photosynthesis and respiration (Figure 1):

Photosynthesis:

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{sunlight and chlorophyll}} C_6H_{12}O_6(aq) + 6O_2(g)$$

Respiration:

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$$

Other factors that contribute to the carbon–oxygen cycle include the:

- decomposition of dead organisms, a slow process that adds CO_2 to the cycle
- burning of fossil fuels and active volcanoes, which adds CO_2 to the cycle
- formation of carbonate ions as atmospheric CO_2 dissolves in the ocean:

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

At present, the largest contributor to increasing levels of carbon dioxide in the atmosphere is the rate and extent to which we burn fossil fuels for energy.

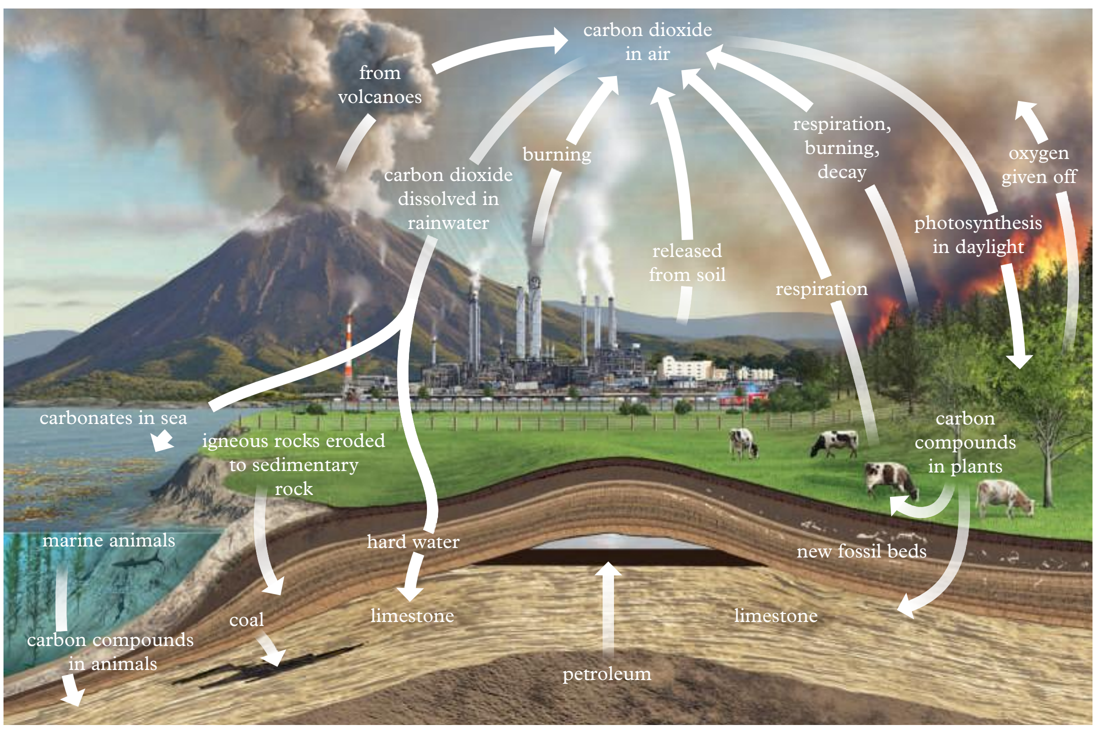


FIGURE 1 The carbon–oxygen cycle

Natural acidity of rain

Normal rain is slightly acidic at a pH of 5.6. This is because carbon dioxide in the atmosphere reacts readily with water vapour to form carbonic acid:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

This reaction contributes to the **natural acidity** of rain, snow and hail.

natural acidity
the slight acidity of rain caused by the reaction of CO_2 with water

acid rain
rain with higher-than-normal acidity because of the atmospheric reactions of water with sulfur dioxide (SO_2) or nitric oxide (NO)

Acid rain

Acid rain, by definition, is any rain, snow or hail hat contains high levels of nitric and sulfuric acids. Acid rain usually has a pH between 4.2 and 4.4. Both nitric oxide and sulfur dioxide are released into the atmosphere by the burning of fossil fuels and other industrial activity. Acid rain is formed when atmospheric sulfur dioxide (SO_2) or nitric oxide (NO) react with oxygen and water vapour.

Sulfur dioxide reacts with oxygen in the atmosphere and produces sulfur trioxide. Sulfur trioxide then reacts with water vapour to form sulfuric acid (H_2SO_4) as shown by the following reactions:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Nitric oxide reacts with oxygen in the atmosphere to form nitrogen dioxide. Nitrogen dioxide then reacts with water vapour in the atmosphere to produce nitrous acid (HNO_2) and nitric acid (HNO_3) as shown by the following reactions:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

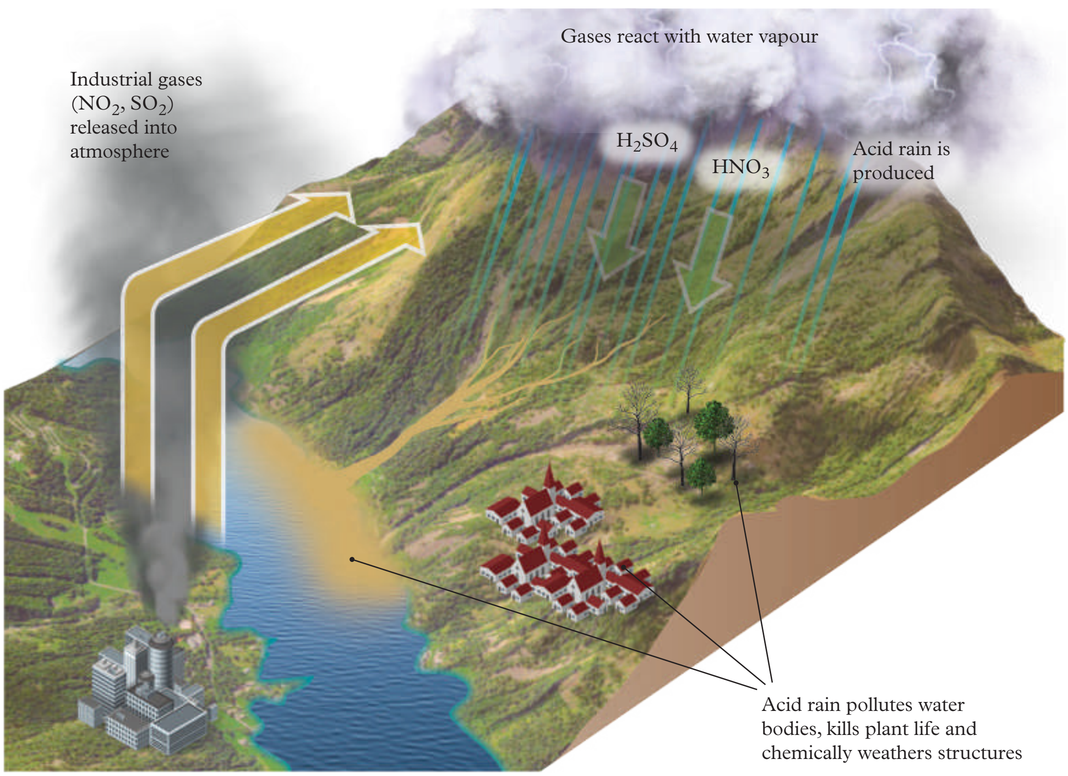


FIGURE 2 The formation of acid rain

Increased instances of acid rain can trigger respiratory issues in animals and humans. Also, when acid rain falls on bodies of water and landscapes, it can alter the pH and acidity levels of the water and soil. Changes to pH can affect the survival of organisms in these systems that have specific pH tolerance levels.

FIGURE 3 The natural acidity of rain can lead to the chemical weathering of objects such as this statue.

Ocean acidification

Ocean acidification is the decrease of ocean pH. Increased levels of carbon dioxide in the atmosphere have directly accelerated the rate of ocean acidification. This is because carbon dioxide gas is soluble in water. CO_2 enters the ocean by the following reactions:

$$CO_2(g) \rightleftharpoons CO_2(aq)$$

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

As the concentration of carbonic acid (H_2CO_3) in oceans rises, pH lowers and oceans become more acidic.

ocean acidification
the continual decrease of ocean pH caused by increased atmospheric carbon dioxide levels

Increased carbonic acid levels are also particularly harmful for shell-building animals and corals that have exteriors or structures made from calcium carbonate ($CaCO_3$). Many marine animals build their shells from free carbonate and calcium ions available in seawater as follows:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)$$

Increased carbonic acid concentrations in the ocean limit the ability for the above reaction to occur. This is because in seawater, carbonic acid will dissociate into H^+ and bicarbonate ions (HCO_3^-) as follows:

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

In the presence of H^+ ions, free carbonate ions in the ocean tend to bond with dissociated H^+ ions to form bicarbonate before bonding with free calcium ions (Ca^{2+}) to form calcium carbonate:

$$CO_3^{2-}(aq) + H^+(aq) \rightleftharpoons HCO_3^-(aq)$$

The formation of bicarbonate prevents marine organisms from using free carbonate ions to grow new shells. Low calcium carbonate levels in the skeletons or structures of marine organisms reduces their ability to grow and reduces the strength of their defensive structures, leaving them more vulnerable to predators.

Study tip

Dissolved carbon dioxide is responsible for the natural acidity of rain, which is different from acid rain, which is formed from dissolved sulfur dioxide and nitrous oxides, released from burning fossil fuels

FIGURE 4 Shell-building marine organisms rely on free carbonate ions present in the ocean to grow new shells.

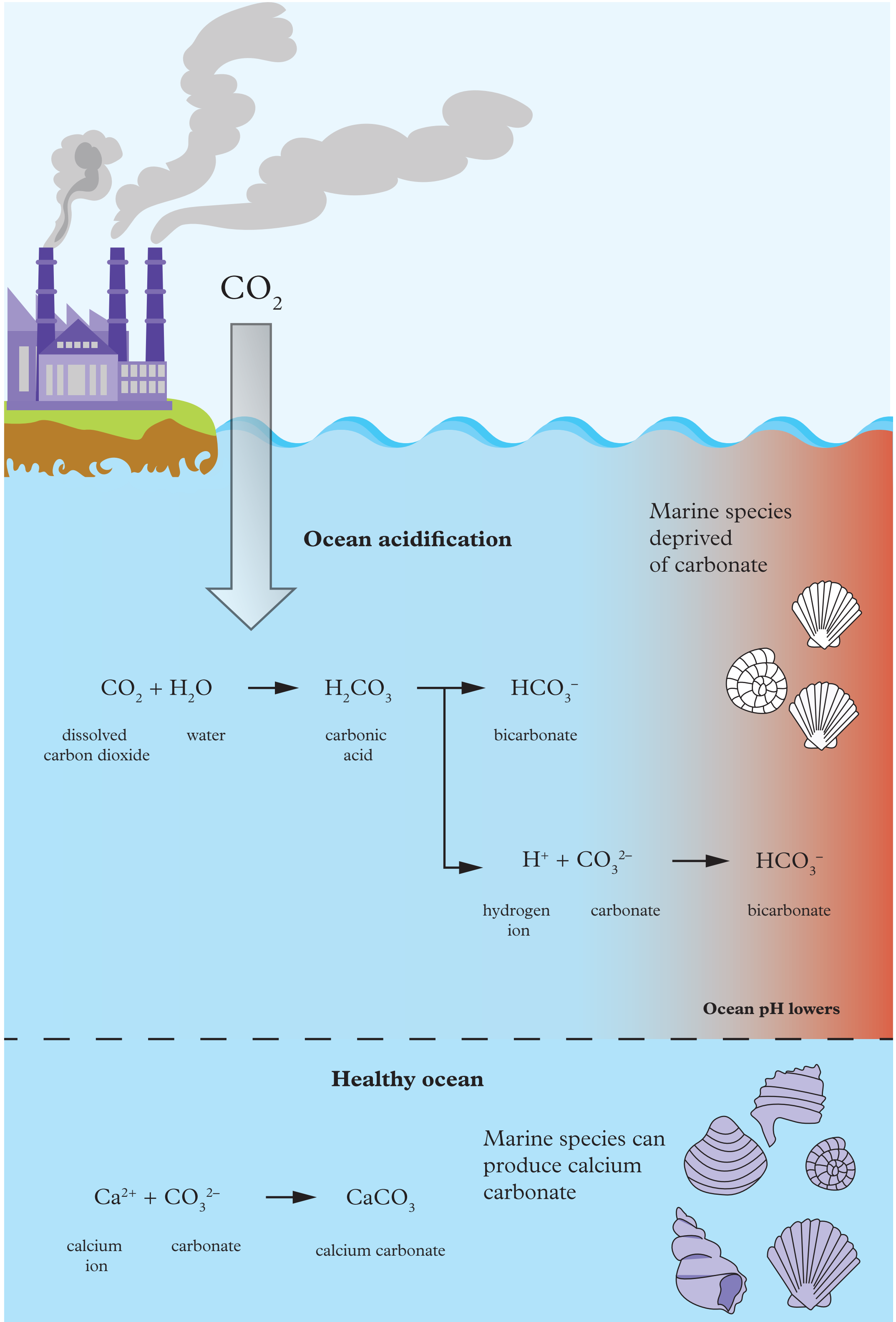


FIGURE 5 A summary of ocean acidification

12.6 SKILL DRILL

Analysing pH data

Key science skill: Analyse and evaluate data and investigation methods

Figure 6 shows a graph of CO_2 levels off the coast of Hawaii and the effect on the pH of seawater.

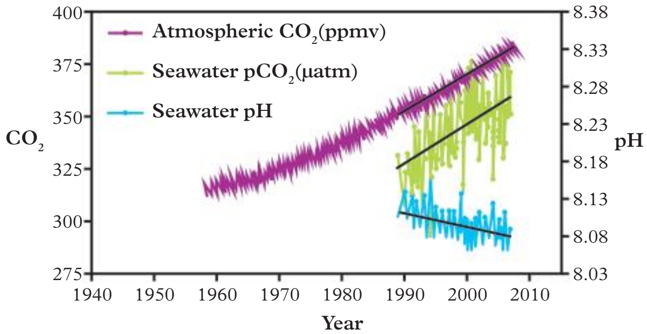


FIGURE 6 The impact of CO_2 levels off the coast of Hawaii

Practise your skills

1 Identify the relationship between the measured levels of CO_2 in the atmosphere over time. Use evidence from Figure 6 to support your answer.

2 Identify the relationship between the measured levels of CO_2 in seawater over time. Use evidence from Figure 6 to support your answer.

3 Using your answers from Questions 1 and 2, identify the relationship between the measured levels of CO_2 in the atmosphere and seawater and the pH of seawater. Use evidence from Figure 6 to support your answer.

Need help analysing data? See Topic 1.8 (page 24).

12.6 CHECK YOUR LEARNING

Describe and explain

1 Define ‘ocean acidification’.

2 Explain how increasing ocean acidity can prevent marine shell-building animals and corals from building their protective skeletons or shells.

3 Describe how carbon dioxide is added and removed from at least two processes in the carbon–oxygen cycle.

Apply, analyse and compare

4 Compare the acidity of normal rain and of acid rain.

Design and discuss

5 Discuss which United Nations Sustainability Goal(s) are most relevant to the problem of ocean acidification.

FIGURE 7 Ocean acidification deprives marine species of carbonate.

CHAPTER

12 Review

Chapter summary

12.1

- The Brønsted–Lowry theory defines acids as hydrogen ion (H^+) or proton donors and bases as hydrogen ion (H^+) or proton acceptors, forming conjugate acid–base pairs.
- Polyprotic acids can donate more than one proton from each molecule.
- Amphoteric substances can act as acids or bases in reactions.
- Acid–base reactions in water can be represented by balanced ionic and full chemical equations, including states.

12.2

- The strength of an acid or a base depends on the amount of ionisation in solution, or the tendency to donate (or accept) hydrogen ions (H^+).
- Acid or base strength should not be confused with concentration and dilution.

12.3

- Acids react with metal hydroxides to produce a salt and water.
- Acids react with metal carbonates to produce a metal salt, water and carbon dioxide gas.
- Ingredients present in antacids assist in the neutralisation of stomach acid.

12.4

- The pH scale is used to compare the acidic, basic or neutral properties of substances in solution.
- The pH scale is a negative logarithmic scale and is a measure of the concentration of hydrogen ions (H^+) in solution; $pH = -\log_{10}[H^+]$.
- The ionic product of water is a constant: $K_w = [H_3O^+]\ [OH^-] = 1.00 \times 10^{-14}\ M^2$ at 25°C.
- The ionic product of water can be used to calculate the pH of strong acid and base solutions.

12.5

- Indicators are chemicals that give a visible sign, such as a colour change, to show whether a substance is acidic, basic or neutral.
- Indicators can be manufactured synthetically or derived from natural ingredients.
- The accuracy and precision of natural and commercial indicators can vary.
- An electronic pH meter can be used for more accurate and precise pH measurement.

12.6

- Dissolved carbon dioxide contributes to the natural acidity of rain.
- Acid rain contains high levels of dissolved sulfur dioxide and nitrous oxides, released from the burning of fossil fuels.
- Dissolved carbon dioxide in the oceans forms a weak acid, which is harmful for shell growth in marine invertebrates.

Key formulas

Ionic product of water	$K_w = [H_3O^+]\ [OH^-] = 1.00 \times 10^{-14}\ M^2$
pH	$pH = -\log_{10}[H^+]$
pOH	$pOH = -\log_{10}[OH^-]$
pH and pOH	$pH + pOH = 14$
Acid reaction with metal carbonate	acid + metal carbonate → water + metal salt + carbon dioxide
Acid reaction with metal hydroxide	acid + metal hydroxide → water + metal salt

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can …	Confidently	Mostly	Not really	Revision link
• describe the Brønsted–Lowry theory of acids and bases, including polyprotic acids and amphiprotic species, and the writing of balanced ionic and full equations, with states, for their reactions in water	☐	☐	☐	Go back to Topic 12.1
• distinguish between strong and weak acids and strong and weak bases, and between concentrated and dilute acids and bases, including common examples	☐	☐	☐	Go back to Topic 12.2
• explain how neutralisation reactions produce salts, including: reactions of acids with metal carbonates and metal hydroxides, including balanced full and ionic equations, with states, and types of antacids and their use in the neutralisation of stomach acid	☐	☐	☐	Go back to Topic 12.3
• use the logarithmic pH scale to rank solutions from most acidic to most basic	☐	☐	☐	Go back to Topic 12.4

Revision questions

Multiple choice

1 For $HS^-(aq)$:

A $H_2S(aq)$ is the strong conjugate acid.

B $H_2S(aq)$ is the weak conjugate acid.

C $H_2S(aq)$ is the strong conjugate base.

D $H_2S(aq)$ is the weak conjugate base.

2 Consider the following reactions:

$HSO_4^-(aq) + NH_3(aq) \rightarrow SO_4^{2-}(aq) + NH_4^+(aq)$

$H_2PO_4^-(aq) + NH_3(aq) \rightarrow HPO_4^{2-}(aq) + NH_4^+(aq)$

$HSO_4^-(aq) + NH_4^+(aq) \rightarrow H_2SO_4(aq) + NH_3(aq)$

$H_3PO_4(aq) + CN^-(aq) \rightarrow H_2PO_4^-(aq) + HCN(aq)$

In the above reactions, which of the following is acting as an amphiprotic substance?

A $H_2PO_4^-(aq)$

B $NH_4^+(aq)$

C $NH_3(aq)$

D $HSO_4^-(aq)$

3 A 4.0 M solution of CH_3COOH is a:

A dilute strong acid.

B dilute weak acid.

C concentrated strong acid.

D concentrated weak acid.

4 Which of the following solutions is *most* acidic?

A Hydrogen chloride in water, $[H^+] = 0.001$ M

B Potassium hydroxide in water, $[OH^-] = 10^{-5}$ M

C Ethanoic acid in water, $[H^+] = 0.000\ 05$ M

D Sodium hydroxide in water in water, $[OH^-] = 0.001$ M

5 The pH of solution X is 8 and the pH of solution Y is 10. Which statement is correct about the hydrogen ion concentration in the two solutions?

A $[H^+]$ in X is 10 times more than in Y.

B $[H^+]$ in X is 100 times more than in Y.

C $[H^+]$ in X is 8 times more than in Y.

D $[H^+]$ in X is 0.1 times more than in Y.

6 Which of the following is *not* a polyprotic acid?

  A H_2CO_3

  B H_3PO_4

  C H_2S

  D CH_3COOH

7 For the ionisation of perchloric acid in water:

  $HClO_4(l) + H_2O(l) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$

  which of the following species would have the lowest concentration?

  A $HClO_4$

  B H_2O

  C H_3O^+

  D ClO_4^-

8 Which of the following is *not* true of acids?

  A They taste sour.

  B They react with magnesium to form hydrogen gas.

  C They contain more H_3O^+ ions than OH^- ions.

  D They turn phenolphthalein indicator pink.

9 What are the products of a reaction between H_2SO_4 and $KHCO_3$?

  A H_2O and CO_2 and KSO_4

  B H_2 and CO_2 and K_2SO_4

  C H_2CO_3 and K_2SO_4

  D H_2O and CO_2 and K_2SO_4

10 The Brønsted–Lowry definition of a base is:

  A an electron acceptor.

  B a substance capable of forming OH^- ions in water.

  C a species that forms water when it reacts with an acid.

  D a proton acceptor.

Short answer

11 Construct balanced chemical equations, including states to show that in water:

  a chromic acid (H_2CrO_4) is a diprotic acid

  b PH_3 is a weak base

  c HS^- is amphiprotic.

12 In aqueous solution, HBr dissociates completely, whereas HOI ionises partially.

  a Construct balanced chemical equations showing these two effects.

  b For each equation, identify the conjugate acid–base pairs.

  c Use your equations to explain the difference between ionisation and dissociation.

13 Formic acid (HCOOH) (from the Latin *formica* = an ant) is a weak monoprotic acid in many insect stings and is partially responsible for the pain they cause. A particular formic acid solution has pH 4.0.

  a **i** State the hydrogen ion concentration of the formic acid solution.

  ii Hence calculate the concentration of the hydroxide ion in the formic acid solution.

  b Write an equation to show formic acid acting as a weak monoprotic acid in water.

14 What is $[H_3O^+]$ and $[OH^-]$ of a solution with pH 8.05? Is this an acidic, a basic or a neutral solution?

15 The pH of rainwater collected in a certain region of Melbourne on a particular day was 4.82. What is the H^+ ion concentration of the rainwater? What is the OH^- ion concentration of the rainwater?

Apply, analyse and compare

16 The pH of a solution of HCl is 2.4. Calculate the concentration of H^+ ions.

17 In a 2 M solution of NaOH, calculate the:

  a concentration of OH^- ions

  b concentration of H^+ ions

  c pH.

18 Determine the acid–base conjugate pairs for the following reactions.

a $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

b $H_2O(aq) + H_2SO_4(aq) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$

c $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

19 Write the balanced full and ionic equations for the reaction of:

a solid magnesium carbonate ($MgCO_3$) and sulfuric acid (H_2SO_4) solution

b lithium hydroxide (LiOH) and ethanoic acid (CH_3COOH) solution

c hydrochloric acid (HCl) and zinc metal (Zn).

20 A chemist dilutes concentrated hydrochloric acid to make two solutions: solution A is 3.0 M and solution B is 0.0024 M. Calculate the $[H_3O^+]$ and $[OH^-]$ of the two solutions at 25°C.

Design and discuss

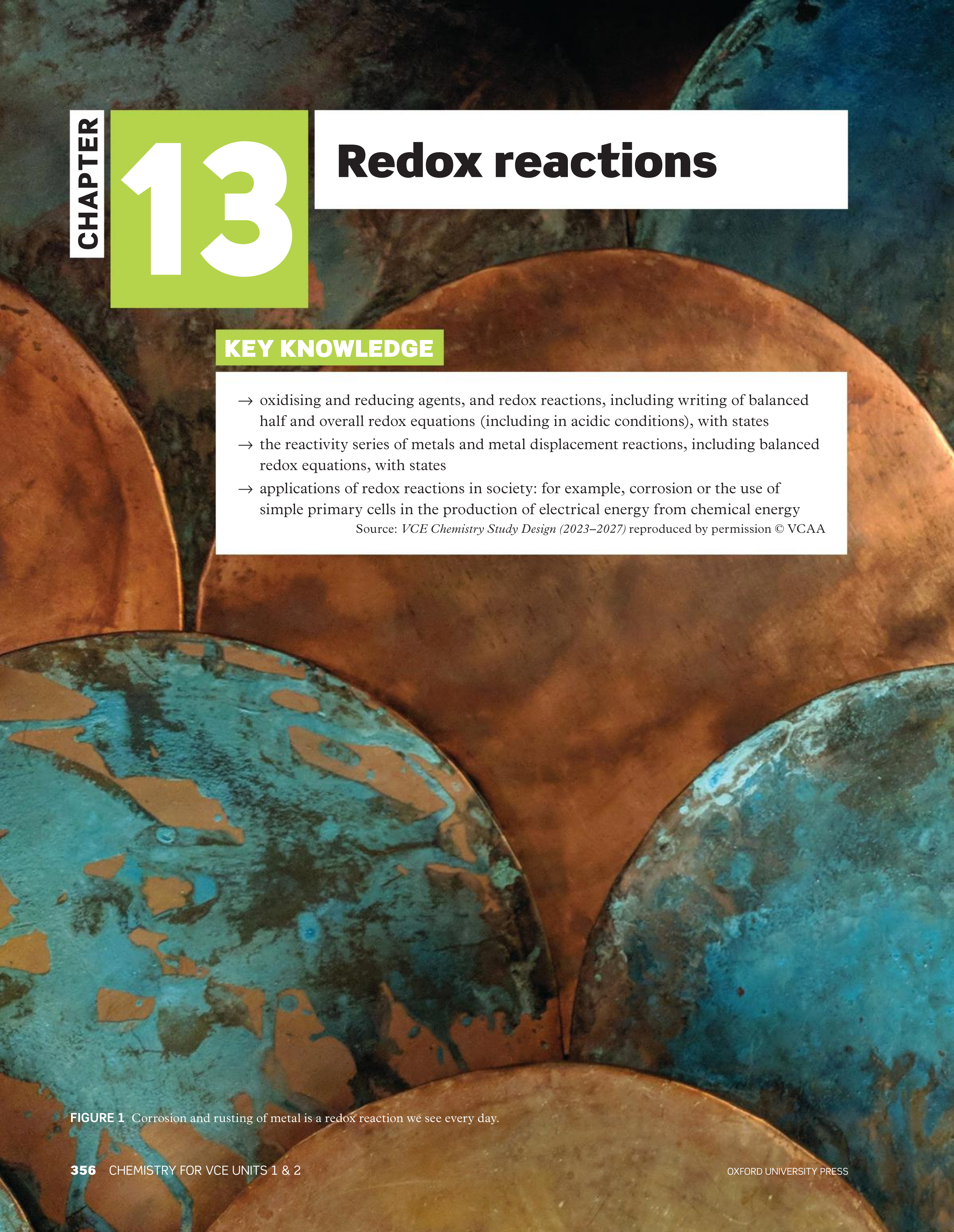
21 Your class has been lent a calibrated electronic pH meter by the local university.

Describe how you would design an experiment, using this apparatus, to measure the pH of household substances and classify them as acids, bases or neutral. Include an aim, a hypothesis, an equipment/materials list, a procedure, your results and how you would communicate your results.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test you knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.



CHAPTER 13 Redox reactions

KEY KNOWLEDGE

- → oxidising and reducing agents, and redox reactions, including writing of balanced half and overall redox equations (including in acidic conditions), with states
- → the reactivity series of metals and metal displacement reactions, including balanced redox equations, with states
- → applications of redox reactions in society: for example, corrosion or the use of simple primary cells in the production of electrical energy from chemical energy

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FIGURE 1 Corrosion and rusting of metal is a redox reaction we see every day.

GROUNDWORK

In Chapter 13, you will learn about redox reactions, including how to identify them, write balanced half-equations and overall redox equations, and predict whether redox reactions will occur. You will also learn about the applications of redox reactions in society.

This chapter will build on concepts you have already learnt in Chapters 2, 5 and 11. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

13A Describe the trend in reactivity of metal atoms in the periodic table.

13A Groundwork resource
Reactivity of metals

13B What type of bond is formed through the transfer of electrons?

13B Groundwork resource
Ionic bonding

13C Balance the following equation, including states:

$$Na_2CO_3 + Ca(NO_3)_2 \rightarrow NaNO_3 + CaCO_3$$

13C Groundwork resource
Balancing equations

PRACTICALS

13.3 PRACTICAL: LITERATURE REVIEW	Are redox reactions beneficial or harmful to society and the environment?	Page 518

13.1 Oxidising and reducing agents

KEY IDEAS

In this topic, you will learn that:

+ a redox reaction involves the transfer of electrons from one species to another
+ one chemical species loses electrons in a process called oxidation while the other gains electrons in a process called reduction
+ half-equations can be written for oxidation and reduction reactions to show the gain or loss of electrons – this can be combined to form an overall redox equation
+ complex redox reactions can be conducted under acidic conditions
+ complex half-equations and complex overall redox equations are generated using the KOHES method.

Some chemical reactions involve the transfer of valence electrons between reactants to form products. This exchange of electrons is like a transaction – one reactant loses one or more electrons, and the other reactant accepts them.

'**Redox**' is an abbreviation for a pair of reactions where the exchange of electrons between reactants occurs simultaneously. These reactions are called **reduction** and **oxidation** reactions. One cannot occur without the other.

redox
a chemical reaction involving the transfer of electrons from one reactant to another

reduction
a process in which one chemical species gains electrons from another

oxidation
a process in which one chemical species loses electrons to another

chemical species
the general name given to an atom, element, ion, molecule or compound

Study tip

A way to remember that oxidation is a loss of electrons and reduction is a gain in electrons is the abbreviation OIL RIG: **O**xidation **I**s **L**oss, **R**eduction **I**s **G**ain.

Redox reactions

Reduction and oxidation reactions involve the transfer of electrons. To form a new product, one reactant loses one or more electrons and the second reactant gains them.

- Oxidation occurs when a **chemical species** loses one or more electrons.
- Reduction occurs when a chemical species gains one or more electrons.

The best way to understand this is through an example. Let's look at a redox reaction between sodium and oxygen to form sodium oxide. This is given by the equation:

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

The electron shell diagrams of the sodium and oxygen atoms (Figure 1) show that sodium has one valence electron, which it must lose to have a complete valence shell and be stable. However, oxygen has six valence electrons. This means it must gain an extra two to complete its octet, have a complete valence shell and be stable. Since sodium will lose one electron and oxygen will gain two, we need two sodium atoms for every oxygen atom.

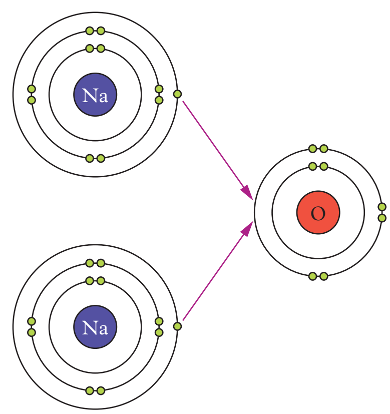

FIGURE 1 Electrons are transferred between sodium and oxygen atoms to form sodium oxide.

Because of this transfer of electrons, Na becomes the cation Na^+ and O becomes the anion O^{2-}. This means that sodium has lost an electron and gained a positive charge. Oxygen has gained two electrons and a negative charge. Therefore, sodium has undergone oxidation and oxygen has undergone reduction. In other words, sodium has been oxidised and oxygen has been reduced. The general process is summarised in Figure 2.

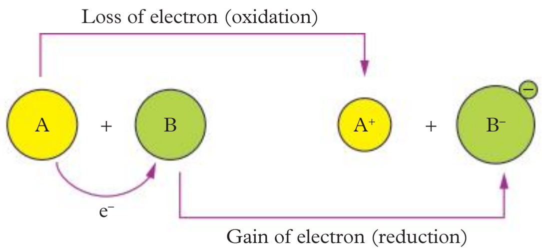

FIGURE 2 The transfer of an electron from A to B results in a positive cation (A^+) and negative anion (B^-), which form an ionic compound (AB).

Oxidising and reducing agents

A chemical species cannot lose electrons unless another is available to accept them. This means that the redox reaction must occur between an **oxidising agent** (oxidant) and a **reducing agent** (reductant).

An oxidising agent or oxidant:

- is responsible for causing oxidation
- is itself reduced in the redox reaction.

A reducing agent or reductant:

- is responsible for causing reduction
- is itself oxidised in the redox reaction.

In the example that we just looked at, oxygen causes sodium to lose an electron, making it the oxidising agent. Sodium causes oxygen to gain an electron, making it the reducing agent.

oxidising agent
a chemical species that causes oxidation and is itself reduced

reducing agent
a chemical species that causes reduction and is itself oxidised

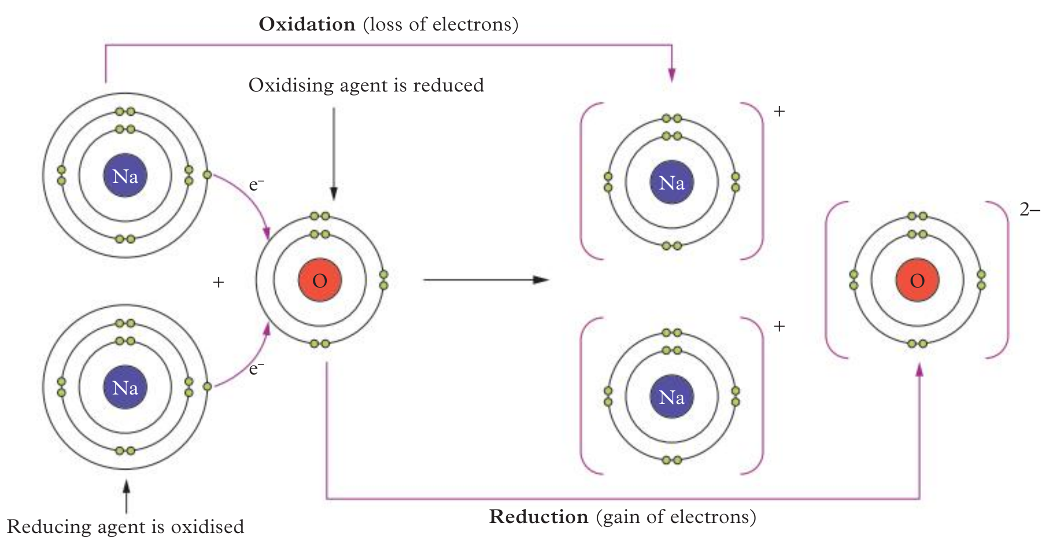

FIGURE 3 A redox reaction occurs between a reducing agent (Na) and an oxidising agent (O).

13.1 Real-world chemistry
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conjugate redox pair
a pair of chemical species that represents the gain or loss of electrons, e.g. Na/Na^+ and O/O^{2-}

13.1A Worked example
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13.1A Worked example
Video demonstration

The redox reaction has resulted in the formation of **conjugate redox pairs**. In the case of Figure 3, sodium loses an electron to form Na^+, making Na and Na^+ a redox conjugate pair. Similarly, oxygen gains electrons to form O^{2-}, making O and O^{2-} a conjugate redox pair. The strong oxidising agent, oxygen, produces a weak reducing agent, O^{2-}. Conversely, the strong reducing agent, sodium, produces a weak oxidising agent, Na^+.

See how to identify oxidising and reducing agents in Worked example 13.1A. You can also explore the origins of the term 'reduction' in Real-world chemistry 13.1.

Oxidation numbers

oxidation number the number of electrons that an atom gains or loses to form a chemical bond with another atom

We can identify the number of electrons transferred between the reducing agent and oxidising agent by using **oxidation numbers**. They help to determine the total number of electrons that have been gained or lost to form a new product. They do not always represent the charge of individual chemical elements but are used as an indicator of how many electrons an atom has. Oxidation numbers can be calculated for elements, ions or covalent molecules. Table 1 outlines the rules for assigning oxidation numbers to different chemical species.

TABLE 1 Rules for assigning oxidation numbers

Rule	Examples
Elements have an oxidation number of 0.	O_2, F_2, He, Fe, Zn, Li
Certain elements have common oxidation numbers when they are present in compounds.	- Group 1 metals (Li^+, Na^+, K^+) are always +1. - Group 2 metals (Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}) are always +2. - Hydrogen is +1 (except in metal hydrides, e.g. LiH, where it is –1). - Oxygen is –2 (except in peroxide (H_2O_2) where it is –1).
For monatomic ions, the oxidation number is given by the charge on the ion.	Cu^{2+} has an oxidation number of +2. Na^+ has an oxidation number of +1.
In polyatomic ions, the sum of the oxidation numbers is equal to the charge of the ion.	In PO_4^{3-}, the sum of the individual oxidation numbers is –3, so P + (4 × O) = –3. Oxygen is –2 and phosphorus is +5. Substituting the numbers into the equation: 5 + (4 × –2) = –3.
In a neutral compound, the sum of oxidation numbers is equal to 0.	HCl is a neutral compound and has an oxidation number of 0. As hydrogen is +1, chlorine must be –1.
The most electronegative element has a negative oxidation number.	NO_2 is a neutral compound. Oxygen is more electronegative and therefore has a negative oxidation number. Oxygen is –2 and 2 × –2 = –4. The sum of the oxidation numbers is 0, so nitrogen has an oxidation number of +4.

Study tip

Charges are written with the + or – symbol after the number, e.g. 3+. Oxidation numbers are written with the + or – symbol before the number, e.g. +3.

Study tip

A decrease in oxidation number = reduction
An increase in oxidation number = oxidation

When an atom gains one or more electrons, it becomes oxidised and gains some negative charge. Here, its oxidation number decreases. On the other hand, in reduction, electrons are gained, resulting in a decrease in oxidation number. You can see this in Worked example 13.1B.

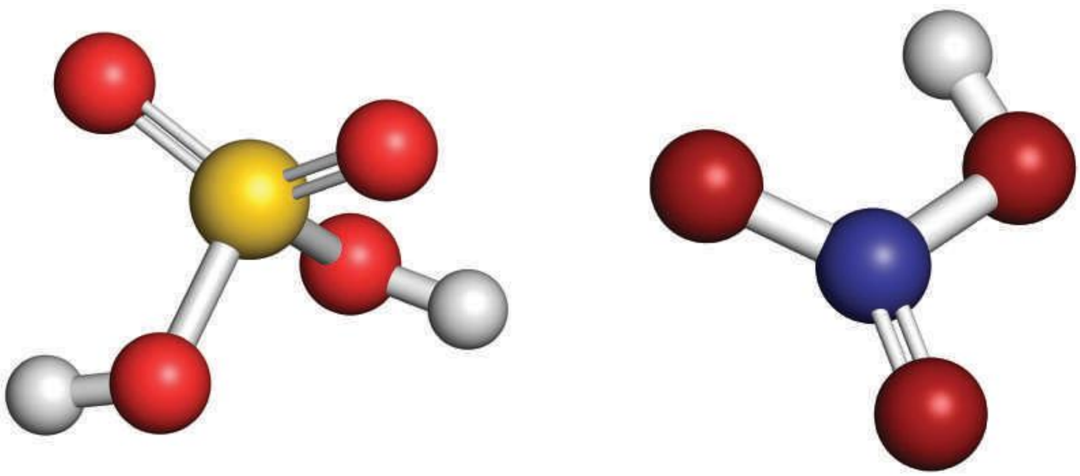

FIGURE 4 Can you determine the oxidation number of sulfur in sulfuric acid (H_2SO_4) or nitrogen in nitric acid (HNO_3), shown here?

13.1B WORKED EXAMPLE

CALCULATING OXIDATION NUMBERS

Calculate the oxidation numbers for the reactants and products and determine which chemical species is the oxidising agent and reducing agent in the following redox reaction:

$$H_2(g) + I_2(g) \rightarrow 2HI(g)$$

Solution

Think	Do
Step 1: Use the rules in Table 1 to assign oxidation numbers to all molecules and compounds.	H_2 is an element, so it has an oxidation number of 0. I_2 is an element, so it has an oxidation number of 0. HI is a neutral compound, so the sum of its oxidation numbers is 0. • When present in a compound, hydrogen always has an oxidation number of +1 (unless in a metal hydride). HI is not a metal hydride, so the oxidation number is +1. • The oxidation number of the compound is the sum of individual oxidation numbers. +1 + iodine = 0; therefore, the oxidation number of iodine is –1.
Step 2: Determine whether the oxidation number for each species has increased or decreased from reactant to product.	Hydrogen has an oxidation state of 0 as a reactant and +1 as a product. Iodine has an oxidation state of 0 as a reactant and –1 as a product.
Step 3: Determine which species are the oxidising agent and reducing agent.	The oxidation state of H_2 has increased, so it has undergone oxidation. Therefore, it is the reducing agent. The oxidation state of I_2 has decreased, so it has undergone reduction. Therefore, it is the oxidising agent.

Study tip

Overall redox reactions should not have electrons. These should be cancelled out.

half-equation
an equation that represents either the oxidation or the reduction half of a chemical equation; it shows the transfer of electrons

overall redox equation
an equation formed by combining the two half-equations; spectator ions are not shown, and electrons are balanced and cancelled out

Simple half- and overall redox equations

As redox reactions involve chemical species which undergo oxidation and reduction, they can be separated into two **half-equations**. When these two half-equations are combined, an **overall redox equation** is formed.

Half-equations

Half-equations show the transfer of any electrons. They do not include spectator ions. These ions have the same oxidation number and state on both sides of the equation. Like a spectator at a football game, they do not participate in the redox reaction. This means that half-equations only show the oxidising agent gaining electrons, and reducing agent losing electrons. Worked example 13.1C walks you through how to write half-equations.

Overall redox equations

Once the oxidation and reduction half-equations have been identified, they can be combined to form an overall redox equation. To do this, both half-equations must have the same number of electrons so that they can cancel out in the final equation. You can see this in Worked example 13.1D.

13.1C Worked example
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13.1C Worked example
Video demonstration

13.1D Worked example
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13.1D Worked example
Video demonstration

complex half-equation
a half-equation for an oxidation and reduction equation that occurs under acidic or alkaline conditions

complex overall redox equation
an overall redox equation for a redox reaction that occurs under acidic or alkaline conditions

13.1F Worked example
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13.1F Worked example
Video demonstration

Complex half- and overall redox equations

Often, redox equations cannot be balanced by simply balancing atoms adding electrons. More **complex half-equations** and **complex overall redox equations** are required for molecules or polyatomic atoms that contain hydrogen or oxygen (e.g. H_2O, CH_3COOH, $Cr_2O_7^{2-}$, MnO_4^- and SO_4^{2-}).

These reactions only occur under acidic or alkaline conditions. In Units 1 and 2, you only need to know the steps for writing redox equations in acidic conditions. This involves the KOHES method, where we use hydrogen ions (H^+) to balance the equations.

1 K – Balance the **key** elements (everything except oxygen and hydrogen).
2 O – Balance the **oxygen** by adding $H_2O(l)$ molecules to the opposite side.
3 H – Balance the **hydrogen** by adding $H^+(aq)$ ions to the opposite side.
4 E – Balance the charge by adding **electrons** to one side.
5 S – Assign **states** to all reactants and products except electrons, which have no state.

See how to write a complex half- and full redox equation in Worked examples 13.1E and 13.1F.

13.1E WORKED EXAMPLE

WRITING COMPLEX HALF-EQUATIONS IN ACIDIC CONDITIONS

Complete the half-equation for the dichromate ion ($Cr_2O_7^{2-}$) forming the chromium ion (Cr^{3+}). Determine whether this is an oxidation or reduction half-equation:

$$Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$$

Solution

Think	Do
Step 1: K – Balance any **key** elements. Key elements are all elements other than oxygen and hydrogen.	Cr is the key element. The left side has two Cr atoms, so we need to add a coefficient of 2 to the product side. $Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq)$
Step 2: O – Balance the **oxygen** by adding H_2O molecules to the opposite side.	There are seven oxygens on the left, so we need to add seven H_2O molecules to the right. $Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O$
Step 3: H – Balance the **hydrogen** by adding H^+ to the opposite side.	There are 14 hydrogens on the right, so we need to add 14 H^+ to the left. $Cr_2O_7^{2-}(aq) + 14H^+ \rightarrow 2Cr^{3+}(aq) + 7H_2O$
Step 4: E – Balance the charge by adding **electrons** to one side.	$Cr_2O_7^{2-}(aq) + 14H^+ \rightarrow 2Cr^{3+}(aq) + 7H_2O$ The total charge on the left is –2 + 14 = +12 The total charge on the right is +3 × 2 = +6 To balance the charges, we need to add six negatively charged electrons to the left to make both sides +6. $Cr_2O_7^{2-}(aq) + 14H^+ + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O$
Step 5: S – Assign **states** to all reactants and products with the exception of the electrons, which have no state.	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
Step 6: Determine whether this is an oxidation or reduction half-equation.	Electrons are gained (on the left side of the equation), so this is a reduction half-equation.

Study tip

Sometimes you can skip some steps in the KOHES method. For example, if the reactant and product do not contain oxygen, water does not need to be added.

Study tip

Look at either the change in oxidation number or the side of the half-equation that the electrons are on, to determine whether it is an oxidation or reduction half-equation.

13.1 CHALLENGE

Writing equations for complex redox reactions

When reacted with a solution of acidified sodium dichromate ($Na_2Cr_2O_7$), methanol (CH_3OH) is oxidised to methanoic acid (CH_3COOH) and the dichromate ($Cr_2O_7^{2-}$) is reduced to chromium(III) (Cr_2O_3).

1 Write the oxidation half-equation.
2 Write the reduction half-equation.
3 Write the balanced overall equation.

13.1 CHECK YOUR LEARNING

Describe and explain

1 Describe how the processes of reduction and oxidation are related.
2 Describe what happens to the valence shell electrons during redox reactions.
3 Explain the difference between an oxidation and a reduction half-equation.

Apply, analyse and compare

4 The following redox reactions were investigated:
   I Lithium metal and gaseous sulfur (S) reacting to form solid lithium sulfide
   II Magnesium metal and fluorine gas (F_2) reacting to form solid magnesium fluoride
   III An aqueous solution of silver nitrate reacting with manganese(II) metal to form aqueous manganese nitrate and silver metal

   For each reaction:
   a identify the species that is being oxidised and the species that is being reduced
   b identify the oxidising and reducing agent
   c write the oxidation and reduction half-equations
   d write the overall redox equation.

5 Complete and balance the following half-equations, then use oxidation numbers to identify the species undergoing oxidation and the species undergoing reduction.
   a $H_2(g) + Cl_2(g) \rightarrow HCl(aq)$
   b $Cu(NO_3)_2(aq) + Fe(s) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$
   c $2AgNO_3(aq) + Pb(s) \rightarrow Pb(NO_3)_2(aq) + 2Ag(s)$
6 Write balanced half-equations for the reactions in Question 5, including states.
7 For the following pairs of chemicals, write balanced half-equations and use oxidation numbers to determine whether oxidation or reduction has occurred.
   a MnO_4^-/MnO_2
   b BiO_3^-/Bi^{3+}
   c SO_3^{2-}/SO_4^{2-}
   d HNO_2/NO_3^{2-}

Design and discuss

8 Oxidation is a name given to many chemical reactions. It can also be defined as a gain in oxygen or a loss of hydrogen. Evaluate these definitions and provide examples of these types of reactions.

13.2 The reactivity series of metals

KEY IDEAS

In this topic, you will learn that:

+ the periodic table is used to predict which elements will more readily gain or lose electrons in redox reactions
+ the reactivity series of metals orders metals from strongest to weakest reducing agents
+ metal displacement reactions are redox reactions in which a more reactive metal displaces a less reactive metal ion to form a new compound
+ more complex molecules or ions can also gain and lose electrons, but this cannot be predicted by the periodic table – instead, the electrochemical series is used.

Some elements, atoms, molecules or ions more readily gain or lose electrons than others. This is referred to as their reactivity. Because oxidation is a loss of electrons, a chemical species that can easily lose its electrons will undergo oxidation more readily and is considered a stronger reducing agent. Similarly, because reduction is a gain in electrons, a chemical species that can more easily gain electrons will undergo reduction more readily and is a stronger oxidising agent.

In this topic, you will learn how to use the periodic table and reactivity series of metals to predict which chemical species will be more reactive than others. The electrochemical series can also be used, but you don't need to learn about this until Units 3 and 4.

Predicting redox reactions using the periodic table

In Chapter 2, you learned about the trends in first ionisation energy (the energy required for each element to lose one electron) on the periodic table.

Group 1 and 2 elements readily lose their valence electrons to form more stable positive cations. By losing these electrons, they become oxidised and act as reducing agents. Some metals are stronger reducing agents than others because they more readily donate their valence electrons. Generally, the fewer valence electrons a metal has, the less energy is required to lose them. Therefore, elements in the bottom left of the table have the lowest ionisation energies and will lose their valence electrons more readily.

Meanwhile, group 17 elements (which have more valence electrons and are more electronegative) readily gain an electron to form a more stable negative anion. By gaining electrons, they become reduced and act as oxidising agents. Some non-metals are stronger oxidising agents than others because they more readily accept valence electrons.

The reactivity series of metals

reactivity series of metals
a list that orders the metals in the periodic table from most reactive (strongest reducing agent) to least reactive (weakest reducing agent)

The **reactivity series of metals** is a list of metals arranged according to the trends in the periodic table. This list orders metal atoms from most reactive to least reactive (Figure 1).

The reactivity series of metals makes it very easy to predict the oxidising and reducing strengths of different chemical species. See how this is done in Worked example 13.2A.

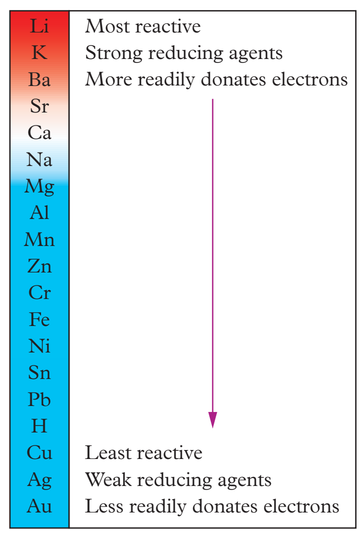

FIGURE 1 The metal reactivity series

13.2A WORKED EXAMPLE

PREDICTING THE STRENGTHS OF OXIDISING AND REDUCING AGENTS

Identify the stronger reducing agent out of aluminium and magnesium. Justify your choice, by using the periodic table and reactivity series of metals.

Solution

Think	Do
Step 1: Look at the periodic table. Compare the number of valence electrons.	Magnesium is in group 2 and aluminium is in group 13. This means that magnesium has two valence electrons, which are easier to lose than aluminium's three valence electrons.
Step 2: Look at the electrochemical series. Compare the strengths of the oxidising and reducing agents.	Magnesium is further down the electrochemical series than aluminium, so it will lose its electrons and undergo oxidation more readily than aluminium.
Step 3: Identify the stronger reducing agent.	Magnesium is a stronger reducing agent than aluminium.

Predicting displacement reactions of metals

A **displacement reaction** is a type of redox reaction that occurs when a more reactive element replaces a less reactive element in a compound. The formation of new compounds by displacement reactions can be predicted by using the reactivity series of metals. Let's consider Worked examples 13.2B and 13.2C.

displacement reaction
a type of redox reaction in which a more reactive element replaces (or displaces) a less reactive element to form a new compound; the metal ions of one metal must be above the other metal in the reactivity series of metals

13.2B WORKED EXAMPLE

PREDICTING DISPLACEMENT REACTIONS USING THE METAL REACTIVITY SERIES

Consider the following possible reactions:

Reaction 1: $ZnSO_4(aq) + Ag(s)$ Reaction 2: $AuNO_3(aq) + Ni(s)$

a Use the metal reactivity series to predict whether a reaction will occur.

b If a reaction will occur, identify the products and write a balanced equation.

Solution

Think	Do
Step 1: Identify the metals in each reactant and compare their reactivity on the metal reactivity series.	Reaction 1: Zn is more reactive than Ag. Reaction 2: Ni is more reactive than Au.
Step 2: Determine whether a displacement reaction will occur. This will happen if the metal in the compound is less reactive than the lone metal.	Reaction 1: Ag will not displace Zn. Therefore, a reaction will not occur. Reaction 2: Ni will displace Au. Therefore, a reaction will occur.
Step 3: Identify the products of the reaction.	Reaction 2 will result in the formation of $Ni(NO_3)_2(aq)$ and Au(s).
Step 4: Write a full balanced equation for the reaction.	$2AuNO_3(aq) + Ni(s) \rightarrow Ni(NO_3)_2(aq) + 2Au(s)$

13.2C Worked example
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13.2C Worked example
Video demonstration

Study tip

A solid metal higher up on the reactivity series of metals will displace a metal (in aqueous solution) lower on the series.

Another common example of a metal displacement reaction is the reaction that occurs when a coiled copper wire is placed into a solution of silver nitrate. The chemical equation for this reaction is:

$$AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

The solid copper (Cu) sits above silver (Ag) on the reactivity series of metals. This means that it will displace the silver in solution. This can be seen in Figure 2. $Cu^{2+}(aq)$ ions have been formed, which gives the solution a blue colour. Ag(s) crystals have started forming on the coil.

FIGURE 2 Copper displaces silver nitrate from solution.

13.2 REAL-WORLD CHEMISTRY

Redox in nature: photosynthesis and cellular respiration

Plants are autotrophs, which means they produce their own energy from sunlight. To generate this energy, they must absorb water from soil into their roots, and carbon dioxide from the atmosphere into their leaves. These compounds are carried to an organelle called the chloroplast (Figure 4). Here, they react in the presence of sunlight to form glucose and oxygen.

This reaction is called photosynthesis, and is given by the following chemical equation:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

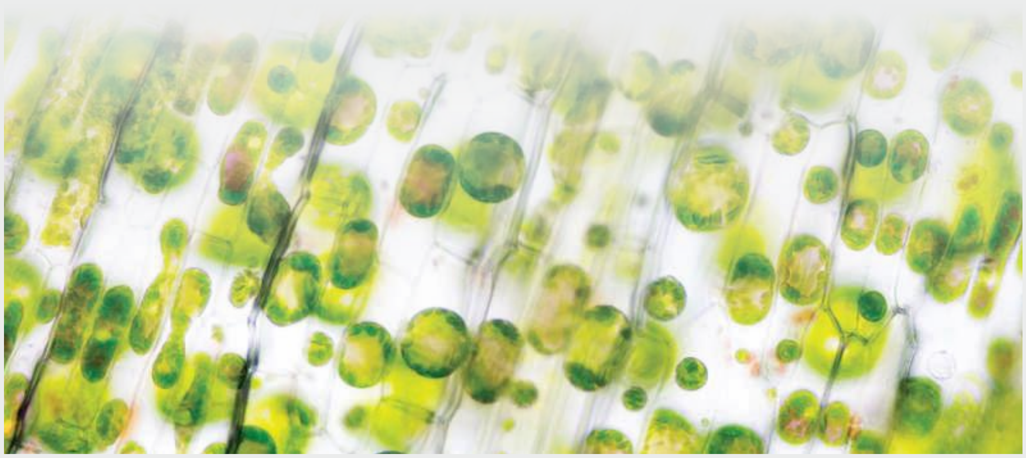
Humans are heterotrophs, which means we cannot make our own energy and must get energy through consuming food, such as plants. When consumed, the sugars that the plant has produced are ingested and broken down within the body for energy. This process, called cellular respiration, occurs according to the chemical equation:

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

Apply your understanding

1 Write the reduction half-equation in the photosynthesis reaction.
2 Write the oxidation half-equation in the photosynthesis reaction.
3 In the cellular respiration equation, identify the oxidising and reducing agents, using oxidation numbers to support your answer.

FIGURE 3 A microscope image of plant cells. The green chloroplasts are where photosynthesis occurs.



13.2 SKILL DRILL

Designing an experiment to compare the reactivity of metals

Key science skill: Plan and conduct investigations

Several metals are reacted with 1 M hydrochloric acid (HCl) to determine which is most reactive. The metals tested are calcium, magnesium and aluminium.

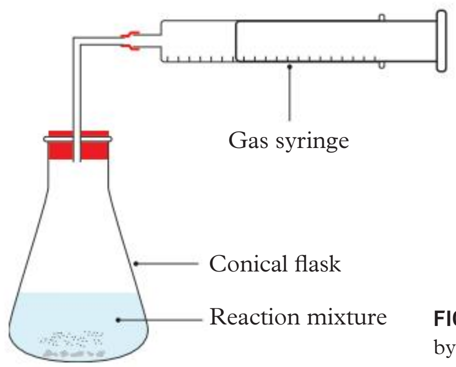


FIGURE 4 The reactivity of metals can be tested by measuring the hydrogen gas produced.

The metals were tested by determining the volume of hydrogen gas produced by each reaction every 10 seconds.

Practise your skills

1 Write an aim and hypothesis for this experiment.

2 Identify the independent and dependent variables.

3 Identify the materials and equipment you would need for this experiment.

4 Explain whether the data will be qualitative or quantitative.

Need help planning investigations? See Topics 1.3 and 1.4 (pages 12–15).

13.2 CHECK YOUR LEARNING

Describe and explain

1 Explain the trend in reactivity across a period in the periodic table.

2 Explain the trend in reactivity down a group in the periodic table.

3 Explain why solid metals can only undergo oxidation and non-metal elements from group 17 (Cl_2 or F_2) can only undergo reduction.

Apply, analyse and compare

4 Predict whether a reaction will occur between the following reactants. If you predict that a reaction will occur, write the products and balance the equation.

- **a** $Al(NO_3)_3(aq) + Sn(s)$
- **b** $Ni(NO_3)_2(aq) + Zn(s)$
- **c** $Cu(NO_3)_2(aq) + Fe(s)$
- **d** $LiNO_3(aq) + Cu(s)$

5 For each of the pairs of metals below, identify the stronger reducing agent. Justify your choice, by using the periodic table and reactivity series of metals.

- **a** Lithium and tin
- **b** Zinc and iron
- **c** Manganese and chromium

13.3 Redox reactions in society

KEY IDEAS

In this topic, you will learn that:

+ corrosion is a type of redox reaction observed in nature
+ redox reactions can be used in primary cells to convert chemical energy into electrical energy.

We use redox reactions every day. They form the basis of some of the most important reactions we rely on, including:

- combustion reactions
- corrosion reactions
- the production of electrical energy using simple primary cells.

In this topic, we will explore these applications of redox reactions.

Combustion of chemical compounds

combustion
a chemical reaction with oxygen to form a metal oxide, covalent compound or carbon dioxide and water

metal oxide
an ionic compound where a metal cation is electrostatically attracted to an oxide anion (O^{2-})

Combustion reactions are reactions with oxygen that produce heat. Both metals and non-metals can undergo combustion to form an ionic or a covalent compound. In this topic, you will be introduced to the combustion of metals. You will learn a lot more about combustion reactions, particularly of non-metal compounds, in Units 3 and 4.

A metal will combust to form a metal oxide, according to the following equation:

$$\text{metal} + \text{oxygen} \rightarrow \text{metal oxide}$$

A **metal oxide** is an ionic compound consisting of a positive metal cation and a negative oxide anion. Consider the reaction between copper and oxygen:

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

Copper metal loses electrons to form a positive metal cation. Therefore, copper undergoes oxidation. Oxygen gas gains electrons to form a negative non-metal anion. Therefore, oxygen undergoes reduction.

Combustion in society

Combustion is the primary reaction used to generate electrical energy from non-metal hydrocarbon fuels. We also see combustion when we burn wood, light a match, and when fireworks light up the sky during New Year's celebrations.

FIGURE 1 Combustion occurring at the end of a lit match

FIGURE 2 Combustion of metal salts gives firework displays their colours.

13.3
PRACTICAL: LITERATURE REVIEW

Are redox reactions beneficial or harmful to society and the environment? Go to page **518.**

Corrosion of metals

Corrosion is a process in which metals react with chemicals in the atmosphere (including rain and atmospheric gases such as oxygen and carbon dioxide) to form a more stable chemical compound. It is a type of reaction. For this to occur, the metal must be more stable as an ion than it is as a solid metal.

corrosion
the degradation of a metal to form a more stable metal oxide when exposed to gases and liquids

In a corrosion reaction, the metal readily donates an electron according to the general equation, where M is the metal:

$$M(s) \rightarrow M^+(aq) + e^-$$

Like combustion reactions, corrosion forms metal oxides. However, corrosion reactions are more complex. Let's look at the reaction between iron(II) and water.

1 When water interacts with the surface of the iron metal, iron loses two electrons to form Fe^{2+}. Therefore, iron is oxidised:

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

2 Oxidation cannot occur without reduction, so reduction occurs at the surface of the water droplet. Here, water reacts with oxygen to form hydroxide ions:

$$O_2(g) + H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

3 Fe^{2+} and hydroxide ions react with each other to form iron(II) hydroxide:

$$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(aq)$$

4 The iron hydroxide also reacts with oxygen at the surface to form iron(III) hydroxide:

$$4Fe(OH)_2(aq) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(aq)$$

5 The iron(III) hydroxide decomposes in the presence of oxygen to form a hydrated iron oxide complex:

$$4Fe(OH)_3(aq) \rightarrow Fe_2O_3.H_2O(aq)$$

Iron oxide gives rust its characteristic brown colour (Figure 3).

Put more simply, the corrosion process is as follows:

1 The metal undergoes oxidation.
2 Water undergoes reduction in the presence of oxygen.
3 The products of these redox reactions react in the presence of oxygen to form a metal hydroxide. This may undergo further reactions to form another metal hydroxide.
4 The final metal hydroxide then decomposes to form a hydrated metal oxide complex (a metal oxide associated with water molecules).

FIGURE 3 The corrosion of iron metal forms red-brown rust.

FIGURE 4 The Statue of Liberty

Corrosion in society

Corrosion is commonly observed for metals exposed to the environment. Examples include the rusting of old cars, water pipes and steel surfaces on buildings. The green colour of the Statue of Liberty in New York is also a result of corrosion.

Reactions in batteries

Reactions can transform the chemical energy stored in a chemical species into other forms of energy. Because redox reactions involve a transfer of electrons, they can be used to generate an electric current. These reactions are used in **batteries** to power many devices.

battery
a container in which reactants are stored to generate electrical energy from chemical energy

Reactants in direct contact

In the previous examples of redox reactions, the reactants are in direct contact with each other. The transfer of electrons between reactants that are in direct contact results in the conversion of chemical energy into thermal energy, the release of heat, and the formation of a solid product (precipitate), as in Figure 5.

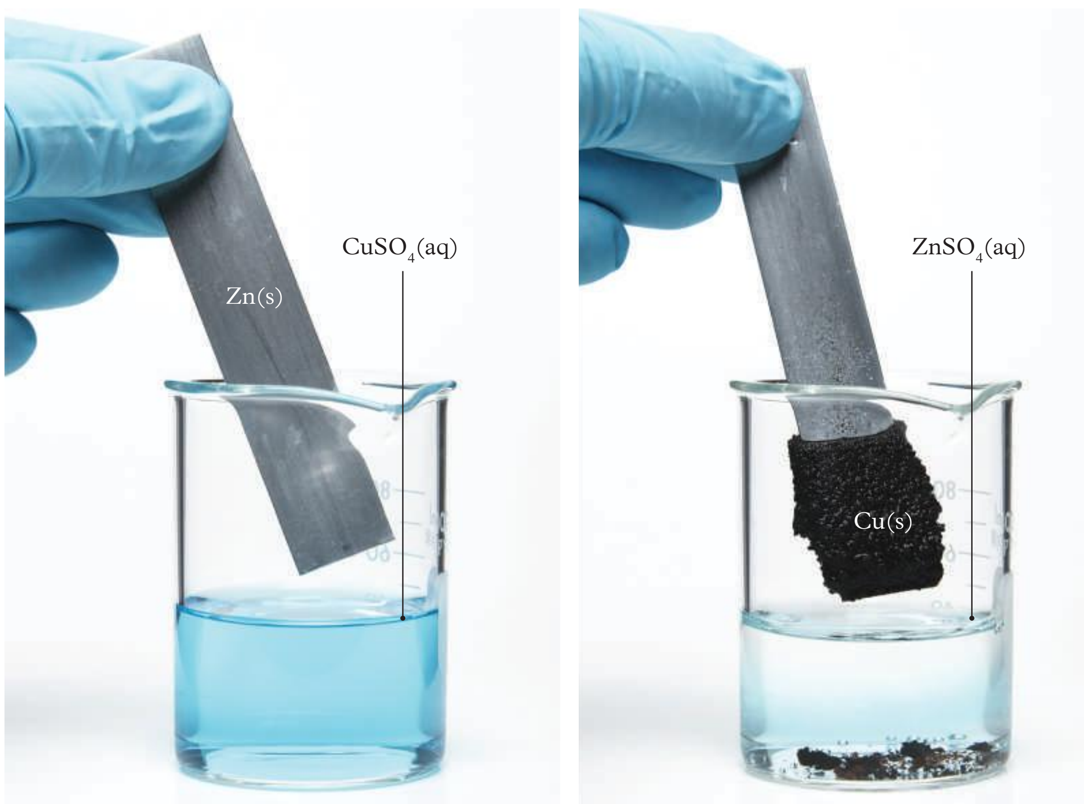


FIGURE 5 Zinc metal displaces copper from copper(II) sulfate solution to form a copper metal precipitate and zinc sulfate solution. This redox reaction generates heat.

Reactants in separate containers – galvanic cells

Redox reactions can also occur when reactants are not in direct contact. These occur in **galvanic cells**, which are circuits where redox reactions occur spontaneously and convert chemical energy to electrical energy. In this arrangement, little heat is released. There are two main types of galvanic cells:

- **primary cells** – which cannot be recharged and must be disposed of after use
- **secondary cells** – which can be reused because they are rechargeable.

In Unit 2 of VCE Chemistry, you will only look at simple primary cells, whereas both primary and secondary cells will be explored in Unit 3. Let's see how they work.

In a galvanic cell, the oxidation half-reaction and reduction half-reaction occur in different compartments or **half-cells**. In Figure 6, a galvanic cell is shown. Each galvanic cell is composed of two half-cells.

galvanic cell
a circuit in which a redox chemical reaction occurs to generate electrical energy; the reduction and oxidation half-reactions occur in separate compartments (or cells)

primary cell
a type of galvanic cell that cannot be recharged; it is disposed of after the redox reaction is complete

secondary cell
a type of galvanic cell that can be recharged; it can be reused because the redox reaction is reversible

half-cell
a compartment that contains either the oxidation or the reduction half-reaction

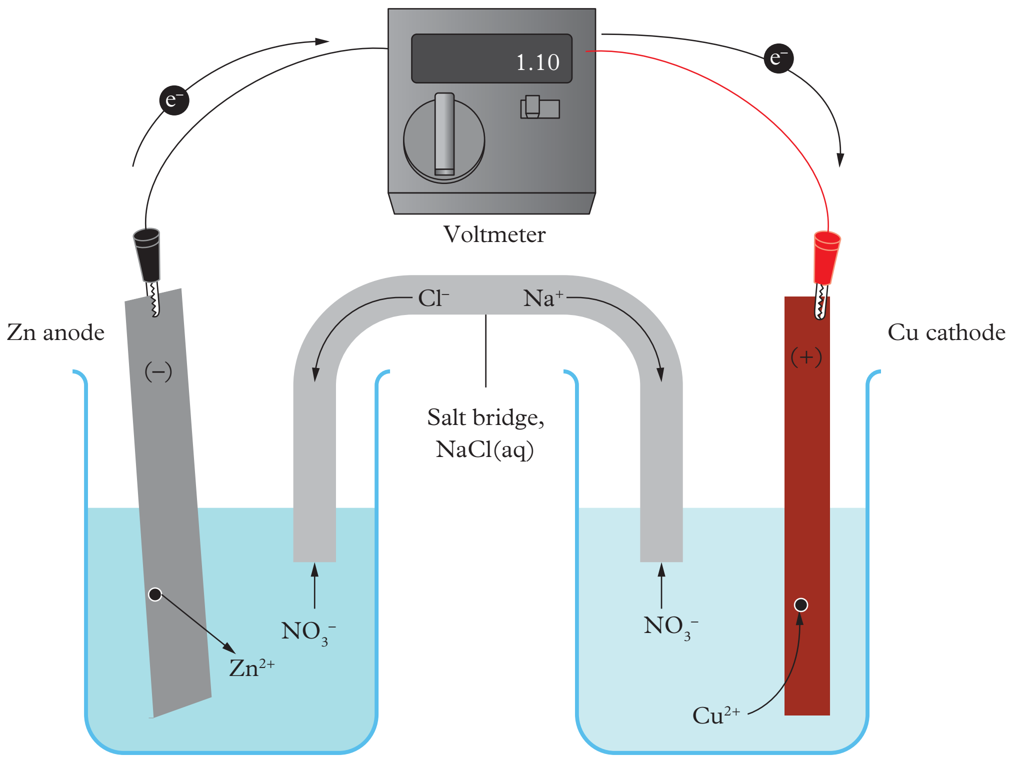


FIGURE 6 A galvanic cell

The oxidation half-equation is: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

The reduction half-equation is: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

The reactants are physically separated, and the half-equation reactions occur at the **electrodes** (metal strips). Oxidation occurs at the negatively charged electrode – the **anode** (Zn). Reduction occurs at the positively charged electrode – the **cathode** (Cu). The redox reactions occur at the surface of each electrode.

Each electrode is placed in a solution containing the reactant in ionic form (Cu^{2+} and Zn^{2+}). This is called the electrolyte. The anion in both solutions is nitrate (NO_3^-). This has been chosen because it is insoluble and will not form a precipitate with any metal cation. For this reason, nitrate is a spectator ion and helps to maintain the charge in each half-cell. The copper half-cell loses positive charge as it becomes Cu(s), so the nitrate ions move out of the half-cell to balance the loss of positive charge.

electrode
a strip or rod that conducts electricity; this is the surface on which the oxidation or reduction reactions occur

anode
the negatively charged electrode

cathode
the positively charged electrode

salt bridge
a connection between solutions in a galvanic cell that allows the flow of charge by acting as a pathway for moving ions

internal circuit
the pathway for movement of ions between half-cells

external circuit
the pathway for movement of electrons from the anode to the cathode

The solutions are connected by a **salt bridge** to form an **internal circuit**. In Figure 6, this is made of sodium chloride (NaCl), which allows the transfer of soluble ions. Na^+ cations move towards the cathode to balance the charge when Cu^{2+} forms Cu, and Cl^- anions move towards the anode to balance the charge when Zn^{2+} increases in concentration as Zn loses electrons.

The anode and cathode are connected to each other by a wire. This wire is how electrons are transferred between half-cells. Electrons move from the anode, where electrons are lost by zinc, to the cathode, where they are gained by copper. Often, a voltmeter is included in the circuit to show the voltage produced by the cell. This is the **external circuit**.

The overall redox equation is: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

In a primary cell, the electrons only move in one direction. Once the reducing agent has lost all of its electrons, or the circuit breaks, the reaction stops. You will learn about this in much more depth in Units 3 and 4.

13.3 REAL-WORLD CHEMISTRY

Repainting the Eiffel Tower

The Eiffel Tower was designed by Gustave Eiffel to be the main attraction in the 1889 Paris world's fair. It was constructed from pure iron, which was the best and most robust construction material at the time. The tower was intended to stand for 20 years.

Although the reaction mechanisms behind corrosion were unknown, the effects of corrosion were. As such, Eiffel knew that he had to cover the metal to protect it. Venetian red paint was applied to the iron structure in the workshop before the tower was assembled.

It took two years, two months and five days to fully assemble the 300-metre-high tower. This was completed by 31 March 1889, in time for its grand opening at the world's fair on 5 May.

Because of the chipping of the paint, general wear and tear and exposure to weather, the tower is repainted approximately every 7 years. Paint choices have changed over time, but the current brown colour was first used in 1968. This was chosen because it coordinated better with the Parisian landscape.

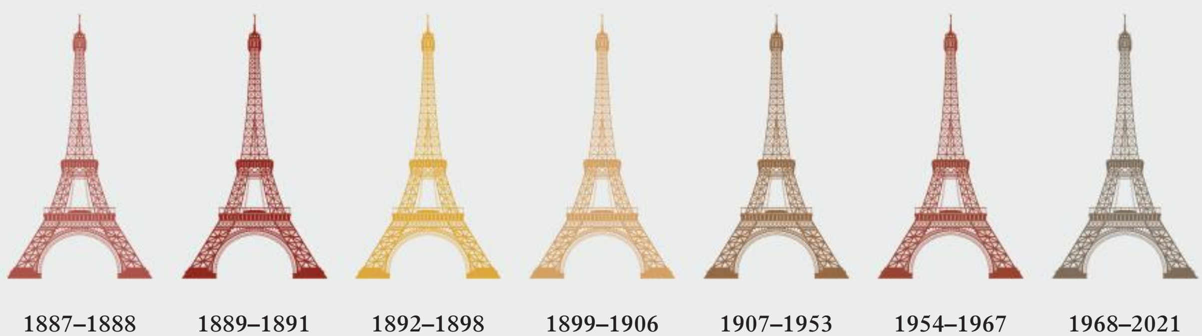


FIGURE 7 Colour changes of the Eiffel Tower since 1887

Apply your understanding

1 The tower is made of iron. Write the redox reactions that occur when iron reacts with oxygen in the atmosphere to form rust (Fe_2O_3).

2 Explain the benefits of painting the Eiffel Tower.

3 Explain why the Eiffel Tower has been repainted approximately every 7 years.

Primary cells in society

You encounter primary cells every day. They are in single-use batteries such as AA or AAA batteries, where the electrical energy produced acts as a power source. The disadvantage of a primary cell is that, once all of the chemical energy has been converted to electrical energy and used up, the battery dies and must be disposed of.

FIGURE 8 Primary cells must be disposed after use.

13.3 CHECK YOUR LEARNING

Describe and explain

1 Describe the process of corrosion, using chemical equations to support your answer.

2 Explain how redox reactions can convert chemical energy into electrical energy.

Apply, analyse and compare

3 Use chemical equations to explain how corrosion can be prevented by adding a coating of zinc onto a piece of iron.

4 For each pair of species a–c, draw a galvanic cell diagram that includes:
- an anode and a cathode
- the material that each electrode is made of
- the direction of electron movement
- a suitable salt bridge solution
- the direction of ion movement in the salt bridge
- the direction of ion movement in each half-cell.

  a $Ag^{+}(aq)/Ag(s)$ and $Ni^{2+}(aq)/Ni(s)$

  b $Pb^{2+}(aq)/Pb(s)$ and $Mg^{2+}(aq)/Mg(s)$

  c $Co^{2+}(aq)/Co(s)$ and $Zn^{2+}(aq)/Zn(s)$

5 Write half-equations and the overall redox equation for the redox reactions in Question 4.

Design and discuss

6 You teacher claims that there are circumstances when the corrosion of a metal can be beneficial but does not explain why this is so. Your friend is absent from this Chemistry lesson and asks you to explain it to them. Investigate a circumstance where this is true and communicate the answer that you would give to your friend.

7 Simple primary cells are non-rechargeable and must be thrown away after use. In contrast, secondary cells can be recharged and reused.

  a Discuss how primary cells are problematic according to the green chemistry principle of preventing waste.

  b Identify the United Nations Sustainable Development Goal(s) that this could relate to.

  c Research the use of secondary cells and evaluate whether they help to address the green chemistry principle of preventing wastes.

CHAPTER 13 Review

Chapter summary

13.1
- Redox reactions occur when electrons are transferred from one reactant to another. A reactant undergoes oxidation if it loses electrons (OIL), whereas a reactant undergoes reduction if it gains electrons (RIG).
- The reactant that loses one or more electrons causes reduction in the other reactant and is called the reducing agent. The reactant that accepts one or more electrons causes oxidation in the other reactant and is called the oxidising agent.
- Oxidation numbers can be used to determine whether a chemical species has undergone oxidation or reduction.
- Half-equations represent the loss or gain of electrons. Simple half-equations can be constructed by balancing the charge by using electrons.
- Half-equations can be combined into an overall redox reaction equation. The overall redox equation does not show the transfer of electrons.
- For redox reactions occurring under acidic conditions, half-equations cannot always be balanced just by adding electrons. The KOHES method must be followed for these complex half-equations. These half-equations can also be combined to form a complex overall redox equation.

13.2
- The reactivity series of metals lists elements in order of reactivity. The more electronegative an element is, the more readily it will accept electrons and undergo reduction. The lower the ionisation energy of an element, the more readily it will lose a valence electron and undergo oxidation.
- More reactive elements will displace less reactive elements in a solution. This is represented in a displacement reaction.
- The reactivity series of metals can be used to predict whether metal displacement reactions will occur.

13.3
- Redox reactions occur often in society. Examples are combustion and corrosion reactions.
- When redox reactants are in direct contact, they will convert chemical to thermal energy.
- When separated into half-cells, redox reactants will convert chemical to electrical energy with a small amount of energy wasted as heat. This involves the transfer of electrons through a circuit, creating a galvanic cell. This is the foundation of batteries.

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can…	Confidently	Sort of	Not really	Revision link
• explain how electrons move between reactants in oxidation and reduction	☐	☐	☐	Go back to Topic 13.1
• determine the oxidation numbers of different chemical species and use them to identify redox reactions, including identification of oxidising agents and reducing agents	☐	☐	☐	Go back to Topic 13.1

I can…	Confidently	Sort of	Not really	Revision link
• write balanced half-equations and overall redox equations: including simple half- and overall equations, and complex reactions in acidic conditions (using KOHES)	☐	☐	☐	Go back to Topic 13.1
• use the reactivity series of metals to predict whether a displacement redox reaction will occur	☐	☐	☐	Go back to Topic 13.2
• explain how redox reactions are used in society, including: combustion of metals, corrosion and simple primary cells	☐	☐	☐	Go back to Topic 13.3

Revision questions

Multiple choice

1 The oxidation number of phosphorus in $P_2O_7^{4-}$ is:

A +10

B +7

C +5

D +3.5

Use the following equation to answer Questions 2 and 3.

$H_2Te(g) + S(s) \rightarrow Te(s) + H_2S(g)$

2 Identify the atoms that undergo oxidation and reduction.

A H_2Te is oxidised; S is reduced.

B S is oxidised; H_2Te is reduced.

C Te is oxidised; H_2S is reduced.

D H_2S is oxidised; Te is reduced.

3 Identify the oxidising agent and reducing agent.

A H_2Te is the oxidising agent; S is the reducing agent.

B S is the oxidising agent; H_2Te is the reducing agent.

C Te is the oxidising agent; H_2S is the reducing agent.

D H_2S is the oxidising agent; Te is the reducing agent

4 Which of the following represents an oxidation half-equation?

A $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$

B $O_2(g) + 4H^+(aq) + 2e^- \rightarrow 2H_2O(l)$

C $AuCl_4(aq) + 3e^- \rightarrow Au(s) + 4Cl^-(aq)$

D $2Ag(s) + S^{2-}(aq) \rightarrow Ag_2S(s) + 2e^-$

5 The oxidation number of phosphorus in HPO_4^{2-} is:

A +2

B +3

C +4

D +5

6 Which of the following reactions is a redox reaction?

A $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

B $CuSO_3(aq) + Zn(s) \rightarrow ZnSO_4(aq) + Cu(s)$

C $NaOH(aq) + NH_4Cl(aq) \rightarrow NaCl(aq) + NH_4OH(aq)$

D $NaClO_3(aq) + KOH(aq) \rightarrow NaOH(aq) + KClO_3(aq)$

7 Consider the following reaction:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

The oxidising agent is:

A Na(s)

B $Cl_2(g)$

C $Na^+(s)$

D $Cl^-(s)$

8 Consider the following reaction:

$$Br_2(l) + Pb(s) \rightarrow PbBr_2(l)$$

The reducing agent is:

A $Br_2(l)$

B Pb(s)

C $Pb^{2+}(l)$

D $Br^-(l)$

9 Consider the following reaction:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

The oxidising agent is:

A $H(g)$

B $O_2(g)$

C $H^+(aq)$

D $OH^-(aq)$

10 Consider the following reaction:

$$Sn(NO_3)_2(aq) + 2K(s) \rightarrow 2KNO_3(aq) + Sn(s)$$

The reducing agent is:

A $Sn(NO_3)_2(aq)$

B $K(s)$

C $KNO_3(aq)$

D $Sn(s)$

Short answer

Describe and explain

11 Describe how an atom's ability to gain or lose electrons is represented by its location on the periodic table.

12 Describe the relationship between an oxidising agent and a reducing agent.

13 Describe the relationship between oxidation and reduction.

14 Explain how the reactivity series of metals can be used to determine the strongest oxidising agent and strongest reducing agent in a reaction.

15 Explain the difference in the outcomes of redox reactions when reactants are:

a in direct contact in the same beaker

b separated into different beakers and connected by wire.

Apply, analyse and compare

16 Determine the oxidation number of:

a Mn in MnO_4^-

b S in SO_4^{2-}

c Te in TeO_3^{2-}

d N in NH_3

17 Three redox reactions are shown below.

I Sodium metal and gaseous bromine form solid sodium bromide.

II Aluminium metal and hydrochloric acid form aluminium chloride and hydrogen gas.

III Liquid water reacts with lithium metal to form lithium hydroxide and hydrogen gas.

For each reaction:

a identify the species that undergoes oxidation and reduction

b identify the reducing and oxidising agents

c write balanced half-equations

d write a balanced overall redox equation.

18 For the following pairs of chemicals, write balanced half-equations under acidic conditions and use oxidation numbers to identify the oxidising agent and reducing agent in each equation.

a NO/NO_2^-

b SO_4^{2-}/SO_2

c TeO_3^{2-}/Te

d IO^-/IO_3^-

e $As/H_2AsO_4^-$

f ReO_4^-/Re

19 Using oxidation numbers, determine which species are oxidised and reduced in the following reactions.

a $2AgNO_3(aq) + Mn(s) \rightarrow Mn(NO_3)_2(aq) + 2Ag(s)$

b $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

c $Ca(s) + Br_2(l) \rightarrow CaBr_2(s)$

d $Hg(NO_3)_2(aq) + Fe(s) \rightarrow Hg(l) + Fe(NO_3)_2(aq)$

e $C(s) + O_2(g) \rightarrow CO_2(g)$

20 Balance each of the following redox equations by separating them into their half-equations and combining them into their balanced overall equation.

a $MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)$

b $H_2S(aq) + NO_3^-(aq) \rightarrow S(s) + NO_2(g)$

c $PbS(s) + NO_3^-(aq) \rightarrow Pb^{2+}(aq) + S(s) + NO(g)$

d $Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$

e $H_3AsO_4(aq) + Sn^{2+}(aq) + H^+(aq) \rightarrow H_3AsO_3(aq) + Sn^{4+}(aq) + H_2O(l)$

21 Predict whether a reaction will occur between the following reactants. Justify your answer.

a $Cl_2(g)$ and Mg(s)

b Zn(s) and $H_2(g)$

c Bromine and lithium sulfate

d Iron(III) nitrate and hydrogen sulfide gas

22 Consider the following redox equations.

I $Br_2(g) + H_2(g) \rightarrow 2HBr(aq)$

II $Fe(NO_3)_2(aq) + Al(s) \rightarrow Al(NO_3)_3(aq) + Fe(s)$

III $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

a Write the half-equations for the reactions.

b Write a balanced overall redox equation, excluding spectator ions.

c Identify the half-equations as oxidation or reduction.

d Identify the oxidising agent and the reducing agent.

23 For each pair of species a–c, draw a galvanic cell diagram that includes:

- the anode and cathode
- the material that each electrode is made of
- the direction of electron movement
- a suitable salt bridge solution
- the direction of ion movement in the salt bridge
- the direction of ion movement in each half cell
- reduction and oxidation half-equations
- the balanced overall redox equation.

a Au^+/Au and Mg^{2+}/Mg

b Sn^{2+}/Sn and Ni^{2+}/Ni

c Cd^{2+}/Cd and Mn^{2+}/Mn

Design and discuss

24 Design a mind map that demonstrates the connections between the following terms.

- Oxidation
- Reduction
- Gain in electrons
- Loss of electrons
- Oxidising agent
- Reducing agent

25 For the following mixtures, discuss which species will undergo oxidation and reduction using the reactivity series of metals and justify your responses.

a $Cu^{2+}(aq)$, $Au^+(aq)$, Zn(s) and Fe(s)

b $H^+(aq)$, $Al^{3+}(aq)$, K(s) and Li(s)

26 A student is required to determine the reactivity of metals by using half-cells, where two can be combined to form a galvanic cell. The metals used are Sn^{2+}/Sn, Cd^{2+}/Cd and Mg^{2+}/Mg.

Discuss how the student could determine which half-cell would undergo oxidation and reduction when they are paired, and how the student could determine the reactivity of the metals. Discuss the expected results, using the reactivity series of metals.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 2

AREA OF STUDY 1

Checkpoint

Multiple choice

Question 1

Select the reaction that involves the transfer of electrons.

A $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$

B $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

C $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

D None of the above

Question 2

The pH of 0.0050 M calcium hydroxide solution that has completely dissociated in water is:

A 12.0

B 11.7

C 2.3

D 2.0

Question 3

When zinc metal is added to a solution of copper(II) sulfate, the zinc becomes coated with copper metal. This is because:

A zinc metal is an oxidising agent, and $Cu^{2+}(aq)$ is a reducing agent.

B zinc metal is a stronger oxidising agent than $Cu^{2+}(aq)$.

C zinc metal is a reducing agent, and $Cu^{2+}(aq)$ is an oxidising agent.

D zinc metal is reduced, and $Cu^{2+}(aq)$ is oxidised.

Question 4

Which of the following properties of water is most important for cooling when people perspire (or sweat)? Water:

A has a relatively high boiling point.

B is polar, making it a good solvent.

C has a high latent heat of vaporisation.

D has a high latent heat of fusion.

Question 5

The pH of 0.1 M HCl is about 1.0 but the pH of 0.1 M arsenic acid (H_3AsO_4) is about 1.5. This is because:

A arsenic acid is amphiprotic in water.

B arsenic acid has more than one hydrogen which can react with a base.

C arsenic acid does not completely dissociate in water.

D arsenic is in the same group as nitrogen in the periodic table.

Question 6

Which of the following lists ranks the metals in order of decreasing reactivity?

A Sn, Fe, Mg

B Ni, Sn, Mg

C Zn, Fe, K

D Ni, Sn, Ag

Question 7

In which of the following acid–base reactions are the underlined species acting as a conjugate pair?

A $NH_3 + \underline{H_2O} \rightarrow \underline{NH_4}+ + OH^-$

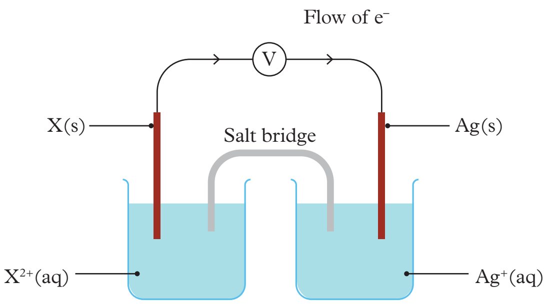
B $HCN + \underline{H_2O} \rightarrow H_3O^+ + \underline{CN}^-$

C $HNO_3 + \underline{H_2O} \rightarrow \underline{H_3O^+} + NO_3$

D $\underline{CH_3COOH} + H_2O \rightarrow \underline{CH_3COO}^- + H_2O$

Question 8

The following galvanic cell is formed by the combination of two half-cells: $X^{2+}(aq)/X(s)$ and $Ag^+(aq)/Ag(s)$.



Which of the following is correct?

A X is the anode, and the reaction is:
$X^{2+}(aq) + 2e^- \rightarrow X(s)$

B X is the cathode, and the reaction is:
$X^{2+}(aq) + 2e^- \rightarrow X(s)$

C X is the anode, and the reaction is:
$X(s) \rightarrow X^{2+}(aq) + 2e^-$

D X is the cathode, and the reaction is:
$X(s) \rightarrow X^{2+}(aq) + 2e^-$

Question 9

Which of the following species is both diprotic and amphiprotic?

A H_2CO_3

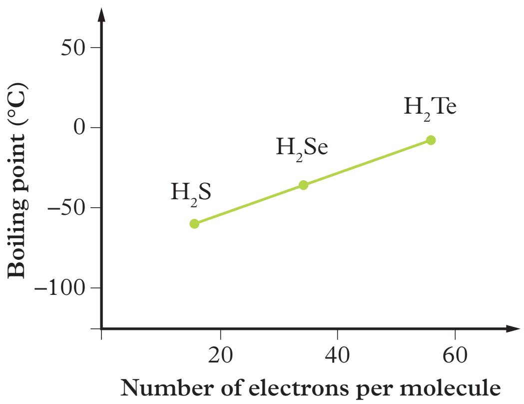
B $H_2PO_4^-$

C $CH_3CH_2COO^-$

D HS^-

Question 10

The following graph shows the variation in the boiling points of some group 16 hydrides.



It can be estimated from the graph that water (a group 16 hydride) should have a boiling point well below –50°C. Which one of the following statements best accounts for water's exceptionally high boiling point of 100°C?

A Water is a small molecule with very strong covalent bonds between the atoms.

B There are strong dispersion forces between the water molecules because they are packed tightly.

C The partial charges on the atoms of this polar molecule result in very strong intermolecular interactions.

D Water, unlike the other hydrides, exhibits hydrogen bonding between molecules.

Short answer

Question 1 (6 marks)

The results of an experiment using acid–base indicators showed the following effects on red and blue litmus.

Solution	Red litmus	Blue litmus
KNO_3	Red	Blue
K_3PO_4	Turns blue	Blue
NH_4NO_3	Red	Turns red

a Construct an equation to show an indicator (HIn) acting as a weak acid. 1 mark

b Classify each solution in the above table as acidic, basic or neutral. 3 marks

c Construct an equation to show potassium sulfide dissolving in water. 1 mark

d Construct an equation to show sulfide ions interacting with water. Would the solution be acidic or alkaline? 1 mark

Question 2 (3 marks)

Toxic gases released during a natural disaster in Cameroon in 1986 were suspected to be hydrogen sulfide, carbon monoxide and carbon dioxide.

Two reactions that can produce hydrogen sulfide in acidic soils are:

$FeS(s) + 2HCl \rightarrow H_2S(g) + FeCl_2(aq)$

$2CH_2O + H_2SO_4 \rightarrow H_2S(g) + 2CO_2(g) + 2H_2O(l)$

(Note: CH_2O is the empirical formula for carbohydrates.)

a Identify the equation that represents a redox reaction. 1 mark

b Identify the oxidising agent from the reaction and explain your choice. 2 marks

Question 3 (4 marks)

When 0.10 M solutions of sodium carbonate and ammonia were prepared and the pH was measured, it was found that sodium carbonate had a higher pH value.

a Explain this observation. 2 marks

b Deduce the hydroxide ion concentration of an aqueous solution with a pH of 9.0. 1 mark

c State the formula of the conjugate base of the ion HCO_3^-. 1 mark

Question 4 (4 marks)

Write balanced full and ionic equations (with states) for the reactions between:

a HNO_3 and $Ca(OH)_2$ solutions 2 marks

b Na_2CO_3 and H_2CrO_4 solutions. 2 marks

Question 5 (7 marks)

A galvanic cell is assembled from the $Cu^{2+}(aq)/Cu(s)$ and $Pb^{2+}(aq)/Pb(s)$ half-cells.

a Draw a diagram for this galvanic cell, labelling the:

- **i** polarity of the electrodes
- **ii** electrode that is the anode and the electrode that is the cathode
- **iii** direction of electron flow
- **iv** direction of anion (negative ion) flow.

4 marks

b For this galvanic cell, write the chemical half-equations for the reaction occurring at the:

- **i** anode
- **ii** cathode. 2 marks

c Construct the overall chemical equation for the cell when it is discharging. 1 mark

Question 6 (7 marks)

In a series of experiments involving the displacement of one metal ion from solution by another metal, the following results were recorded by a group of students.

Combination	Result
Copper nitrate(II) + lead	Reaction occurred
Copper nitrate(II) + zinc	Reaction occurred
Iron(II) sulfate + zinc	Reaction occurred
Lead(II) nitrate + copper	No reaction occurred
Lead(II) nitrate + iron	No reaction occurred
Lead(II) nitrate + tin	Reaction occurred
Magnesium sulfate + zinc	No reaction occurred
Tin(II) chloride + iron	Reaction occurred
Tin(II) chloride + magnesium	Reaction occurred
Zinc sulfate + lead	No reaction occurred
Zinc sulfate + tin	No reaction occurred

a Identify which of the following shows the correct order of reactivity of the five metal elements used in this series of experiments (with most reactive element first). Explain your reasoning.

1. Copper, lead, tin, zinc, magnesium
2. Magnesium, tin, zinc, copper, lead
3. Magnesium, zinc, tin, lead, copper
4. Tin, zinc, magnesium, lead, copper
5. Zinc, magnesium, lead, copper, tin

2 marks

b Which experimental result do you think is likely to be in error? Explain your reasoning. 2 marks

c For the second experiment of copper(II) nitrate and zinc shown in the table, construct:

- **i** a balanced chemical equation (with states) for the reaction
- **ii** equations for the half-reactions (with states).

3 marks

Question 7 (9 marks)

a As a result of increasing carbon dioxide levels in recent years, seawater can have a pH as low as 5.7.

i Construct a balanced equation to show water dissolving atmospheric carbon dioxide to produce carbonic acid.

ii Explain, using a balanced equation, how this acid can cause ocean pH to decrease.

Marine animals and corals have shells or skeletons that are made of calcium carbonate ($CaCO_3$).

iii Construct a balanced chemical equation to show how marine animals and corals build their shells or skeletons.

iv Explain, using balanced equations, the impact of ocean acidity on shell-building animals and corals.

5 marks

b In contrast to ocean acidification, acid rain in cities often results from a mixture of oxide pollutants containing nitrogen and sulfur.

i Construct balanced full equations for the reaction of oxide pollutants in water to form acid rain.

ii State one problem caused by acid rain.

4 marks

TOTAL MARKS

/50 marks

CHAPTER 14

Measuring solubility and concentration

KEY KNOWLEDGE

→ solution concentration as a measure of the quantity of solute dissolved in a given mass or volume of solution (mol L^{-1}, g L^{-1}, %(m/v), %(v/v), ppm), including unit conversions

→ the use of solubility tables and solubility graphs to predict experimental determination of ionic compound solubility; the effect of temperature on the solubility of a given solid, liquid or gases in water

→ the use of precipitation reactions to remove impurities from water

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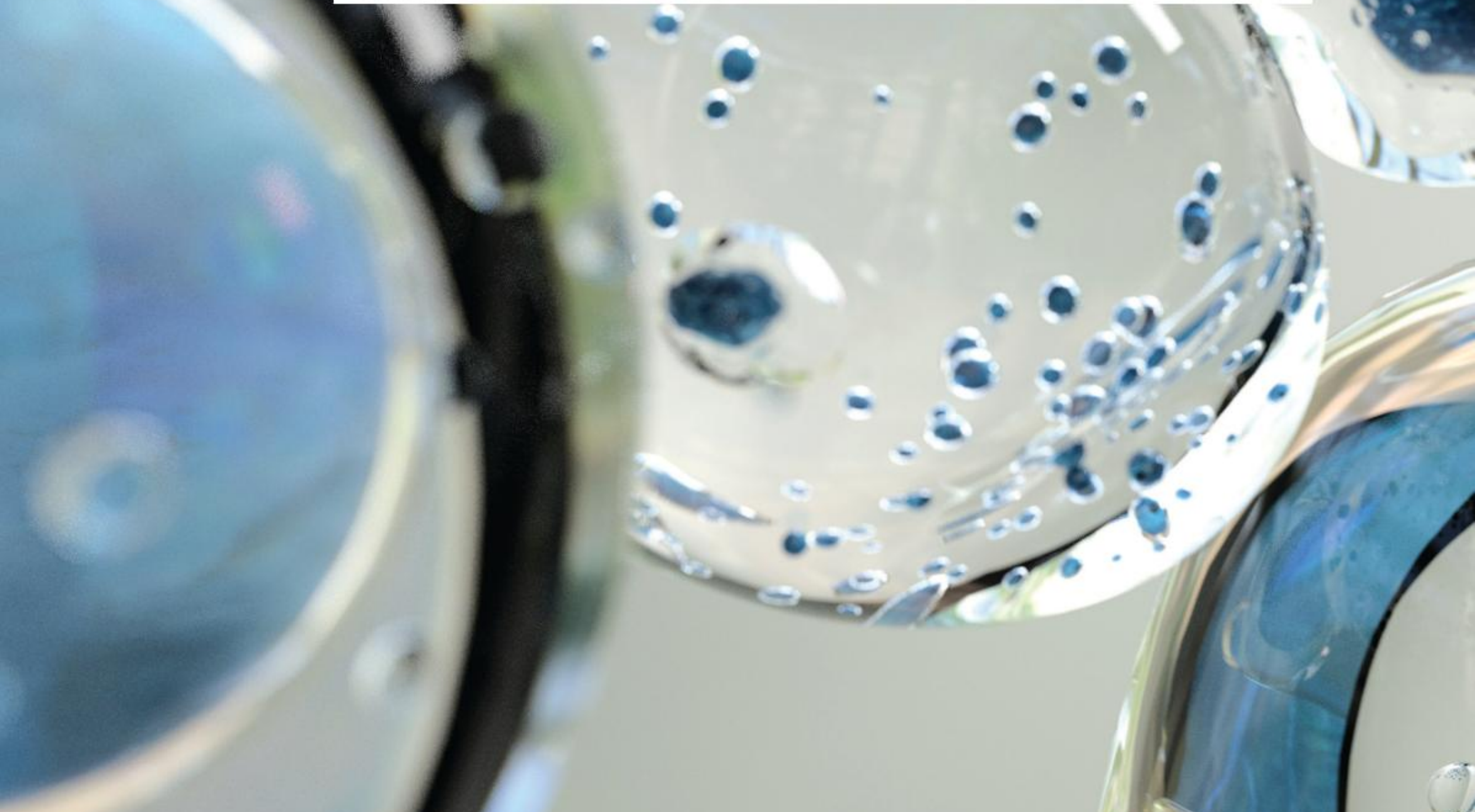

FIGURE 1 Water is a universal solvent that can dissolve more substances than any other solvent on Earth.

GROUNDWORK

In Chapter 14, you will learn about the different units you can use to represent solubility and how to predict the solubility of substances in water by using solubility tables and graphs. You will also learn about how to apply your understanding of precipitation reactions to the purification of water.

This chapter will build on concepts you have already learnt in Chapters 3, 5 and 6. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

14A Write a balanced equation for the reaction between solutions of calcium chloride and sodium carbonate.

14A Groundwork resource
Writing balanced equations

14B Calculate the amount (in mol) of potassium iodide in 5.3 g.

14B Groundwork resource
Moles

14C Explain how *like dissolves like* relates to intermolecular bonds.

14C Groundwork resource
Like dissolves like

14D How does intermolecular bonding affect the physical state of a chemical?

14D Groundwork resource
Physical state

PRACTICALS

14.2 PRACTICAL: CLASSIFICATION & IDENTIFICATION	Can you determine the solubility of molecules in water from their structure?	Page 519
14.3 PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT	How can you remove cations from hard water?	pro

14.1 Solution concentration

KEY IDEAS

In this topic, you will learn that:

+ concentration is a measure of the amount (moles), mass or volume of substance in a given volume of solution
+ concentration can be measured in many different units depending on the nature of the solution
+ molar concentration can be calculated if the moles and volume of the substance are known.

solution
an even mixture that forms when a solute dissolves in a solvent

degree of saturation
the amount of solute dissolved in a volume of solvent

unsaturated
a solution in which more solute can be dissolved

saturated
a solution in which the maximum amount of solute is dissolved

supersaturated
a solution in which more than the usual maximum amount of solute is dissolved

metastable
a highly energetic and unstable state; in a metastable supersaturated solution, the dissolved solute is likely to 'crash out' of solution and form crystals

crystal
solid material that forms

Water is considered the 'universal solvent' because it can dissolve more substances than any other chemical. This means that it can form many different **solutions**. The molecular shape of water is what gives it this property.

Because water is bent, a dipole is formed in which oxygen has a partial negative charge and each hydrogen has a partial positive charge.

This allows water molecules to form hydrogen bonds with neighbouring molecules and, therefore, dissolve them (Figure 1).

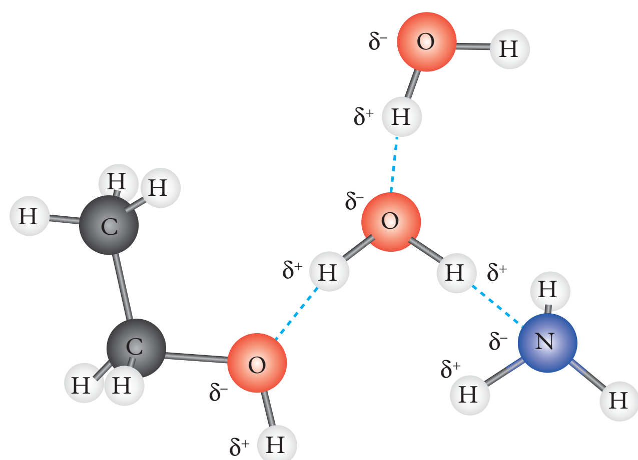

FIGURE 1 Water forms hydrogen bonds with ethanol, ammonia and other water molecules.

Solutions

A solution is a mixture that forms when a solute dissolves in a solvent. To prepare a solution, a sample of solute is measured (either by weighing on a balance or measuring a volume) and then added to a volume of solvent. The solution can be stirred or heated to make the solute dissolve faster.

Depending on the amount of solute in a solution, the **degree of saturation** changes (Figure 2):

- If all the solute dissolves and it is possible to dissolve more, the solution is **unsaturated**.
- If no more solute will dissolve, the solution is **saturated**.
- If a saturated solution is heated, more solute often dissolves. If the solute remains in solution as it cools, this is a **supersaturated** solution. At this point, the solution is considered to be **metastable** and the solute can 'crash out' of the solution to form **crystals**.

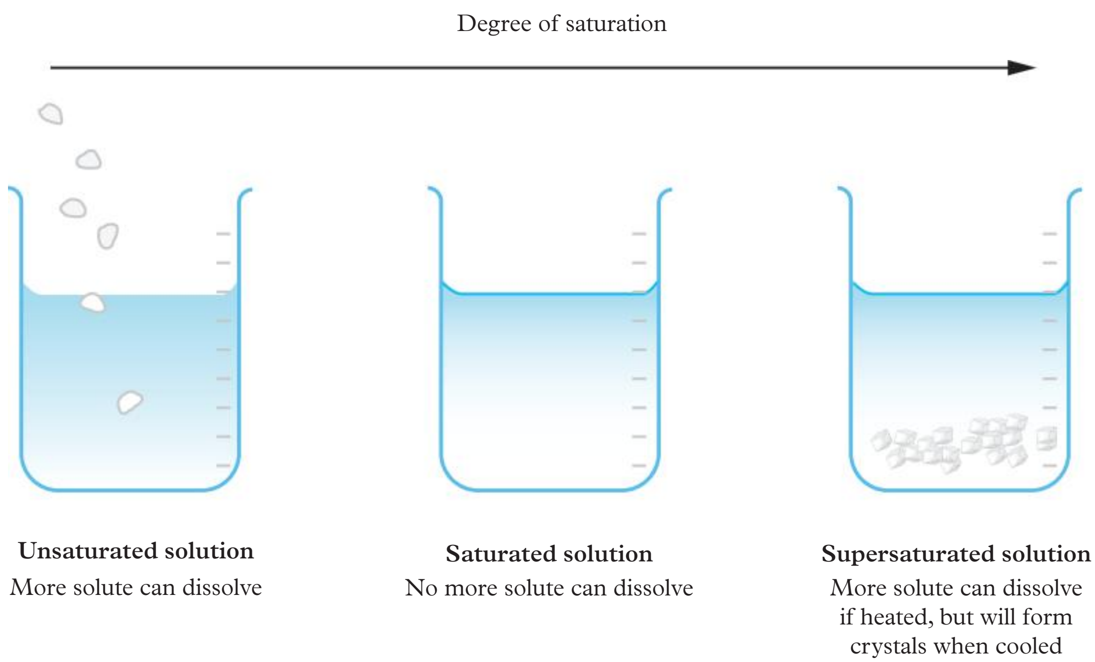

FIGURE 2 The degree of saturation increases as more solute is dissolved.

A solution in which water acts as the solvent is called an **aqueous solution** (from the word *aqua*, meaning 'water'). It is given its own state in chemical equations, represented as (aq). Analyse and evaluate data about an aqueous solution of copper(II) sulfate in Skill drill 14.1.

14.1 Skill drill
Find me in your obook pro

aqueous solution
the type of solution formed when water is used as the solvent

Concentration

The amount of solute that has dissolved in an accurately known mass or volume of solvent to form a solution can be described by referring to its concentration. For a specific volume, a concentrated solution contains a higher amount of solute, whereas a dilute solution has a lower amount of solute. Concentration is measured in many ways, depending on the nature of the solute and the solvent.

Units of concentration

Concentration can be described in a range of units, including:

- molarity (mol L^{-1} or M)
- grams per litre (g L^{-1})
- parts per million (1 ppm = 1 mg L^{-1})
- percentage mass per volume (% m/v)
- percentage volume per volume (% v/v).

Molarity (mol L^{-1} or M)

Molarity, or molar concentration, is the most common unit of concentration used to measure the amount (in mol) of a substance per litre of solvent. It is represented as moles per litre, mol L^{-1} or M. When calculating concentration, either mol L^{-1} or M can be used as units.

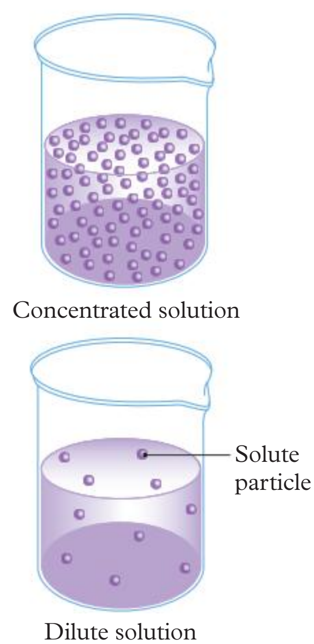

FIGURE 3 The same volume of a concentrated solution contains a higher amount of solute than a dilute solution.

molarity (mol L^{-1} or M)
a unit of concentration measured by the number of moles per litre of solution; also known as molar concentration

Study tip

When doing multistep calculations, highlight the important information (numbers, units, chemicals or key words), then add the letters that represent them above the values. Put a question mark next to the value you are asked to calculate. Look at all of the formulas and chemical equations, and find a pathway that will allow you to calculate the answer.

Concentration in moles per litre is calculated by using the formula:

$$c = \frac{n}{V}$$

where c is concentration (mol L^{-1} or M), n is number of moles (mol) and V is volume (L). If volume is given in millilitres, then it must be converted to litres first.

Concentration can also be represented by using square brackets around the formula of the substance. For example, if a potassium iodide solution has a concentration of 0.8 M, you can express this as [KI] = 0.8 M.

Often, the concentration of ions in a solution must be determined. This is essential when calculating the pH of a solution (covered in Chapter 12). An extra step must be added where the amount (in mol) of the ion is determined from the amount (in mol) of the solute. This requires a mole ratio statement. For example, Na_3PO_4 contains three Na^+ ions. Therefore, the amount (in mol) of Na^+ is three times the amount (in mol) of Na_3PO_4. The amount of Na^+ is then used to calculate the concentration of the Na^+ ions in the solution.

Study tip

For questions that ask you to calculate the mass of solute required to make a solution of known concentration, you can rearrange the concentration formula to find mass (m).

Study tip

Always convert volume into litres when performing concentration calculations.

14.1A WORKED EXAMPLE

CALCULATING THE MOLARITY OF A SOLUTION USING MOLES

A 200 mL solution contains 1.2 mol of calcium chloride. Calculate its concentration. Then, calculate the concentration of chloride ions in solution.

Solution

Think	Do
Step 1: Convert volume to litres.	200 mL ÷ 1000 = 0.200 L
Step 2: Rearrange the concentration formula depending on the unknown value.	$c = \frac{n}{V}$ Does not need to be rearranged because you want to calculate the concentration (c).
Step 3: Use the formula to calculate the unknown value.	$[CaCl_2] = \frac{1.2}{0.200} = 6.0$ M (2 sig fig)
Step 4: Look at how many chloride ions are in calcium chloride. Multiply the number of chloride ions by the mol of calcium chloride to find the mol of chloride ions.	There are two chloride ions. $n(Cl^-) = 2 \times n(CaCl_2) = 2.4$ mol
Step 5: Use the formula to calculate the concentration of chloride ions.	$[Cl^-] = \frac{n}{V}\ \frac{2.4}{0.200} = 12$ M (2 sig fig)

14.1B Worked example
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14.1B Worked example
Video demonstration

In a laboratory, it is more realistic to weigh a mass of a chemical and dissolve it in water, then calculate its concentration. This calculation is a multistep process where moles must be calculated first, followed by concentration. See how this is done in Worked example 14.1B.

Grams per litre (g L^{-1})

g L^{-1}
a unit of concentration measured in the mass of solute (g) per litre of solution

Another common unit of concentration is grams of substance per litre. There are several reasons for using **g L^{-1}** instead of mol L^{-1}. For example, you may not know the molar mass of a compound if it is a new compound and still being investigated, or the substance may be impure or a mixture.

Concentration in grams per litre is calculated by using the formula:

$$c = \frac{m}{V}$$

where c is concentration (g L^{-1}), m is mass (g) and V is volume (L). This is the same calculation you used previously, but amount (in mol) is replaced with mass (in grams). Worked example 14.1C shows you how to calculate the concentration of a solution in g L^{-1}.

14.1C Worked example
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14.1C Worked example
Video demonstration

Parts per million (ppm or mg L^{-1})

The unit **parts per million** (ppm) is used to measure very low concentrations. Let's use the density of water as an example.

The density of water is 1.00 g mL^{-1} at room temperature. It helps to convert the units so that they are the same. In this case, we can consider them both in units of mass.

- This means that 1 mL of water has a mass of 1.00 g.
- Because there are 1000 mL of water in a litre, 1 L of water has a mass of 1.00 kg.
- To determine how many parts per million this is, we take one-millionth of 1 kg. This is 1 mg.
- Because 1 L of water = 1 kg of water, one-millionth of a litre is 1 mg.
- So, 1 ppm is equivalent to 1 mg per litre of water or 1 mg L^{-1}.

Concentration in ppm is calculated by using the formula:

$$c = \frac{m}{V}$$

where c is concentration (mg L^{-1} or ppm), m is mass (mg) and V is volume (L). This is also the same calculation you used previously, but mass is now in milligrams.

parts per million (ppm) or mg L^{-1}
a unit of concentration measured in the mass of solute (mg) per litre of solution

Study tip

Make sure you are confident using the $c = \frac{n}{V}$ and $c = \frac{m}{V}$ equations. Concentration calculations are very important in VCE Chemistry.

14.1D WORKED EXAMPLE

CALCULATING THE CONCENTRATION OF A SOLUTION IN ppm

125 mL of a sodium chloride solution was made from 0.0024 g of the salt. Calculate its concentration in ppm.

Solution

Think	Do
Step 1: Convert volume to litres.	125 mL ÷ 1000 = 0.125 L
Step 2: Convert mass to mg.	0.0024 × 1000 = 2.4 mg
Step 3: Rearrange the concentration formula depending on the unknown value.	$c = \frac{m}{V}$ Does not need to be rearranged because you want to calculate the concentration (c).
Step 4: Use the formula to calculate the unknown value.	$c(NaCl) = \frac{2.4}{0.125}$ = 19 mg L^{-1} = 19 ppm (2 sig fig)

Percentage mass per volume (% m/v)

The concentration of a solution can also be expressed as the mass of solute (in g) per 100 mL of solution. This is called percentage mass per volume or **% m/v**.

Percentage mass per volume is calculated by using the formula:

$$\% \text{ m/v} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$$

% m/v
a unit of concentration measured in the mass of solute (g) per 100 mL of solution

Study tip

You may see percentage mass per volume expressed as % w/v instead of % m/v. This is percentage weight per volume. They are the same thing.

When working with percentage mass per volume (and percentage volume per volume, which you will learn about next), you may have to convert between units. You can use the unit conversion tool in Figure 4 to help you. For percentage mass per volume, it is handy to convert the concentration units to g L^{-1} first.

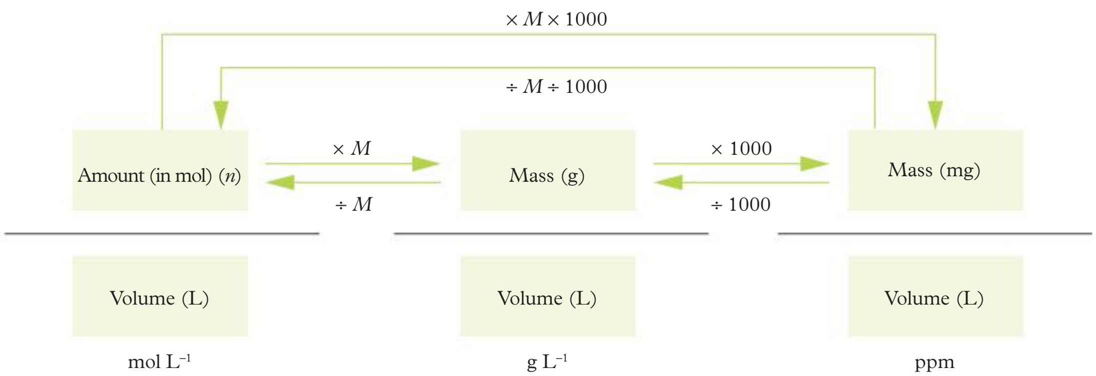


FIGURE 4 A conversion tool for concentration units mol L^{-1}, g L^{-1} and ppm, where M is molar mass

14.1E Worked example
Video demonstration

Study tip

Units are very important for these calculations. Always double check that you're using the right ones; otherwise, your answer could be off by a factor of 1000.

14.1F Worked example
Find me in your obook pro

14.1F Worked example
Video demonstration

14.1E WORKED EXAMPLE

CONVERTING BETWEEN UNITS OF CONCENTRATION

Convert 7.6 ppm LiOH to:

a g L^{-1}

b M

Solution

Think	Do
Step 1: Remember that ppm is the same as mg L^{-1}.	7.6 ppm is the same as 7.6 mg L^{-1}.
Step 2: Convert the mass to grams.	**a** 7.6 mg ÷ 1000 = 0.0076 g 0.0076 g L^{-1} (2 sig fig)
Step 3: Remember that M is mol L^{-1}. Calculate the number of moles of LiOH.	$n = \frac{m}{M}$ $n(\text{LiOH}) = \frac{0.0076}{23.9}$ $= 0.000318$ mol in 1 L **b** 0.00032 M or 3.2×10^{-4} M (2 sig fig)

Percentage volume per volume (% v/v)

If the solute that has been dissolved into a solution is a liquid, it is more accurate to calculate the volume of the solute (in mL) per 100 mL of solution. Often this is used for fuels, chemical stock solutions or drinks. This is called percentage volume per volume or **% v/v**.

Percentage volume per volume can be calculated using the formula:

$$\% \text{ v/v} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$$

% v/v
a unit of concentration measured in the volume of solute (mL) per 100 mL of solution

14.1G WORKED EXAMPLE

CALCULATING THE VOLUME OF A SOLUTE USING % v/v

The manufacturer of an alcoholic drink claims that it contains 4.90% v/v alcohol (ethanol). Calculate the volume of ethanol in a standard drink of 285 mL.

Solution

Think	Do
Step 1: Rearrange the concentration formula depending on the unknown value.	You want to calculate volume of the solute, so $\% \text{ v/v} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$ becomes $\text{volume of solute (mL)} = \frac{\% \text{ v/v}}{100} \times \text{volume of solution (mL)}$
Step 2: Use the formula to calculate the volume of the solute.	$\text{volume of solute (mL)} = \frac{4.90}{100} \times 285 = 14.0 \text{ mL}$ (3 sig fig)

14.1 Challenge
Find me in your obook pro

Try Challenge 14.1 to see if you can calculate concentration using density.

You can also compare the concentration units in nutrition information panels in Real-world chemistry 14.1.

14.1 Real-world chemistry
Find me in your obook pro

14.1 CHECK YOUR LEARNING

Describe and explain

1 Describe when g L^{-1} and ppm might be used instead of mol L^{-1}.

2 Explain why % v/v is used in alcoholic drinks rather than mol L^{-1}.

Apply, analyse and compare

3 Determine the concentration (mol L^{-1}) of:
   a a solution containing 0.3 mol of NaCl in 50 mL of water
   b iodide ions when 0.9 g of KI was dissolved in 20 mL of water
   c a solution containing 4.93×10^{21} HCl molecules in 125 mL of water.

4 Determine the concentration (in ppm) of:
   a 0.0032 mol of HNO_3 in 345 mL of water
   b 0.093 g of NaOH in 100 mL of water
   c 4.32×10^{15} molecules of H_2SO_4 in 150 mL of water.

5 Calculate the:
   a volume of ethanol required to make 50.0 mL of a 20% v/v drink
   b mass of $AgNO_3$ in a 25.00 mL bottle of 0.5 M solution
   c volume of water required to make a 0.5 M solution of $CuSO_4$ if 7.98 g of copper sulfate is available (assuming the whole mass is used)
   d mass of octane (density of 0.703 g mL^{-1}) in a 40 L tank of octane
   e molar concentration (M) of a bottle of fuming HCl which has a concentration of 37% v/v and a density of 1.2 g mL^{-1}.

6 Convert the following units.
   a 0.056 M of $K_2Cr_2O_7$ to ppm
   b 25 ppm KOH to g L^{-1}
   c 82 g L^{-1} of H_3PO_4 to M
   d 0.58 mg L^{-1} to % m/v

14.2 Solubility tables and graphs

KEY IDEAS

In this topic, you will learn that:

+ ionic and polar covalent substances can be soluble in water
+ the solubility of ionic substances can be predicted using a solubility table
+ the solubilities of solids, liquids and gases at different temperatures can be predicted using solubility graphs.

soluble
will dissolve

insoluble
will not dissolve

Solubility is the ability of a solute (solid, liquid or gas) to dissolve in a solvent. If it dissolves, the substance is **soluble**. If it does not dissolve, it is **insoluble**.

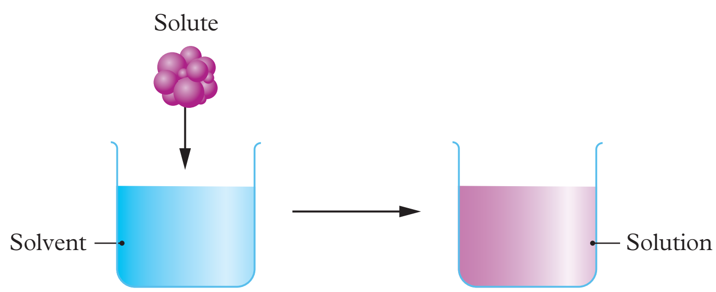


FIGURE 1 A solute dissolves in a solvent to form a solution.

Study tip

Refresh your understanding of intermolecular forces by going back to Chapter 3.

Study tip

Remember *like dissolves like!*

Solubility of substances in water

At a given temperature, the solubility of a substance in water depends on the:

1. breaking of solute–solute and solvent–solvent bonds
2. formation of solute–solvent bonds.

Do ionic compounds dissolve?

In Chapter 5, you learnt that an ion is an atom or molecule that has gained electrons to become a negatively charged anion or lost electrons to become a positively charged cation (Figure 2). An ionic bond is an electrostatic attraction between a cation and an anion. Because this involves whole charges, ionic bonding is stronger than the dipole–dipole attractions formed between partially charged polar covalent molecules.

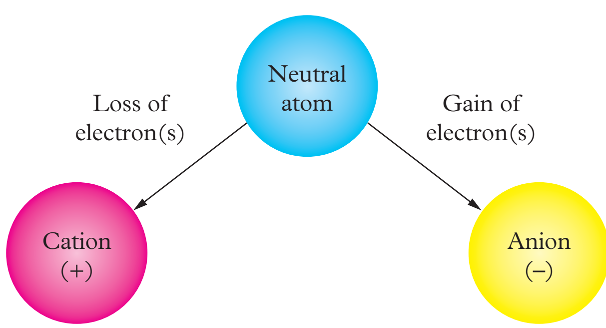


FIGURE 2 A neutral atom loses one or more electrons to become a cation or gains one or more electrons to become an anion.

14.2
PRACTICAL: CLASSIFICATION & IDENTIFICATION

Can you determine the solubility of molecules in water from their structure? Go to page **519**.

For an ionic compound to dissolve (or be soluble) in water, it must form ion–dipole attractions with the water, which replace any existing bonds that it has between cations and anions. Let's look at an example. When sodium chloride (NaCl) is added to water, the:

1. ionic bonds between Na^+ and Cl^- in solid NaCl dissociate (break apart)
2. hydrogen bonds between water molecules break apart
3. Na^+ becomes attracted to the partially negative oxygen and Cl^- becomes attracted to the partially positive hydrogens of water molecules
4. ions become hydrated and form **ion–dipole attractions** with water (Figure 3).

ion–dipole attraction
the interaction between an ion with a whole charge and the partial charge caused by a polar bond or molecule

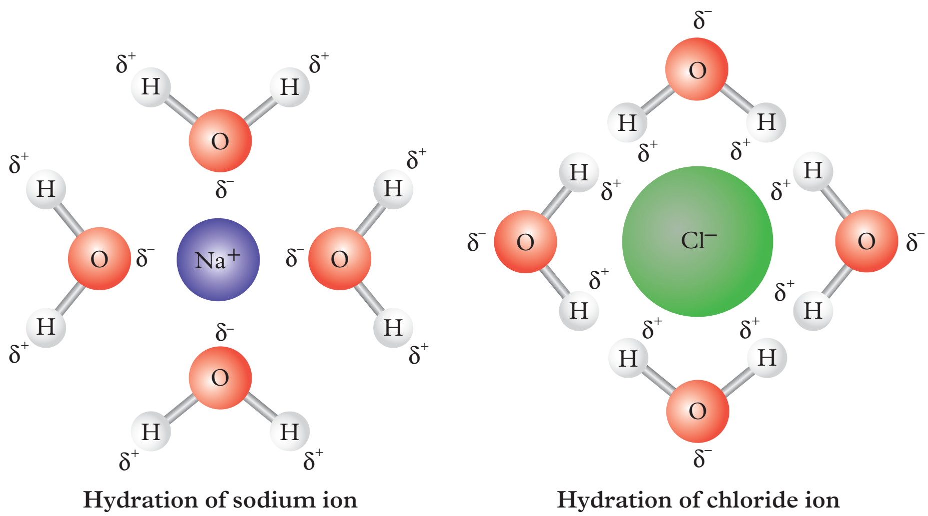


FIGURE 3 The Na^+ cation and Cl^- anion are attracted to water molecules and become hydrated.

ionic substance
a cation (metal) interacting with an anion (non-metal)

The opposite charges are the reason ionic bonds or interactions are so strong. However, a common misconception is that because **ionic substances** have whole positive and negative charges, the ionic compound will not dissociate to form ion–dipole attractions with the partially charged solvent molecules. This is true if there is only one water molecule interacting with one unit of the compound but, in reality, there is not. Many water molecules will interact with the cations and anions in the ionic compound. The sum of all the partial charges attached to many water molecules provides them with a very strong attractive force to pull the ionic substance apart (Figure 5). This process is called dissolving.

FIGURE 4 Did the Wicked Witch of the West melt or dissolve?

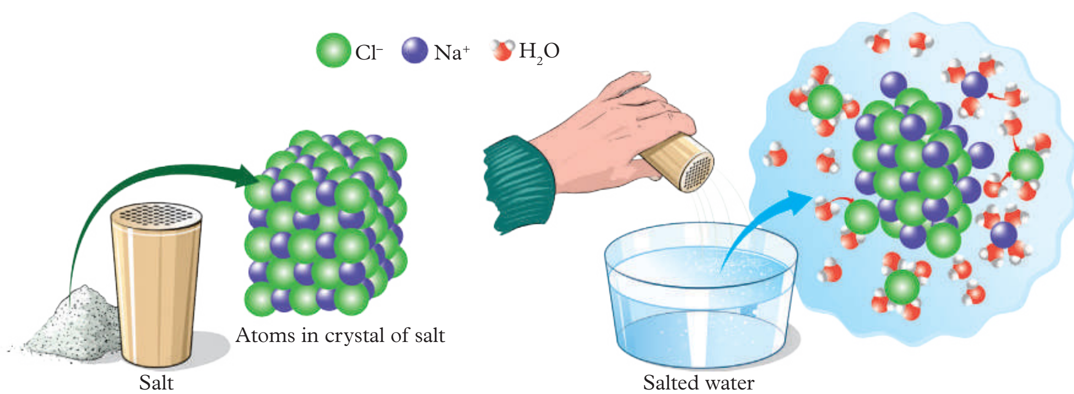


FIGURE 5 Water molecules dissociate anions from cations when they interact.

Study tip

Dissolving and melting are not the same! Dissolving happens when one substance mixes evenly with another to form a mixture. Melting is when one substance changes from solid to liquid form. So, when water was poured onto the Wicked Witch of the West, did she melt or *dissolve?*

covalent molecular substance
a molecule that is covalently bonded because of the sharing of valence electrons

Do covalent molecular substances dissolve?

Covalent molecular substances can also dissolve in water if they are polar. The permanent dipoles within the molecules form intermolecular forces with the solvent molecules. For example, the partially positive hydrogen of water interacts with the partially negative oxygen in ethanol to form a hydrogen bond (Figure 6).

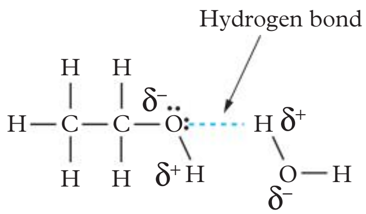


FIGURE 6 The hydrogen bonding interaction between ethanol and water

Non-polar molecular substances can only interact with water solvent molecules through weaker dispersion forces, so they have a much lower solubility in water.

The interaction between molecular substances and water is not as strong as the interaction between ionic substances and water. This is because ionic substances have whole charges on their ions, whereas molecular substances are only partially positive and partially negative. This means that, generally, ionic substances have a greater solubility than molecular substances do.

Solubility tables

14.2A Worked example
Find me in your obook pro

14.2A Worked example
Video demonstration

Remember from Chapter 5, a solubility table (Table 1) is a set of rules that provide a guideline for the behaviour of ionic substances at 25°C and 1 atm. Specifically, it helps you tell whether an ionic compound is soluble or insoluble. Soluble salts exist in an **aqueous** (aq) state. Insoluble salts are solids, indicated by (s). Take a look at Worked example 14.2A if you need a refresher on how to use a solubility table.

aqueous
the state given to a substance that is dissolved in water, (aq)

TABLE 1 A solubility table with simple rules for solubility of salts in water

Rule	Soluble	Exceptions
1	Nitrates (NO_3^-)	None
2	Group 1 ions: Li^+, Na^+, K^+	None
3	Ammonium (NH_4^+) salts	None
4	Ethanoates (CH_3COO^-)	None
5	Most halide salts: chlorides (Cl^-), bromides (Br^-) and iodides (I^-)	Ag^+, Pb^{2+} and Hg^{2+} salts
6	Most sulfate (SO_4^{2-}) salts	Ag^+, Ba^{2+}, Ca^{2+}, Pb^{2+}, Hg^{2+} and Sr^{2+} salts
	Insoluble	**Exceptions**
7	Most hydroxides (OH^-)	Li^+, Na^+, K^+ and NH_4^+ salts
8	Most sulfides (S^{2-}), carbonates (CO_3^{2-}) and phosphates (PO_4^{3-})	Li^+, Na^+, K^+ and NH_4^+ salts

Study tip

The SNAPE rule states that ionic substances that contain one or more of the following substances will always be soluble.

- **S**odium (Na^+)
- **N**itrate (NO_3^-)
- **A**mmonium (NH_4^+)
- **P**otassium (K^+)
- **E**thanoate (CH_3COO^-)

Slightly soluble substances

Some ionic compounds are **slightly soluble**. For example, some ions are so strongly attracted to one another (i.e. the attraction between the anion and cation is so great) that they are less likely to dissociate or dissolve. This means that not all of the substance will be in solution, and some will remain as a solid.

slightly soluble
only partially dissolves in a solvent

Insoluble substances

Some ionic compounds are insoluble. This means that they do not dissolve in water and therefore remain as a solid. This is because the attractive forces between the anions and cations are much stronger than any interactions the ions may have with water. As seen in the solubility table, most hydroxides, sulfides, carbonates and phosphates are insoluble in water.

Temperature and solubility graphs

The solubility of a substance is measured as the maximum amount that will dissolve in a known quantity of solvent at a given temperature. Once the maximum amount of solute has been reached and no more can dissolve, the solution becomes saturated.

If the temperature of a solution is increased, the solubility of a substance changes. An increase in temperature causes an increase in kinetic energy. This energy either causes substances to become more or less soluble in water, depending on:

- polarity – polar or non-polar
- state – solid, liquid or gas.

A **solubility graph or curve** shows the relationship between temperature and the solubility of substances. It is a graphical representation of the mass of solute that can dissolve in 100 g of water (y-axis) at a certain temperature (x-axis). Some solubility graphs for different substances are shown in Figure 7.

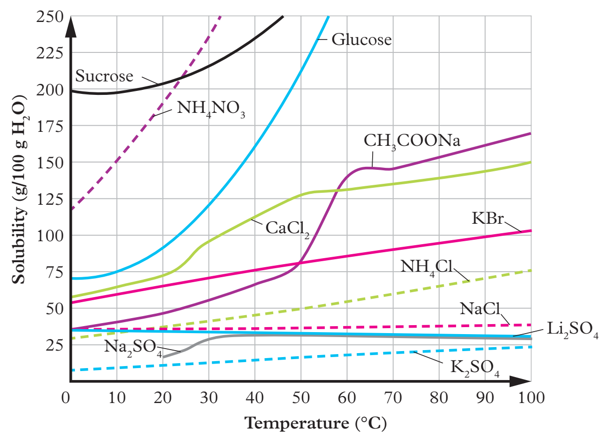


FIGURE 7 The solubility graphs of some substances in water

A solubility graph provides several key pieces of information.

- The curve represents the **saturation point** of a substance at a given temperature.
- A solution is unsaturated if its concentration sits below the curve.
- A solution is supersaturated if its concentration sits above the curve.
- A positive trend indicates that solubility increases as temperature increases.
- A negative trend indicates that solubility decreases as temperature increases.

Worked example 14.2B shows you how to calculate solubility if the values are provided to you. Worked example 14.2C shows you how to predict solubility using a solubility graph.

Study tip

You may encounter solubility tables that look slightly different. What's important is you remember the rules and their important exceptions.

solubility graph or curve
a graph that shows the relationship between the solubility of substances and temperature

saturation point
the point at which the maximum amount of solute has dissolved in the known quantity of solute at a given temperature

14.2B WORKED EXAMPLE

CALCULATING THE SOLUBILITY OF A SUBSTANCE

At 21°C, a maximum of 4.2 g of solute dissolves in 25 g of water. What is its solubility in g/100 g of water?

Solution

Think	Do
Step 1: If there is 4.2 g in 25 g of water, determine the mass in 100 g of water.	In 100 g of water, there would be: $\frac{100}{25}$ = 4 lots of 4.2 g of solute Therefore, 4 × 4.2 g = 16.8 g of solute. The solute has a solubility of 17 g/100 g water at 21°C (2 sig fig).

Study tip

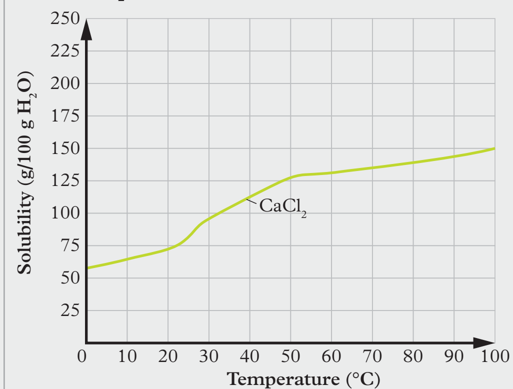
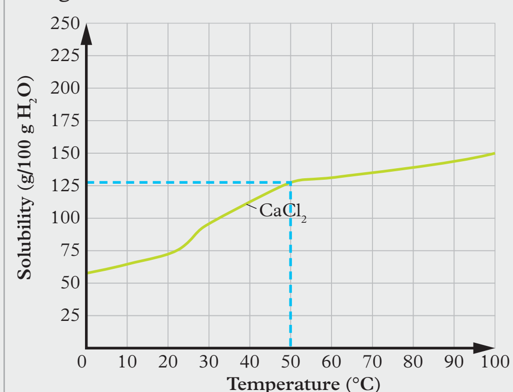
On the internet, you may come across solubility values for substances in other solvents, such as dimethyl sulfoxide (DMSO) and ethanol. In VCE Chemistry, you will only need to deal with solubility of substances in water.

14.2C WORKED EXAMPLE

USING THE SOLUBILITY GRAPH TO PREDICT SOLUBILITY

At 50°C, 140 g of solid $CaCl_2$ solute was added to 100 g of water. Will all of the solute dissolve?

Solution

<table>
<tr><th>Think</th><th>Do</th></tr>
<tr><td>Step 1: Look for the $CaCl_2$ line on the solubility graph in Figure 7.</td><td>The $CaCl_2$ line is green.

</td></tr>
<tr><td>Step 2: Find 50°C on the x-axis and read the solubility (g/100 g H_2O).</td><td>At 50°C, the solubility of $CaCl_2$ is a bit more than 125 g in 100 g of water.

</td></tr>
<tr><td>Step 3: Compare the solubility with the amount of solid solute added.</td><td>140 g was added, which is more than the amount soluble in the 100 g of water. Therefore, not all of the solute will dissolve.</td></tr>
</table>

Effect of temperature on the solubility of gases

Atmospheric gases, such as nitrogen (N_2), oxygen (O_2) and carbon dioxide (CO_2), have very low solubilities in water. As temperature increases and kinetic energy increases, the intermolecular forces between dissolved gas particles and the water solvent (solute–solvent bonds) are broken. The increase in kinetic energy is so significant that the solubility of all gases decreases when temperature increases, regardless of whether they are polar or non-polar (Figure 8).

When writing equations for dissolved gases, their state is aqueous (aq). When present in the atmosphere, their state is gaseous (g).

Practise your science skills on some gas solubility data in Skill drill 14.2. If you're up for a challenge, compare the effect of temperature on gas and ionic salt solubility in Challenge 14.2.

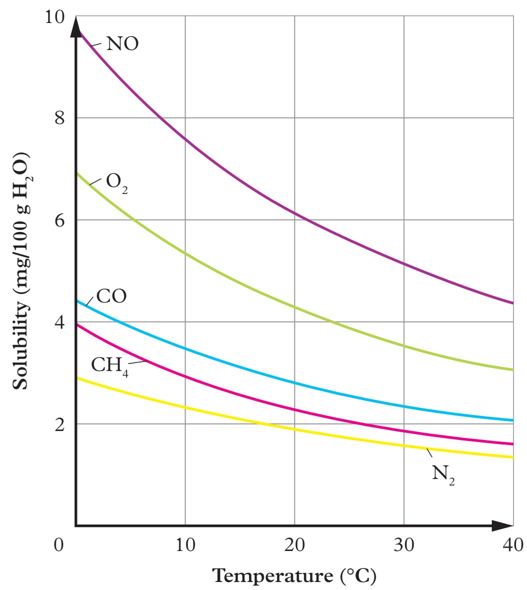


FIGURE 8 The solubility of gases in water decreases as temperature increases.

14.2 SKILL DRILL

Presenting, analysing and evaluating gas solubility data

Key science skills: Analyse and evaluate data and investigation methods

To determine the solubility of carbon dioxide, a student uses a soda maker. They start by measuring the mass of a 1 L bottle and then filling it with water and re-weighing it to determine the mass of the water in the bottle. They then use the soda maker to add carbon dioxide to the water until no more will dissolve and they re-weigh the bottle. This experiment is repeated over several days, each of which had a different maximum temperature.

Practise your skills

1. Construct a solubility graph from the data. Make sure to include the following elements: title, labels, units, legend (if applicable).
2. Describe the trend in the data and use the theory of solubility to explain this trend.
3. Identify a possible error in this experiment and discuss the impact it may have on the results.
4. Evaluate the accuracy, precision and reliability of the data.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

Temperature of bottle (°C)	Mass of bottle (g)	Initial mass of bottle and water (g)	Final mass of carbonated water (g)
18	145	1133.7	1134.7
25	145	1138.2	1139.0
30	145	1135.9	1136.6

14.2 CHALLENGE

Why does the solubility of covalent gases decrease as temperature increases, but the solubility of ionic compounds increases as temperature increases?

14.2 CHECK YOUR LEARNING

Describe and explain

1 Explain why most substances are more soluble as temperature increases.

2 Explain why a substance may be insoluble in water.

Apply, analyse and compare

3 Order these molecules by their solubility in water, from highest to lowest: non-polar molecular substance, ionic substance and polar molecular substance. Justify your answer by referring to the intermolecular forces.

4 Determine whether the following ionic substances are soluble in water and identify their state.

   a $NaNO_3$

   b $Pb(OH)_2$

   c Hg_2Cl_2

   d KCl

   e $BaSO_4$

5 Compare solubility tables and solubility graphs.

6 Use the solubility graph in Figure 9 to answer the following questions.

   a Identify the solubility of NH_4Cl at 70°C.

   b Identify the solubility of $NaNO_3$ at 50°C.

   c Identify the lowest possible temperature required to create a saturated solution containing 30 g of $KClO_3$ in 100 g of water.

   d Determine the solubility of $NaNO_3$ in 30 g of water at 40°C.

   e Determine the solubility of KI in 2.5 g of water at 10°C.

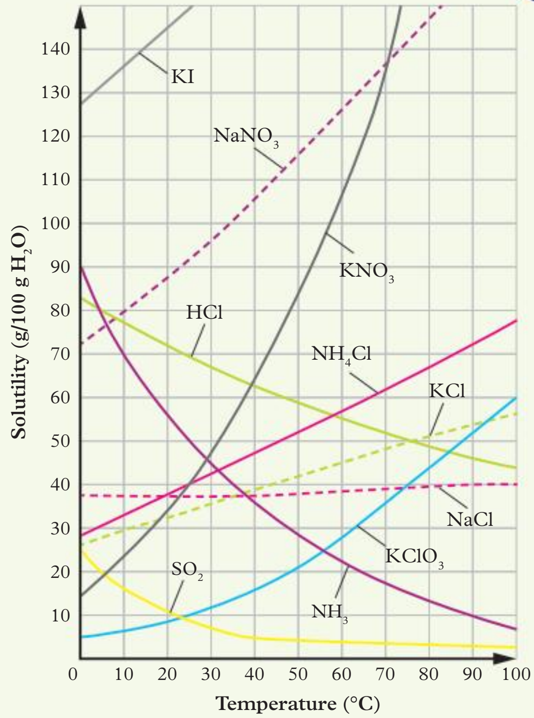


FIGURE 9 The solubilities of some ionic and molecular (covalent) substances in water

Design and discuss

7 Use the solubility graph in Figure 9 to describe how a student could make a supersaturated solution of NH_4Cl.

8 Ammonia (NH_3), sulfur dioxide (SO_2) and hydrogen sulfide (H_2S) are polar molecules that are gases at room temperature. Discuss whether their solubility graphs would differ from that of a non-polar molecule.

9 The solubility of KCl at 76°C is the same as the solubility of KNO_3 at 32°C (50 g/100 g of H_2O). Discuss why this may be.

14.3 Precipitation reactions

KEY IDEAS

In this topic, you will learn that:

+ precipitation reactions involve the formation of a low solubility solid
+ water can be purified by precipitating ions out of solution.

Study tip

In ionic substances, always write the cation first and then the anion.

14.3A Worked example
Find me in your obook pro

14.3A Worked example
Video demonstration

Some reactions between ionic substances form products with low solubility. If the new solute–solute bonds are too strong for them to interact with water, they do not dissolve and will instead form solids. In Chapter 5, you learnt how to predict and identify precipitation reactions between ions in solution. Your knowledge about precipitation reactions can also be applied to the removal of impurities from water.

Precipitation reactions

You might remember from Chapter 5 that a precipitation reaction occurs when two aqueous ionic solutions are mixed and a precipitate (solid) is formed. During this reaction, the cation and anion swap partners to produce new substances.

$$AB + CD \rightarrow AD + CB$$

However, a precipitate is not always formed. To determine whether a precipitation reaction has occurred and a solid has been formed, the solubility rules are used (Worked example 14.3A).

Ionic equations

Ionic equations are used to represent the ions that combine to form a precipitate in a precipitation reaction. Here, the spectator ions are excluded from the chemical equation. A spectator ion does not participate in a chemical reaction and its state remains unchanged throughout the reaction.

14.3B WORKED EXAMPLE

WRITING AN IONIC EQUATION

Write an ionic equation for the reaction between lead(II) nitrate and sodium chloride.

Solution

Think	Do
Step 1: Write the chemical equation for the reaction.	$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$
Step 2: Rewrite the equation to show the aqueous ions separately. The solid does not dissociate into ions, so keep it as a solid.	$Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s) + 2Na^+(aq) + 2NO_3^-(aq)$ $PbCl_2$ is kept as a solid.
Step 3: Identify the spectator ions. These are the ones that do not change states in the chemical reaction.	Na^+ and NO_3^- are aqueous on both sides of the reaction. Therefore, they are the spectator ions.
Step 4: Cancel out the spectator ions.	$Pb^{2+}(aq) + \cancel{2NO_3^-(aq)} + \cancel{2Na^+(aq)} + 2Cl^-(aq) \rightarrow PbCl_2(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$
Step 5: Rewrite the equation. This is the ionic equation.	$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

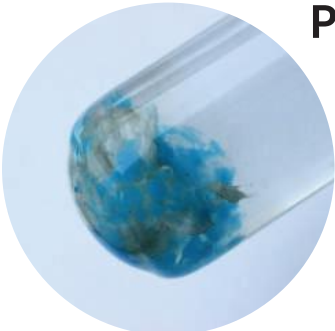


FIGURE 1 Copper silicate and manganese silicate precipitate at the bottom of a test tube.

Precipitating ions to purify water

Precipitation reactions can be used to purify water containing unwanted dissolved salts. This involves adding chemicals that react to form a precipitate with these ions. The solid impurity can then be filtered from the water.

One example is the treatment of hard water to remove magnesium (Mg^{2+}) and calcium (Ca^{2+}) ions. Rain water has a low concentration of dissolved minerals, so it known as 'soft water'. However, as the water runs through rocks, it collects minerals and becomes 'hard water'. The presence of these minerals is unwanted because:

- calcium carbonate ($CaCO_3$) forms and can build up on plumbing as limescale
- they make it difficult to form a lather when you use soap; instead, a poorly soluble film called soap scum forms that is hard to rinse away
- they dry out skin and hair; they form a film on the surface of skin and hair, which prevents moisture from penetrating
- they can disrupt the pH balance of skin and damage its protective barrier properties.

FIGURE 2 **a** Limescale build-up on a kitchen tap. **b** Soap scum forms when soap mixes with Ca^{2+} and Mg^{2+} ions in hard water. **c** Hard water can dry out and damage hair.

A good understanding of the solubility rules has helped scientists find ways to remove Ca^{2+} and Mg^{2+} ions from water. Ca^{2+} can be precipitated out as calcium carbonate (rule 8 in the solubility table, page 392) and Mg^{2+} as magnesium hydroxide (rule 7). These precipitates are then filtered from the water.

Heavy metals, such as mercury and lead, can also be removed from contaminated water by precipitation reactions. These heavy metals form precipitates with chlorides (rule 5), iodides (rule 5), bromides (rule 5) and sulfates (rule 6).

14.3 REAL-WORLD CHEMISTRY

Removing mercury pollutants

Pollutants in the atmosphere have caused problems such as acid rain, respiratory issues, poisoning of waterways and the bioaccumulation of toxic chemicals in marine life. One example is mercury(II) ($Hg^{2+}(aq)$), which (even at low concentrations) has many negative health effects.

How can you remove cations from hard water?
Go to your **Student obook pro**.

When bonded to any soluble anion (anion$^-$), $Hg^{2+}(aq)$ can be precipitated out of a water sample by adding sodium iodide. The reaction forms mercury(II) iodide ($HgI_2(s)$), a red-orange solid:

$$Hg^{2+}(aq) + \text{anion}^-(aq) + Na^+(aq) + 2I^-(aq) \rightarrow HgI_2(s) + Na^+(aq) + \text{anion}^-(aq)$$

This reaction was important for identifying dissolved mercury and removing it from water, particularly during the mercury pollution of Minamata Bay, Japan. In 1908, a factory was built close to the bay to produce chemicals required for fertilisers. Around 1932, the factory expanded to synthesise other, more dangerous chemicals, such as ethanal (CH_3CHO). This required a catalyst, mercury sulfate ($HgSO_4$). The workers did not realise that a side reaction was taking place, forming methyl mercury chloride (CH_3HgCl). This toxic compound was dumped into the bay between 1951 and 1968.

In 1956, many residents of Minamata Bay were taken to hospital with symptoms of convulsions and difficulty in speaking and walking. Many people living in the area were found to be suffering the same symptoms. The doctors named it Minamata disease. Later, the cause was found to be heavy metal poisoning due to consumption of tuna fished from the contaminated water.

Apply your understanding

1 Write an ionic equation for the formation of mercury iodide.
2 Identify the source of the mercury pollutants.
3 Identify the green chemistry principle(s) that are related to removal of pollutants from water.
4 Explain how the principle(s) you identified in Question 3 address the United Nations Sustainability Development Goals.

14.3 CHECK YOUR LEARNING

Describe and explain

1 Explain what happens to the anions and cations during a precipitation reaction.
2 Describe how precipitation reactions can be used to remove impurities from water.

Apply, analyse and compare

3 Determine whether the following reactions are precipitation reactions. Explain why.
  a $Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2 + 2NaNO_3$
  b $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$
  c $FeSO_4 + Mg(NO_3)_2 \rightarrow Fe(NO_3)_2 + MgSO_4$
4 Write the ionic equations for the precipitation reactions that you identified in Question 3.

Design and discuss

5 A student wants to perform a qualitative and quantitative test on a precipitation reaction.
  a Why is a precipitation reaction qualitative?
  b What could be done to make a precipitation reaction quantitative?
  c Design a method to determine the amount of chloride ions in a seawater sample (containing sodium chloride), using a precipitation reaction as the first step.
6 A student in your class claims that all ions participate in a chemical reaction. Evaluate this statement and justify your response.

CHAPTER 14 Review

Chapter summary

14.1

- Water is a universal solvent. Its bent molecule and permanent dipoles allow it to form strong intermolecular bonds with other polar molecules. It can also form hydrogen bonds with other polar molecules containing a H atom bonded to a strongly electronegative atom such as F, O or N.
- Solutions are formed when a solute is dissolved in a solvent. If more solute can dissolve, the solution is unsaturated. If no more can dissolve, it is saturated. If a saturated solution is heated to dissolve more solute and then cooled, it can become supersaturated.
- Concentration is a measure of the number of particles in a specific volume of solution. This is measured in moles per litre (mol L^{-1} or M), grams per litre (g L^{-1}), ppm (equivalent to mg L^{-1}), percentage mass per volume (% m/v) and percentage volume per volume (% v/v).

14.2

- Soluble substances can dissolve in a solvent. For example, solutes with similar properties to water are soluble in water – remember *like dissolves like*.
- Ionic and polar covalent substances have whole or partial charges that result in a significant interaction between the solutes and water molecules. Non-polar substances will only interact through dispersion forces, and therefore have the lowest solubility.
- Solubility tables include rules for the solubility of ionic substances in water at a specific temperature. They can be used to predict whether a substance is soluble (aqueous) or insoluble (solid).
- There is a direct relationship between temperature and solubility. In general, an increase in temperature (and an increase in kinetic energy) results in increased solubility of solids. However, the opposite is seen with gases, which decrease in solubility when temperature increases.

14.3

- Precipitation reactions occur when two aqueous ionic substances react to form at least one solid precipitate:
$$AB(aq) + CD(aq) \rightarrow AD(s) + CB(aq)$$
- An ionic equation is a simplified version of a precipitation reaction in which only the ions involved in the reaction are represented. Spectator ions (which do not change state) are not involved in the reaction and are removed from the equation:
$$A^+(aq) + D^-(aq) \rightarrow AD(s)$$
- Precipitation reactions can be used to purify unwanted ions from an aqueous solution. This involves adding a chemical substance that reacts with the ions to form an insoluble compound, which can then be filtered out and removed from the water.

Key formulas

Concentration (mol L^{-1})	$c = \frac{n}{V}$
Concentration (g L^{-1})	$c = \frac{m}{V}$
Number of moles from mass	$n = \frac{m}{M}$
Concentration (% m/v)	$\% \text{ m/v} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$
Concentration (% v/v)	$\% \text{ v/v} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• calculate the concentration of a solution and express it in units of mol L^{-1} or M, g L^{-1}, ppm, % m/v and % v/v	☐	☐	☐	Go back to Topic 14.1
• convert between different units of concentration	☐	☐	☐	Go back to Topic 14.1
• use solubility tables and solubility graphs to predict the solubility of an ionic compound	☐	☐	☐	Go back to Topic 14.2
• describe the effect of temperature on the solubility of a given solid, liquid or gas	☐	☐	☐	Go back to Topic 14.2
• explain how precipitation reactions can be used to remove impurities from water	☐	☐	☐	Go back to Topic 14.3

Revision questions

Multiple choice

1 The substance in which a solute dissolves to form a solution is called a:

A solvent.
B saturate.
C solute.
D substance.

2 A solution in which no more solute can dissolve is referred to as:

A a solution.
B supersaturated.
C saturated.
D unsaturated.

3 Which of the following is not a unit of concentration?

A M
B mol L^{-1}
C ppm
D kJ L^{-1}

4 What volume of water must be added to 40 mL of 2 M HCl to make a 0.4 M solution?

A 200 mL
B 2 L
C 160 mL
D 240 mL

5 0.056 g of LiOH is dissolved to make 9.6 L of solution. Which answer is *not* a correct concentration of this solution?

A 2.4×10^{-4} M
B 5.8×10^{-3} g L^{-1}
C 5.8 ppm
D 5.8×10^{-3} M

6 A solution that forms crystals is referred to as:

A supersaturated.
B saturated.
C unsaturated.
D none of the above.

FIGURE 1 Crystals forming in a solution

7 Which of the following ionic compounds is insoluble in water?

A $NaOH$

B KCl

C $AgCl$

D $NaNO_3$

8 Which of the following ionic compounds is soluble in water?

A Hg_2Cl_2

B $BaSO_4$

C $PbCl_2$

D K_2SO_4

9 Ionic compounds have the highest solubility in water because they:

A are polar molecules containing partial charges.

B are non-polar molecules.

C are polar molecules containing whole charges.

D contain whole charges on ions.

10 Non-polar molecules are the least soluble in water because they:

A contain dispersion forces, creating a temporary dipole and therefore weaker attractions between molecules.

B contain a permanent dipole but a large amount of kinetic energy and therefore have weaker attractions between molecules.

C contain ionic interactions, but these become weaker as the substance dissolves because cations and anions move away from one another in the water.

D contain dispersion forces, creating a permanent dipole and therefore weaker attractions between molecules.

Short answer

Describe and explain

11 Explain why some molecular substances are soluble in water, while others form layers with water.

12 Explain why a precipitation reaction forms a solid. Use an example to demonstrate.

13 Explain why the solubilities of all substances are converted to or measured in g/100 g of water.

14 KBr and K_2SO_4 both contain potassium cations but have very different solubility data. What differences between these substances is causing this change? Why does KBr have a higher solubility than K_2SO_4?

15 Explain why the solubility of gases in water increases as the temperature of the water decreases.

16 Water is referred to as a 'universal solvent'. Explain why.

Apply, analyse and compare

17 Apply your knowledge to complete the following molarity calculations.

a What volume of solvent is required to make a 0.2 M KNO_3 solution from 0.5 g of KNO_3?

b What is the molarity of a solution made from 0.2 g of KOH in 45 mL of water?

c What mass of $NaCl$ must be added to 450 mL of water to generate a 0.50 M solution?

18 Determine the:

a concentration of a solution prepared from 4.22×10^{21} molecules of $CaCl_2$ in 50 mL of water. Express your answer in g L^{-1}

b concentration of dichromate ions in a solution containing 2.6 mg $K_2Cr_2O_7$ in 105 mL of water. Express your answer in mol L^{-1}

c mass of potassium permanganate, in grams, in a 246 mL solution of 0.6 g L^{-1} $KMnO_4$.

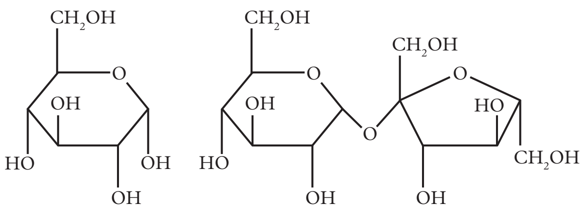
19 Determine the:

a concentration (in ppm) of an iron chloride solution made from 0.045 g of $FeCl_3$ dissolved in a 2.70 L solution

b mass of $NaOH$ required to make a 2.7 ppm solution in 80 mL of water

c volume of water that should be added to 0.01 g of $NaNO_3$ to make a 5 ppm solution.

20 The structures of glucose and sucrose are shown in Figure 2. Using the solubility graph in Figure 3 and the structures of these molecules, explain the difference in their solubility data.



Glucose **Sucrose**

FIGURE 2 The structures of glucose and sucrose

21 Write balanced precipitation reactions for the following reactions and determine whether their products are soluble or insoluble.

- **a** Sodium sulfate + barium nitrate
- **b** Lead(II) nitrate + sodium hydroxide
- **c** Aluminium chloride + potassium hydroxide

22 Use the solubility graph in Figure 3 to answer the following questions.

- **a** Determine the solubility of NH_4NO_3 at 20°C.
- **b** Determine the solubility of K_2SO_4 at 80°C.
- **c** Identify the minimum temperature required to create a saturated solution of glucose, if 25 g of glucose is dissolved in 20 g of water.
- **d** Determine the solubility of sucrose in 40 g of water at 30°C.
- **e** Determine the solubility of $CaCl_2$ in 9 g of water at 40°C.
- **f** Explain how a student could make a supersaturated solution of NH_4NO_3 at 20°C.

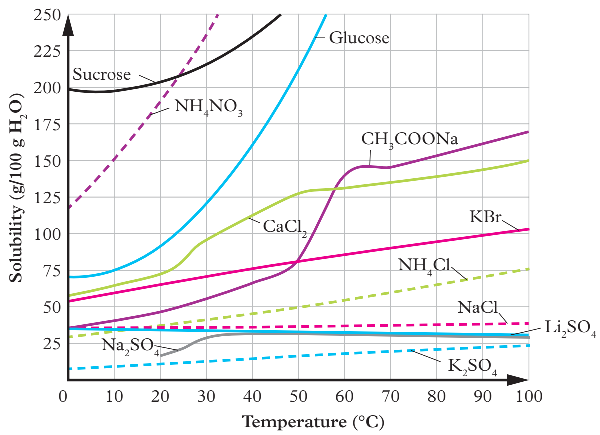


FIGURE 3 Solubility graphs of some substances in water

Design and discuss

23 A student claims that a supersaturated solution cannot be made in a laboratory without a hot plate. Discuss this claim, using the solubility rules to justify your response.

24 Design a method to determine whether carbonate, bromide, aluminium and calcium ions are present in an unknown sample. Present your method as numbered steps, detailing the equipment that could be used in a laboratory.

25 Design a method to purify a sample of wastewater. In your response, you must write balanced precipitation reactions and explain which ions are removed by each reaction.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 15

Analysis for acids and bases

KEY KNOWLEDGE

→ volume–volume stoichiometry (solutions only) and application of volumetric analysis, including the use of indicators, calculations related to the preparation of standard solutions, dilution of solutions, and use of acid–base titrations (excluding back titrations) to determine the concentration of an acid or a base in a water sample

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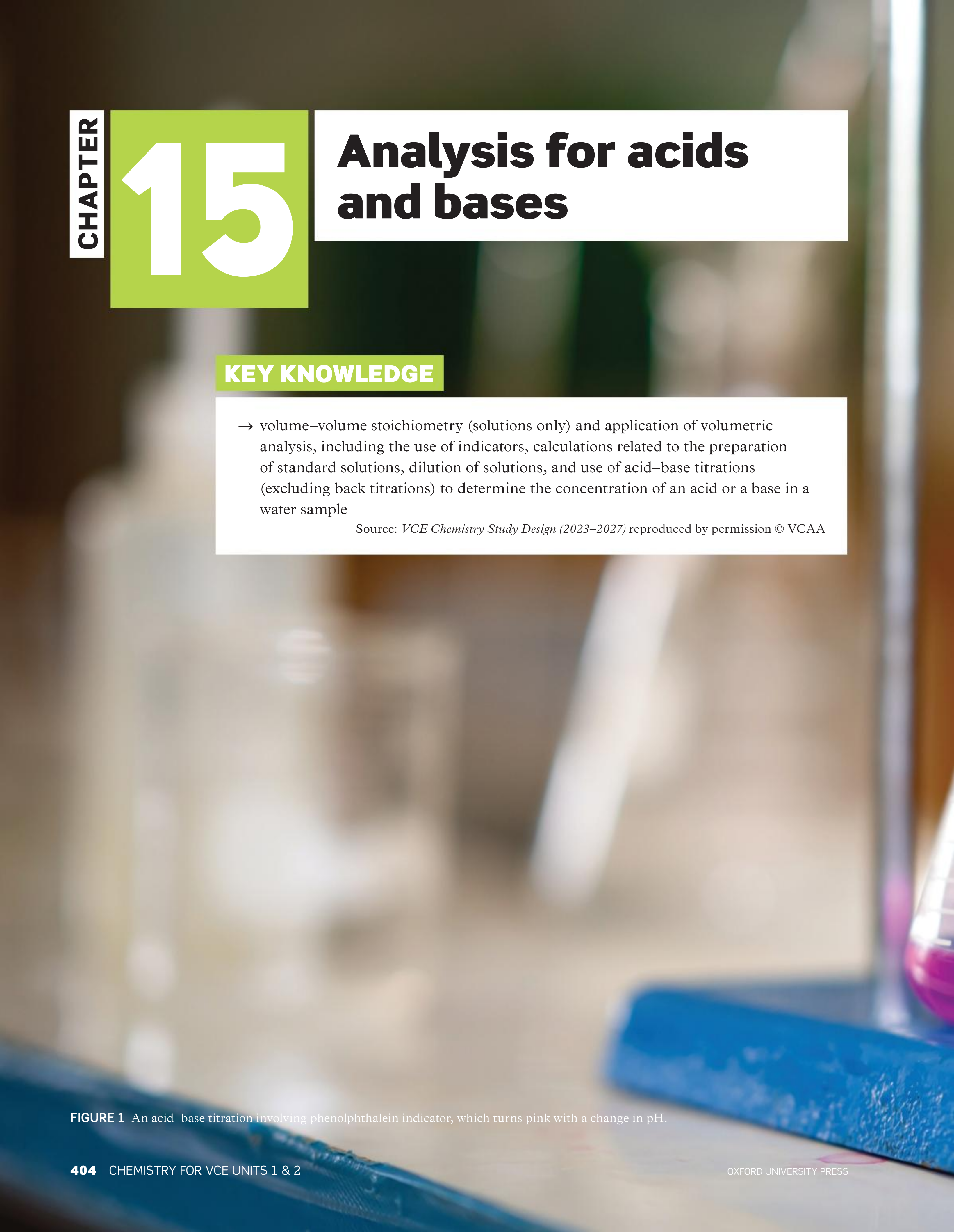
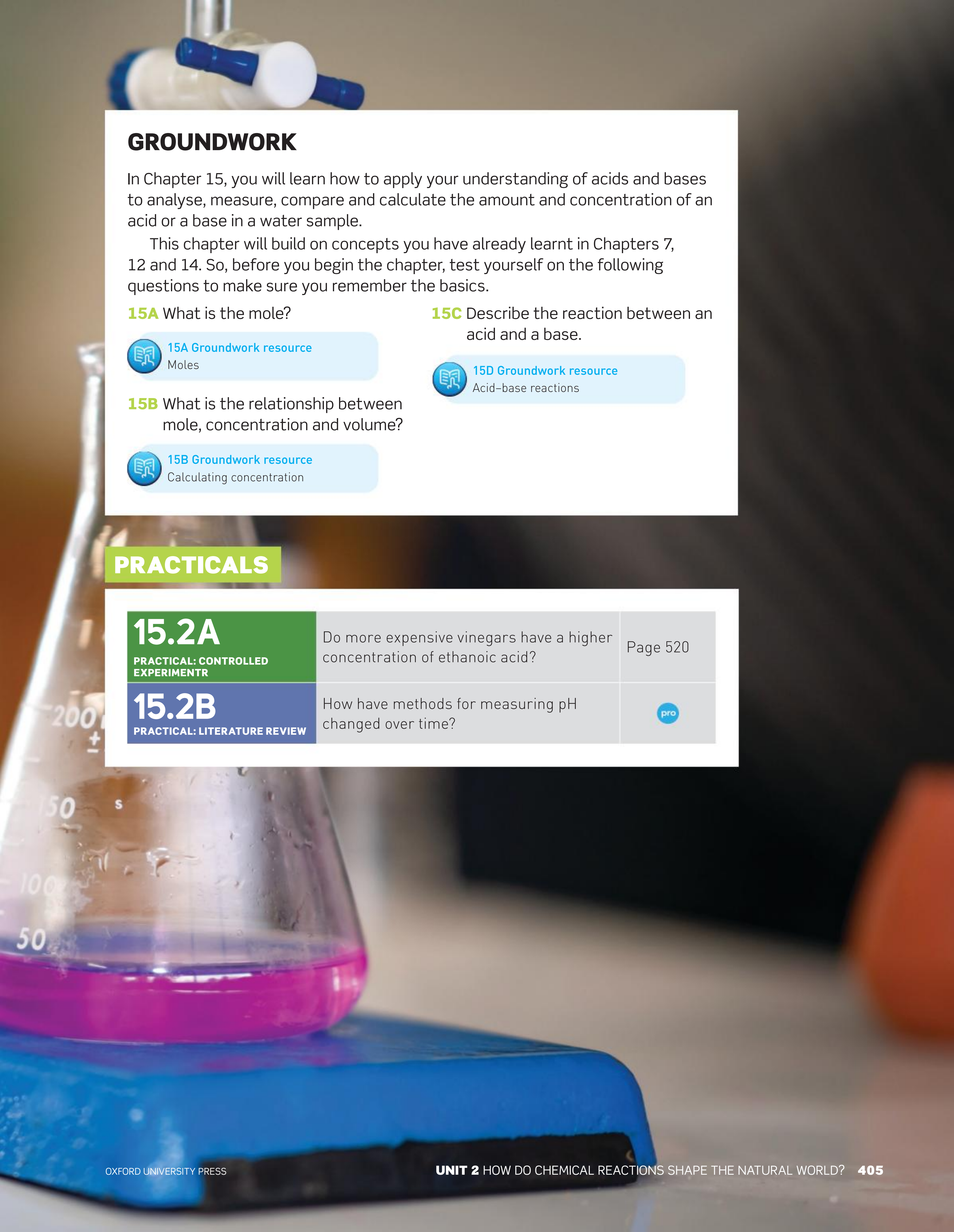


FIGURE 1 An acid–base titration involving phenolphthalein indicator, which turns pink with a change in pH.



GROUNDWORK

In Chapter 15, you will learn how to apply your understanding of acids and bases to analyse, measure, compare and calculate the amount and concentration of an acid or a base in a water sample.

This chapter will build on concepts you have already learnt in Chapters 7, 12 and 14. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

15A What is the mole?

15A Groundwork resource
Moles

15B What is the relationship between mole, concentration and volume?

15B Groundwork resource
Calculating concentration

15C Describe the reaction between an acid and a base.

15D Groundwork resource
Acid–base reactions

PRACTICALS

15.2A PRACTICAL: CONTROLLED EXPERIMENTR	Do more expensive vinegars have a higher concentration of ethanoic acid?	Page 520
15.2B PRACTICAL: LITERATURE REVIEW	How have methods for measuring pH changed over time?	pro

15.1 Volume–volume stoichiometry

KEY IDEAS

In this topic, you will learn that:

+ you can use the mole ratio to calculate an unknown volume or concentration of a solution
+ when excess reactant is added to a reaction, the limiting reagent will 'drive' the reaction
+ diluting a solution will increase the volume and decrease the concentration.

In Chapter 12, you learnt that when an acid and a base are added together, they undergo a neutralisation reaction that produces a salt and water. This can be represented by the following equation:

$$\text{Acid} + \text{base} \rightarrow \text{salt} + \text{water}$$

Acids and bases are often analysed by **volumetric analysis**. This technique allows us to determine the concentration of acidic or basic solutions.

volumetric analysis a quantitative analytical technique for determining concentration of a solution by titrating it against another solution of known concentration and volume

standard solution solute in a solution with a precisely known concentration

stoichiometry the numerical relationship between the amounts of reactants and products in a reaction

Volumetric analysis

In volumetric analysis, a solution with a known volume and concentration (called a **standard solution**) is reacted with a second solution of known volume and unknown concentration. This technique allows us to calculate the unknown concentration of the second solution by **stoichiometry** – the use of mathematical calculations to determine the relative quantities of reactants or products formed in a chemical reaction.

For example, we could use volumetric analysis to determine the concentration of a basic solution, such as milk of magnesia. Milk of magnesia is made of magnesium hydroxide ($Mg(OH)_2$) and is taken as an antacid to neutralise stomach acid (hydrochloric acid, HCl). The reaction for this is:

$$Mg(OH)_2(aq) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$

Once you have a balanced chemical equation and you know the volume and concentration of the hydrochloric acid, you can use the mole ratio to calculate the unknown concentration of a volume of magnesium hydroxide solution.

FIGURE 1 Magnesium hydroxide is a basic solution that can be used to neutralise stomach acid.

Volume–volume stoichiometry

We use **volume–volume stoichiometry**, also known as solution or concentration stoichiometry, to calculate the unknown concentration or volume of solutions.

volume–volume stoichiometry calculations using the mole ratio for a chemical reaction where one solution with known volume and concentration is used to find the unknown volume or concentration of another solution

FIGURE 2 This beaker contains a solution of known volume, but unknown concentration. Volumetric analysis can be used to work out its concentration.

To calculate the unknown concentration, we use the mole ratio.

In a balanced chemical equation, the coefficients give the ratio in which substances will react. For example:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

In this reaction, 2 mol of sodium hydroxide (NaOH) reacts with 1 mol of sulfuric acid (H_2SO_4) to produce 1 mol of the salt (sodium sulfate (Na_2SO_4)) and 2 mol of water (H_2O). This is an acid–base neutralisation reaction, in which an acid and a base react to form a salt and water.

Calculating unknown solution concentration or volume

Using the ratios from a balanced chemical equation, you can calculate a solution's unknown concentration or volume by transposing the equation:

$$n = c \times V$$

where n is amount of substance (mol), c is concentration (mol L^{-1} or M) and V is volume (L).

There are four main steps to performing calculations for volume–volume stoichiometry.

1 Write a balanced chemical equation, define the variables and identify the substance with known concentration.

2 Calculate the amount (in mol) of the substance with the known concentration and volume, using $n = c \times V$.

3 Use the mole ratio for the equation to calculate the unknown amount of substance (in mol) of the other reactant: mole ratio $= \dfrac{\text{unknown coefficent}}{\text{known coefficient}}$

4 Calculate the unknown volume or concentration using $n = c \times V$.

Volume (V) or concentration (c) of known substance $n = c \times V$ → Number of moles of known substance (n) $\xrightarrow{\text{From coefficients in chemical equation}}$ Number of moles of unknown substance (n) → Volume (V) or concentration (c) of known substance $V = n/c$ or $c = n/V$

FIGURE 3 A flow chart of volume–volume stoichiometry

Worked examples 15.1A and 15.1B demonstrate the steps for calculating the known and unknown number of moles, volumes and concentrations using stoichiometry.

15.1B Worked example
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15.1B Worked example
Video demonstration

15.1A WORKED EXAMPLE

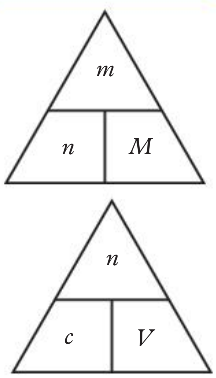
CALCULATING THE CONCENTRATION OF A SOLUTION

What is the concentration of 20.00 mL of potassium hydroxide (KOH) solution that reacts completely with 17.6 mL of 0.50 M phosphoric acid (H_3PO_4) solution?

Solution

Think	Do
Step 1: Write the balanced equation for the reaction, define the variables and identify the known substance. Remember that volume must be in litres.	Known substance $3KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$ $c = ?$ $c = 0.50$ M $V = \frac{20.00 \text{ mL}}{1000}$ $V = \frac{17.6 \text{ mL}}{1000}$ $= 0.020\ 00$ L $= 0.0176$ L
Step 2: Calculate the amount, in mol, of the substance with the known concentration and volume.	We have identified our known substance as H_3PO_4, so we calculate the number of mol of H_3PO_4 using $n = c \times V$. $c(H_3PO_4) = 0.50$ M $V(H_3PO_4) = 0.0176$ L $n(H_3PO_4) = c \times V = 0.05 \times 0.0176 = 0.00\ 88$ mol $n(H_3PO_4) = 0.000\ 88$ mol
Step 3: Use the mole ratio for the equation to calculate the unknown mol of the other reactant.	mole ratio = $\frac{\text{unknown coefficent}}{\text{known coefficient}}$ $\frac{n(KOH)}{n\ (H_3PO_4)} = \frac{3}{1}$ To calculate the unknown mol of KOH, multiply the known mol of H_3PO_4 by the mole ratio. $n(KOH) = n(H_3PO_4) \times$ mol ratio $n(KOH) = 0.000 88 \text{ mol} \times \frac{3}{1}$ $n(KOH) = 0.0264$ mol
Step 4: Calculate the unknown concentration using $n = c \times V$.	$n(KOH) = 0.02\ 64$ mol $V(KOH) = 0.02000$ L $c(KOH) = \frac{n}{V} = \frac{0.02 64}{0.020 00}$ $c(KOH) = 1.3$ M (2 sig fig)

Study tip



These mole equation triangles are useful for transposing equations quickly. You use them by covering up the value you want to calculate. This leaves two that are now in the correct position for calculating (e.g. cover V, calculate $\frac{n}{c}$).

15.1C Worked example
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15.1C Worked example
Video demonstration

Mass–volume stoichiometry

As well as volume–volume stoichiometry, we can use mass–volume stoichiometry to determine the amounts of reactants or products. The principle is the same, but the equation to find the amount of a substance in mol, can be used:

$$n = \frac{m}{M}$$

where n is amount of substance (mol), m is mass (g) and M is molar mass (g mol^{-1}).

Depending on the reactants or products being calculated, a combination of $n = \frac{m}{M}$ and $n = c \times V$ can be used in the same stoichiometry steps as with volume–volume stoichiometry.

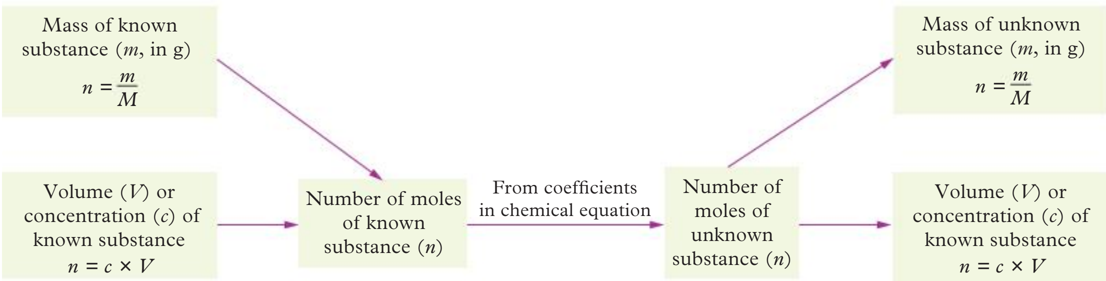

FIGURE 4 A flow chart for mass–volume stoichiometry

Mass solution stoichiometry can be used to calculate the amount of product produced in a reaction also. You can see this in Worked example 15.1C.

Excess reactants in stoichiometry

Reactants are not always mixed in their stoichiometric amounts. Sometimes there is more of one reactant, called an **excess reagent** (or excess reactant). At the end of a chemical reaction, there is excess reactant left over.

The **limiting reagent** (or limiting reactant) is the reactant that is used up first. Once the limiting reagent is used up, the chemical reaction stops and no more product can be formed. The limiting reagent therefore 'limits' the reaction.

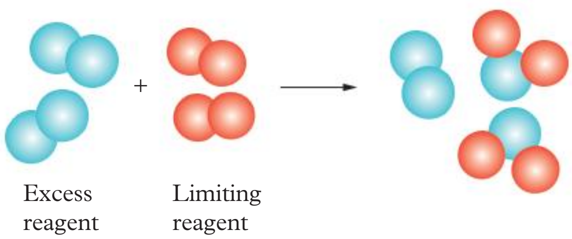

FIGURE 5 The limiting reagent 'limits' how much product can form in a reaction. At the end of the reaction, there will be excess reactant left over.

To calculate the amount of product formed or excess reagent remaining after an acid–base reaction, you need to do the following.

1 Calculate the amount (in mol) of each of the reagents
2 Determine which reagent is in excess and which is the limiting reagent
3 Use the limiting reagent amount (mol) to calculate the amount of product formed *or* the amount of reagent in excess.

Worked example 15.1D shows you how to calculate excess and limiting reagents. If you're looking for an additional challenge, try Challenge 15.1.

Dilutions

Sometimes, the concentration of a solution is too high for it to be analysed by volumetric analysis, so the solution needs to be diluted. A **dilution** involves adding extra solvent, in most cases water. The amount (in mol) of solute remains the same, but the volume and concentration after the dilution changes.

Study tip

You don't need to learn mass–volume stoichiometry in Units 1 & 2 Chemistry, *but* it is part of the Units 3 & 4 course. If you feel confident with these concepts, learning mass–volume stoichiometry would be good preparation for next year.

excess reagent (reactant)
a reactant that is available in a greater amount than is needed for a reaction and is not completely consumed

limiting reagent (reactant)
a reactant that is completely consumed in a reaction and which determines the amount of product formed or excess reactant used

15.1D Worked example
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15.1D Worked example
Video demonstration

15.1 Challenge
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dilution
a decrease in the concentration of a solution by adding solvent

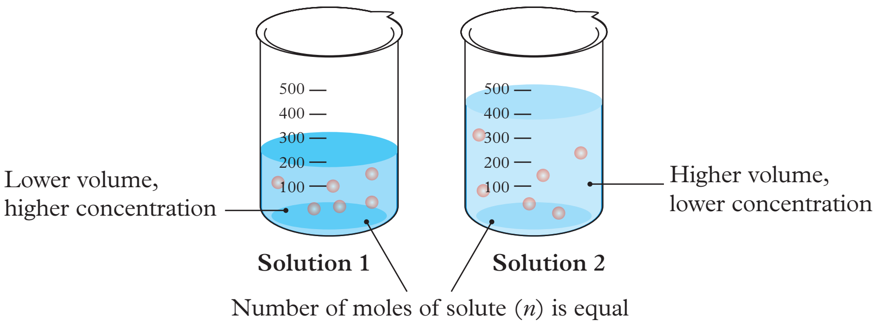


FIGURE 6 Diluting a solution decreases the concentration and increases the volume.

Calculating the concentration of a dilution

After dilution, the amount (in mol) is unchanged (i.e. the amount in the original solution n_1 and the amount in the diluted solution n_2 are the same). You can use the equation $n = c \times V$ to calculate the concentrations of the original and diluted solutions. Therefore, the equation for a dilution is:

$$c_1 \times V_1 = c_2 \times V_2$$

where the subscript of 1 is for the original solution and the subscript of 2 is for the diluted solution. For example, Figure 7 shows 6 mL of a stock solution of 6 M HCl being diluted to form a new 100 mL solution.

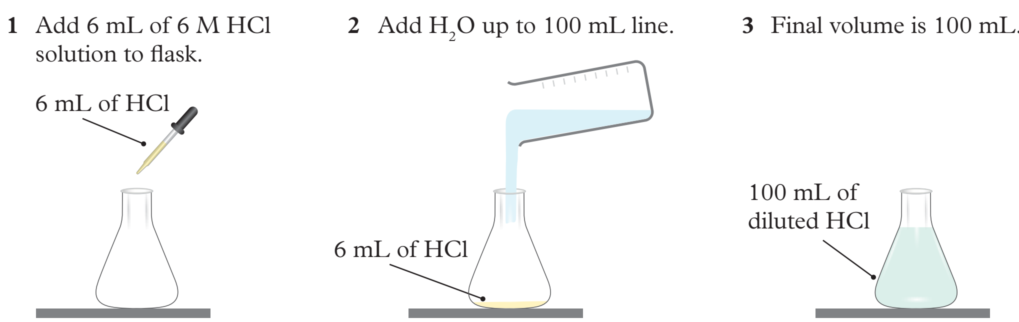


FIGURE 7 Diluting a HCl solution

The new concentration of the diluted HCl can then be calculated using $c_1 \times V_1 = c_2 \times V_2$ and the following steps.

1 Define the variables:

$c_1 = 6$ M

$V_1 = 10$ mL

$c_2 = ?$

$V_2 = 100$ mL

2 Insert the variables into the equation $c_1 \times V_1 = c_2 \times V_2$:

$6 \times 10 = c_2 \times 100$

3 Transpose and solve for the unknown:

$$c_2 = \frac{6 \times 10}{100}$$

$$= 0.6 \text{ M}$$

The new concentration of diluted HCl is 0.6 M. For these calculations, it doesn't matter if the units of volume are mL or L. As long as they are the same units for V_1 and V_2 (e.g. both in mL), they will cancel out.

The **dilution factor** for the HCl dilution is $\frac{100}{10} = 10$. This means the original solution is 10 times more concentrated than the new diluted solution.

dilution factor
the ratio of the final volume to the initial volume in a dilution

15.1E WORKED EXAMPLE

CALCULATING THE CONCENTRATION OF A DILUTION

Calculate the volume of 8.0 M NaOH stock solution that is required to make 250.0 mL of 0.20 M NaOH solution.

Solution

Think	Do
Step 1: Define the variables.	$c_1 = 8.0$ M $V_1 = ?$ $c_2 = 0.20$ M $V_2 = 250.0$ mL
Step 2: Insert the variables into the equation $c_1 \times V_1 = c_2 \times V_2$.	$8.0 \times V_1 = 0.20 \times 250.0$
Step 3: Transpose and solve for the unknown.	$V_1 = \frac{0.20 \times 250.0}{8.0}$ $= 6.25$ mL $= 6.3$ mL (2 sig fig)

15.1 CHECK YOUR LEARNING

Describe and explain

1 Identify a purpose for volume–volume stoichiometry calculations.

2 Outline the steps required for volume–volume stoichiometry calculations.

3 Describe the process of dilution.

Apply, analyse and compare

4 Determine the reactant that is in excess in a reaction between 20.00 mL of 1 M HCl and 30.00 mL of 0.8 M KOH.

5 An acid–base reaction takes place in which 20.00 mL of 0.20 M sulfuric acid (H_2SO_4) is reacted with an unknown volume of 0.50 M potassium hydroxide (KOH).

- **a** Write a balanced chemical equation for the reaction.
- **b** Calculate the volume of potassium hydroxide required to completely neutralise the sulfuric acid.

6 Calculate the concentration of 25.00 mL of HCl that is required to completely neutralise 15.00 mL of 0.400 M $Ca(OH)_2$.

7 Calculate the volume of 0.90 M H_2SO_4 that is required to completely neutralise 30.00 mL of 0.50 M NaOH.

8 Calculate the dilution factors for the following dilutions.

- **a** A 25.00 mL solution of 0.40 M HCl is diluted to 100.0 mL.
- **b** A 15.00 mL solution of 1.5 M NaOH is diluted to 250.0 mL.
- **c** A 50.00 mL solution of 20 M nitric acid is diluted to 500.0mL.

9 For the dilutions in Question 8, calculate each unknown concentration.

10 Determine the final volume of a 30.00 mL solution of 5.0 M sulfuric acid that is diluted to a concentration of 0.25 M.

11 A 25.00 mL solution of 0.28 M sodium hydroxide is reacted with 10.00 mL of 0.36 M hydrochloric acid solution.

- **a** Identify which reactant is the limiting reagent.
- **b** Calculate the mass of the product NaCl of the reaction when it is complete.

15.2 Acid–base titrations

KEY IDEAS

In this topic, you will learn that:

+ volume–volume stoichiometry is used to calculate the unknown concentration in an acid–base titration
+ specific equipment is used in titrations to ensure their accuracy
+ indicators for acid–base titrations are selected so their end point is close to the equivalence point.

titration
a quantitative analytical technique that is used to find the unknown concentration of a solution

acid–base titration
a titration that uses an acid–base neutralisation reaction to find an unknown concentration

Titrations are a technique used in volumetric analysis that can easily be conducted in the laboratory. There are different types of titrations but, in this topic, we will focus on acid–base titrations.

An **acid–base titration** uses volume–volume stoichiometry to calculate an unknown concentration of an acid or a base in a reaction. To do so, this analytical technique uses very specific equipment, indicators and solutions, which we'll examine shortly.

Chemists can use acid–base titrations for a wide range of applications, such as determining the:

- purity of medicines (e.g. the concentration of ephedrine in cough medicine)
- concentration of contaminants in wastewater
- correct number of microbes to ferment milk into other products.

With such varied applications, acid–base titrations are very important to society.

Setting up an acid–base titration

During VCE Chemistry, you will need to set up and conduct a titration.

An acid–base titration uses a volume of standard solution that has a known concentration to react with a volume of solution with an unknown concentration.

The steps for conducting an acid–base titration (Figure 1) are as follows.

aliquot
a volume of liquid delivered by the pipette

1 Use a pipette to measure an **aliquot** of unknown concentration of acid into a conical flask.
2 Add an appropriate indicator to the acid; the indicator should be chosen to change colour when the solution is neutralised.

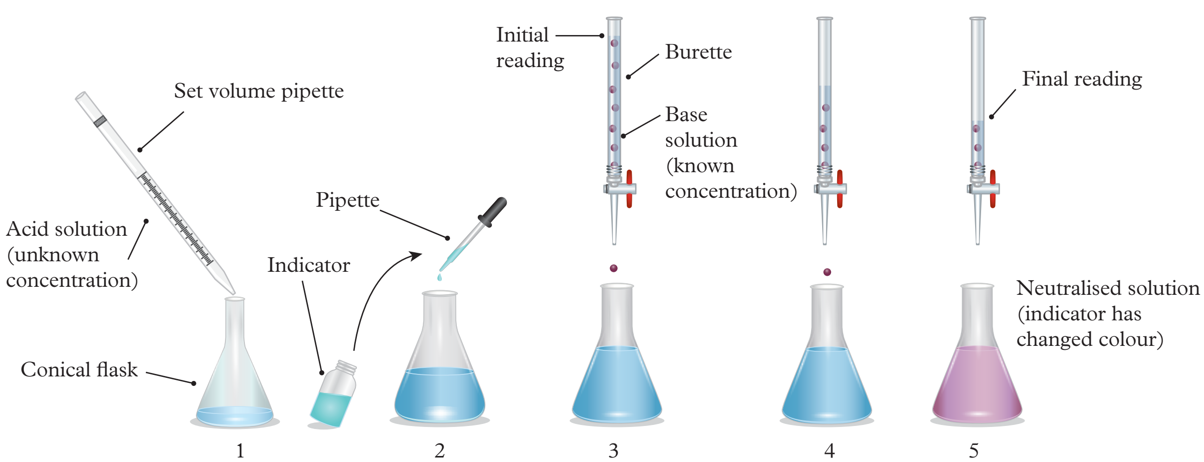


FIGURE 1 The steps for conducting an acid–base titration

3 Fill the burette with a standard solution of base of known concentration and take an initial reading.
4 Carefully release small volumes from the burette into the conical flask, swirling the conical flask to ensure the solutions mix well.
5 A colour change will occur when the end point is reached. You will learn about end points later in this topic.
Calculations can then be done by using volume–volume stoichiometry.

Standard solutions

A standard solution is a solution with an accurately known concentration. There are generally two methods for finding the accurate concentration of a standard solution:

1 using a **primary standard** and dissolving it in a known volume of water. A primary standard is a pure substance that has a specific set of criteria
2 reacting it with another solution whose concentration is known accurately.

primary standard
a very pure substance that is used to make a primary standard solution

Primary standard

To be suitable as a primary standard, a substance must:

- have a known chemical formula
- have a high state of purity
- have a relatively high molar mass so that the errors when weighing a sample are minimised
- be easy to store without reacting with the atmosphere or deteriorating
- be cheap and readily available.

Not all substances make suitable primary standards. Sodium carbonate (Na_2CO_3) is a good primary standard for titration of aqueous acids (such as hydrochloric acid, sulfuric acid and nitric acid solutions) because it meets the criteria for a primary standard.

Preparing a standard solution

The steps for preparing a standard solution (Figure 2) are: as follows.

1 Weigh the pure solid.
2 Transfer the solid into a volumetric flask, using a clean funnel.
3 Wash the remaining solid into the flask with deionised water.
4 Fill the flask halfway with deionised water. Seal the flask with the stopper.
5 Swirl the sealed flask to mix the contents thoroughly.
6 Fill the flask to the calibration line on the neck of the flask. The bottom of the meniscus should be at the top of the line when measured at eye level. Add the stopper again and shake the flask to mix.

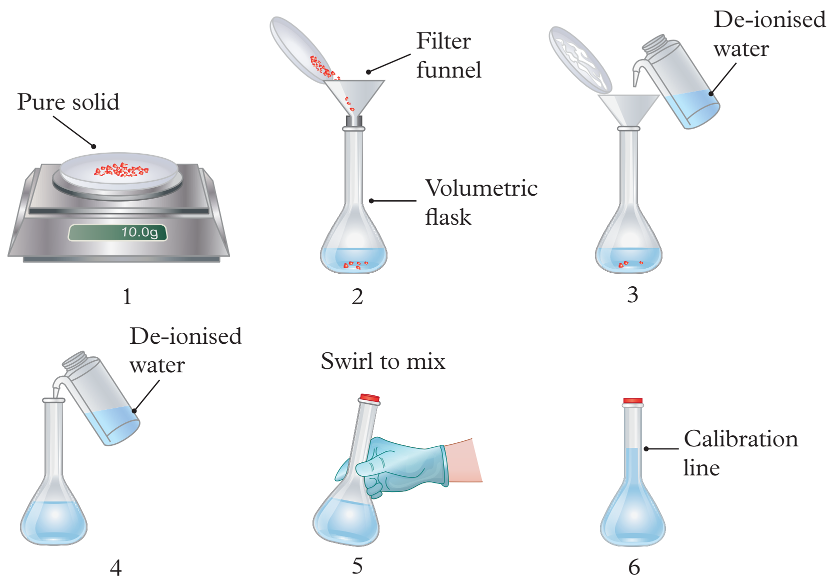


FIGURE 2 The steps for preparing a standard solution

Calculations involved in preparing a standard solution

To calculate the concentration of a standard solution, use:

1 the chemical formula to determine the molar mass (M) of the compound
2 $n = \frac{m}{M}$ to calculate the amount (in mol) of the compound
3 the amount (mol) of the compound, to determine the concentration of the solution by rearranging $n = c \times V$ to $c = \frac{n}{V}$. The volume of a standard solution can also be calculated by rearranging the equation to $V = \frac{n}{c}$.

FIGURE 3 An acid–base titration is performed in a laboratory to determine an unknown concentration of a solution.

15.2A WORKED EXAMPLE

CALCULATING THE CONCENTRATION OF A STANDARD SOLUTION

Calculate the concentration of the standard solution of sodium carbonate (Na_2CO_3) when a mass of 62.1 g of the primary standard is accurately weighed and used to make up the standard solution in a 500.0 mL volumetric flask.

Solution

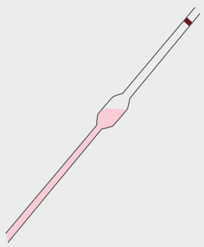
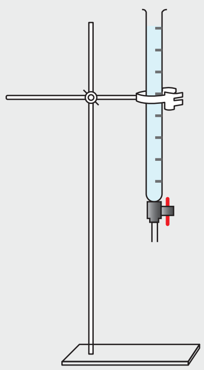
Think	Do
Step 1: Use the chemical formula to determine the molar mass (M) of the compound.	$M(Na_2CO_3) = (23.0 \times 2) + 12.0 + (16.0 \times 3)$ $M(Na_2CO_3) = 106.0$ g mol^{-1}
Step 2: Use $n = \frac{m}{M}$ to calculate the number of mol of the compound.	$m(Na_2CO_3) = 62.1$ g $M(Na_2CO_3) = 106.0$ g mol^{-1} $n(Na_2CO_3) = \frac{m}{M}$ $= \frac{62.1}{106.0}$ $= 0.586$ mol
Step 3: Use the amount in mol of the compound to determine the concentration of the solution using $c = \frac{n}{V}$	$n(Na_2CO_3) = 0.586$ mol Volume must be in litres. $V(Na_2CO_3) = \frac{500.0}{1000}$ $= 0.5000$ L $c(Na_2CO_3) = \frac{n}{V}$ $= \frac{0.586}{0.5000}$ $= 1.17$ M (3 sig fig)

Glassware used for titrations

When performing titrations, we need to use highly accurate glassware. This is because measurements must be accurate to ensure the calculations are correct.

If the glassware is not used correctly, the measurements in a titration can be inaccurate and the calculations incorrect. Table 1 outlines the glassware used in titrations, their measurement uncertainties and errors that can be involved during their use.

TABLE 1 Glassware used in titrations

Glassware	Description	Measurement uncertainties	Rinsed with	Errors involved in use
Pipette 	Used to accurately measure a fixed volume of the unknown solution. The volume from the pipette is known as an aliquot.	20 mL pipette: ±0.02 mL 30 mL pipette: ±0.03 mL	The solution to be used in it	• Rinsing with incorrect solution (e.g. water) • Not emptying pipette fully • Not reading from bottom of the meniscus • Parallax error (not reading from eye level)
Burette 	The standard solution is placed in the burette, which is held in a retort stand using a boss head and clamp. The burette delivers a variable volume of the standard solution called a titre.	50 mL burette: ±0.02 mL for each reading	The solution to be used in it	• Rinsing with incorrect solution (e.g. water) • Not reading from bottom of the meniscus • Not reading correctly to 2 decimal places • Parallax error
Volumetric flask 	Used to prepare a standard solution to a specific volume.	100 mL volumetric flask: ±0.08 mL 250 mL volumetric flask: ±0.1 mL	Deionised water	• Not filling to bottom of meniscus • Not dissolving all solid properly • Never go above the mark; you cannot remove solution from the volumetric flask once it is in
Conical flask 	The aliquot from the pipette is transferred into a conical flask, an indicator is added and the titration is completed in it.	Not used to take accurate measurements: approx. ±5 mL	Deionised water	• Cracks in glassware • Forgetting to add indicator before titrating solutions

As outlined in Table 1, to reduce errors in titrations, you need to rinse the glassware properly. A burette and pipette are rinsed with the solutions that are to be used in them because their measurement uncertainties are so small. Other glassware is rinsed with deionised water before being used in a titration.

titre
the volume delivered by the burette in a titration

concordant
volumes of three or more titres that fall within 0.1 mL of each other

To minimise the errors associated with reading a burette, an average **titre** is determined. To do this, you should obtain three **concordant** titres. A concordant titre is ±0.1 mL; that is, within 0.1 mL of the highest and the lowest titres.

Looking at the table of titration data in Table 2, we can find the concordant titres and calculate an average titre for the burette readings.

TABLE 2 Sample titration data

Titration number	1	2	3	4	5
Titre volume (mL)	15.63	14.94	14.97	14.83	14.89

You can see in Table 2 that:

- The first titration and the fourth titration are outside the 0.1 mL range of the others
- Trials 2, 3 and 5 are concordant because the difference between the highest (14.97 mL) and the lowest (14.89 mL) is 0.08 mL and titre 2 falls between those two volumes, so they can be averaged:

$$\text{Average titre} = \frac{14.94 + 14.97 + 14.89}{3} = 14.93 \text{ mL}$$

Indicators

end point
the point during an acid–base reaction where the colour of an indicator changes

An indicator is a substance (usually an organic dye) that is added to a reaction to signal the **end point** of a titration. It does this by changing colour. They are used in acid–base titrations to visually show the end of the titration. Because each indicator changes colour within a certain pH range, selecting the correct indicator for a titration is important. Table 3 has a list of indicators, their pH range and colour changes.

TABLE 3 Some acid–base indicators

Indicator	pH range	Colour at low pH	Colour at high pH
Thymol blue (first change)	1.2–2.8	Red	Yellow
Methyl orange	3.1–4.4	Red	Yellow
Bromophenol blue	3.0–4.6	Yellow	Blue
Methyl red	4.4–6.2	Red	Yellow
Bromothymol blue	6.0–7.6	Yellow	Blue
Phenol red	4.8–8.4	Yellow	Red
Thymol blue (second change)	8.0–9.6	Yellow	Blue
Phenolpthalein	8.3–10.0	Colourless	Pink

Selecting an indicator for titration

equivalence point
the point in a titration where the acid and base are mixed in exact stoichiometric ratios given by the balanced chemical equation

When selecting an indicator for an acid–base reaction, it is important to select one that changes colour (at the end point) as close to the **equivalence point** as possible. The equivalence point is where the acid and base are present in the exact stoichiometric ratio given by the balanced chemical equation.

The curve in Figure 4 shows the titration of a strong acid into a strong base. It also shows the pH colour change ranges of two indicators: phenolphthalein and methyl orange. Neither of these indicators changes colour exactly at the equivalence point. But both can be used because the graph is so steep that there is no difference in the volume of acid added for whichever indicator you select. If you used phenolphthalein, you would start with pink and titrate until it just turns clear. If you used methyl orange, you would start with yellow and stop as it turned red.

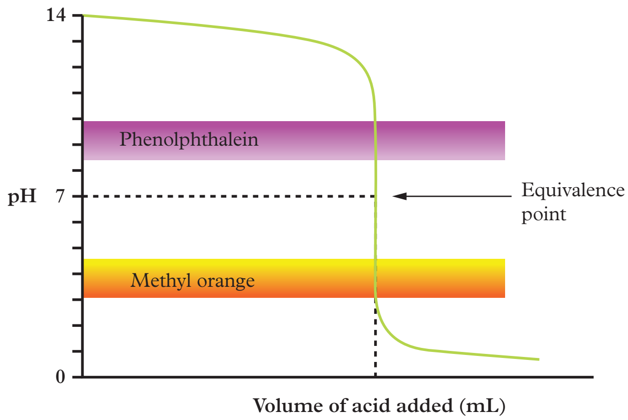


FIGURE 4 The titration of a strong acid and a strong base

In Figure 5a, you can see that phenolphthalein would be useless as an indicator for the titration of a strong acid into a weak base as an indicator as it is outside the pH range (although methyl orange changes colour very close to the equivalence point). In Figure 5b, the phenolphthalein has a colour change exactly at the equivalence point, so it is the best indicator to select for this titration.

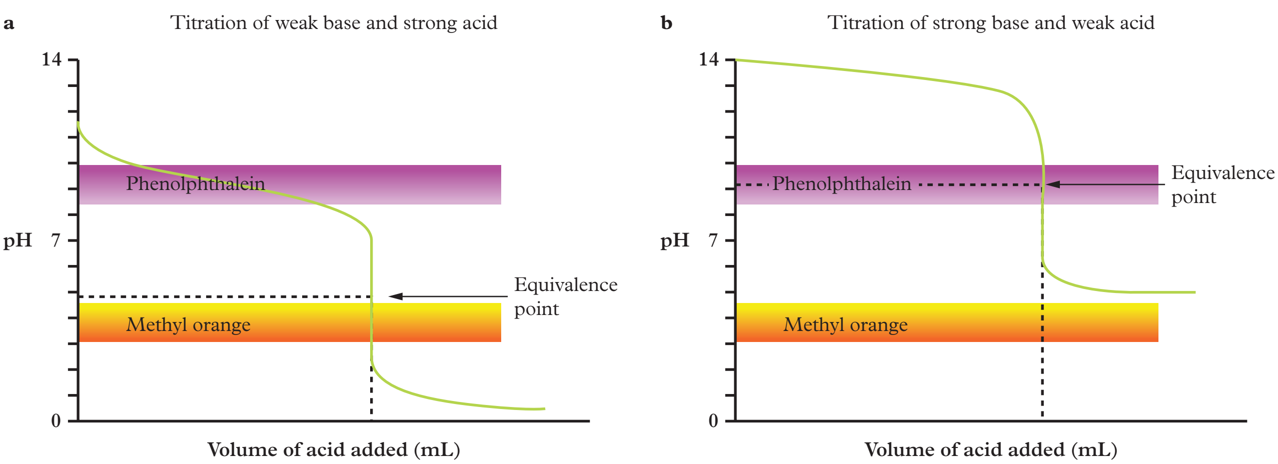


FIGURE 5 The titration curves for **a** a strong acid into a weak base and **b** a weak acid into a strong base

Study tip

When selecting an indicator, you need it to change colour at the steep part of the titration curve.

titrant
the solution being delivered by burette in a titration

Titration curves

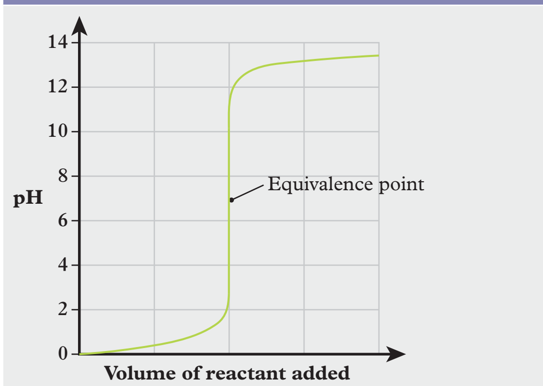
Data for titrations can be graphed using titration curves. They have the following features:

- End point – the volume at which the titration would stop if an indicator was used to find the concentration of the unknown.
- Equivalence point – the midpoint of the steep section of the curve.
- Buffer region – where there is little change in pH when the **titrant** is added. There is another section when this occurs, but only need to worry the first section as it occurs before the equivalence point.

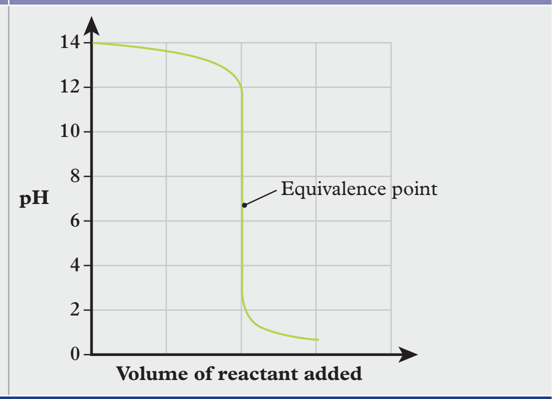
TABLE 4 Some common types of titration curves

Strong acid with strong base

Strong acid with strong base titrant

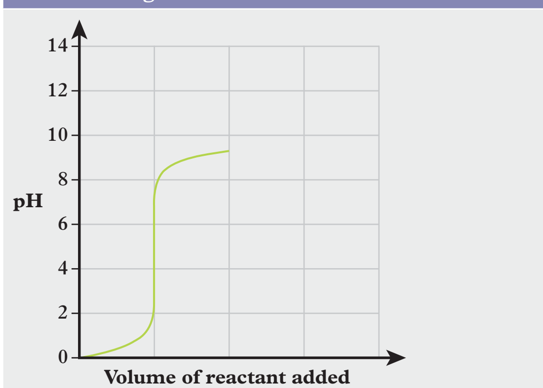


Strong base with strong acid titrant

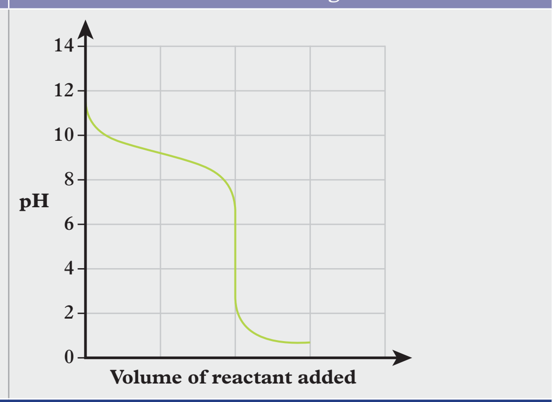


Strong acid with weak base

Strong acid with weak base titrant

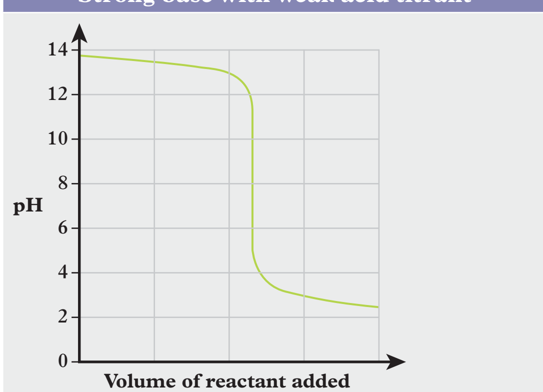


Weak base with strong acid titrant

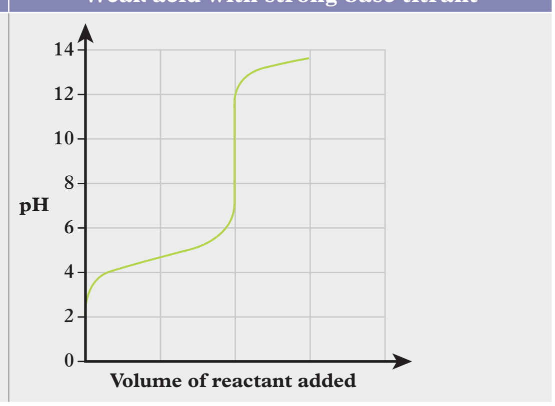


Weak acid with strong base

Strong base with weak acid titrant



Weak acid with strong base titrant



Calculations involved in acid–base titrations

The purpose of completing an acid–base titration is to accurately calculate an unknown concentration of an acid or a base in a solution.

After performing the titration, the steps in calculating the concentration are similar to those you learnt in Topic 15.1 for volume–volume stoichiometry. The steps for the titration calculations are outlined in the flow chart in Figure 6.

Step 1. Calculate the average titre (volume). → Step 2. Write the balanced equation for the reaction, define the variables and identify the known substance. → Step 3. Calculate the amount (in mol) of the substance with the known concentration and volume. → Step 4. Use the mole ratio for the equation to calculate the unknown amount (in mol) of the other reactant. → Step 5. Calculate the unknown concentration using $c = \frac{n}{V}$.

FIGURE 6 A flow chart of steps for volumetric analysis calculations

Titrations that involve dilutions

Sometimes it is necessary to perform a dilution of an unknown solution before proceeding with a titration so that you can use:

- lower and safer concentrations of solutions
- smaller titre volumes that are within the range of the burette.

Calculating the concentration of a solution by the titration of a diluted solution requires the same steps 1–5 that we have already used for titrations. After you have determined the unknown concentration of the diluted solution, there are two new steps:

6 Calculate the dilution factor, using the formula:

$$\text{Dilution factor} = \frac{\text{diluted final volume}}{\text{concentrated initial volume}}$$

7 Multiple the dilution factor by the concentration you calculated in step 5.

You can see the steps in the flow chart in Figure 7.

Step 1. Calculate the average titre (volume). → Step 2. Write the balanced equation for the reaction, define the variables and identify the known substance. → Step 3. Calculate the amount (in mol), of the substance with the known concentration and volume. → Step 4. Use the mole ratio for the equation to calculate the unknown amount (in mol) of the other reactant. → Step 5. Calculate the unknown concentration in the dilute solution, using $c = \frac{n}{V}$. → Step 6. Calculate the dilution factor = $\frac{\text{diluted final volume}}{\text{concentrated initial volume}}$ → Step 7. Calculate the concentration of the undiluted solution = (concentration of diluted solution) × (dilution factor).

FIGURE 7 A flow chart of the steps involved in calculating concentrations with a titration involving a dilution

15.2B Worked example
Find me in your obook pro

15.2B Worked example
Video demonstration

15.2 Real-world chemistry
Find me in your obook pro

Study tip

A dilution factor is the larger 'new' diluted volume divided by the smaller original volume that was taken from the sample.

15.2C Worked example
Find me in your obook pro

15.2C Worked example
Video demonstration

See Worked examples 15.2B and 15.2C to see how these steps are applied. Have a look at Real-world chemistry 15.2 to see how titrations are used in pregnancy tests.

15.2 CHALLENGE

Titrations that involve dilutions

A floor cleaner contains sodium hydroxide. A 10.00 mL sample of the floor cleaner was diluted into a 250.00 mL volumetric flask.

A titration was performed using a standard solution of 0.35 M sulfuric acid to determine the concentration of sodium hydroxide in the cleaner, and obtained the following data:

NaOH aliquot volumes	25.00 mL
Concentration of standard H_2SO_4 solution	0.35 M
H_2SO_4 titre	24.31 mL, 22.05 mL, 23.05 mL, 21.95 mL, 22.00 mL

1 Write a balanced equation for the titration.
2 Calculate the concentration of NaOH in the diluted floor cleaner.
3 Calculate the original concentration of NaOH in the floor cleaner.

15.2 SKILL DRILL

Calculating and evaluating titration data

Key science skills: Analysing and evaluating data and investigation methods

A packet of antacid tablets claims to contain 48% aluminium hydroxide ($Al(OH)_3$), which is a base used to neutralise stomach acid.

A group of students wanted to determine the percentage of aluminium hydroxide in an antacid tablet. They dissolved the 5.6 g tablet in a 100.00 mL volumetric flask of deionised water and used 20.00 mL aliquots of the solution for the titration.

They used 0.40 M hydrochloric acid solution to represent stomach acid and obtained an average titre of 15.80 mL, with three concordant results.

Practise your skills

1 Write a balanced equation for the reaction that occurred during the titration.
2 Calculate the amount (in mol) of HCl used in the titration.
3 Calculate the amount (in mol) of aluminium hydroxide in the antacid tablet.
4 Calculate the mass (in g) of the aluminium hydroxide in the antacid tablet.
5 Calculate the percentage of aluminium in the antacid tablet.
6 Compare the experimental result with the data found on the antacid packet.
7 Identify a systematic and a random error that could have occurred during their experiment.
8 Suggest a modification to the experiment that could limit each error if it were repeated.

Need help analysing and evaluating data? See Topic 1.8 (page 24).

15.2 CHECK YOUR LEARNING

Describe and explain

1 Explain why you need at least three concordant titres for a titration.

2 Define:

a aliquot

b titre.

3 Explain the results you would obtain if you incorrectly rinsed a pipette with water.

4 Explain how judging the line of the meniscus in a burette incorrectly is a random error and not another type of error.

Apply, analyse and compare

5 A 15.00 mL aliquot of barium hydroxide ($Ba(OH)_2$) with an unknown concentration is titrated with 0.80 M hydrochloric acid (HCl). The average titre volume is 22.64 mL.

a Write a balanced equation for the reaction.

b Calculate the amount (in mol) of HCl in the titre.

c Calculate the amount (in mol) of $Ba(OH)_2$ in the sample.

d Calculate the molar concentration of $Ba(OH)_2$ in the sample.

6 A bathroom cleaner contains ammonium hydroxide (NH_4OH). A 25.00 mL sample of the cleaner was diluted in a 200.00 mL volumetric flask.

A titration was performed using a standard solution of 0.500 M phosphoric acid (H_3PO_4) to determine the concentration of NH_4OH in the bathroom cleaner.

The following titration data was obtained.

NH_4OH aliquot: 25.00 mL

Concentration of standard H_3PO_4 solution: 0.500 M

Titres: 24.41 mL, 24.15 mL, 24.25 mL, 25.95 mL, 24.21 mL

a Calculate the average titre (in mL).

b Write a balanced equation for the titration.

c Calculate the molar concentration of NH_4OH in the diluted bathroom cleaner.

d Calculate the original molar concentration of NH_4OH in the bathroom cleaner.

7 A 20.00 mL volume of vinegar containing acetic acid (CH_3COOH) was pipetted into a flask and several drops of phenolphthalein indicator were added. Using a burette, 1.00 M NaOH was slowly added until the indicator turned permanently pink. The volume of sodium hydroxide required to reach this point was 19.56 mL.

a Write a balanced equation for the reaction.

b Calculate the concentration of acid in the vinegar.

c Explain what error was made in this titration.

Design and discuss

8 A group of chemistry students participated in the annual chemistry titration competition. The aim of the competition is to determine the concentration of a sample of sodium hydroxide (NaOH) and get an answer as close to the known concentration as possible.

The sample of NaOH was analysed by titrating with an aqueous solution of hydrochloric acid (HCl). 20.00 mL aliquots of the NaOH were analysed against a 0.400 M HCl solution. An average titre of 21.65 mL was obtained.

a Write a balanced equation for the reaction.

b Calculate the concentration of NaOH.

c The known concentration of the NaOH was 0.43 M.

Discuss if the students were accurate and precise in their titration.

d Suggest one systematic and one random error that could have occurred during their titration.

9 Design a method that is reproducible and would ensure minimal errors for the titration the students performed in Question 8.

CHAPTER

15 Review

Chapter summary

15.1

- Volumetric analysis is a quantitative analysis.
- The coefficients of a balanced chemical equation give the mole ratio.
- Using the mole ratio, you can determine the concentration or volume of a solution from a solution with a known concentration and volume.
- A dilution will increase the volume of the new solution and decrease the concentration; the amount (in mol) will stay the same.

15.2

- The concentration of an acid or a base can be determined by volumetric analysis.
- A standard solution is one where the concentration is accurately known.
- A pipette delivers an aliquot; it must be rinsed with the solution that it will be used with.
- A burette delivers a titre; it must be rinsed with the solution that it will be used with.
- Concordant titres are ±0.1 mL; that is, within 0.1 mL from the highest to the lowest titres.
- The end point for a titration is the colour change of the indicator.
- The equivalence point of a titration is where the acid and base are mixed in the exact stoichiometric ratio given by the balanced chemical equation.
- An indicator should be selected that has an end point as close as possible to the equivalence point.
- The dilution factor is calculated in volumetric analysis by comparing the final and initial volumes.

Key formulas

Amount (in mol) of a substance	$n = \frac{m}{M}$ or $n = c \times V$
Mole ratio	Mole ratio $= \frac{\text{unknown coefficient}}{\text{known coefficient}}$
Dilution	$c_1 \times V_1 = c_2 \times V_2$
Dilution factor	Dilution factor $= \frac{\text{diluted final volume}}{\text{concentrated initial volume}}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can …	Confidently	Mostly	Not really	Revision link
• use the mole ratio and $n = cV$ to calculate an unknown volume or concentration of a solution	☐	☐	☐	Go back to Topic 15.1
• calculate the concentration of dilutions using $c_1 \times V_1 = c_2 \times V_2$	☐	☐	☐	Go back to Topic 15.1
• prepare a standard solution, including calculating its concentration	☐	☐	☐	Go back to Topic 15.2
• identify the equipment used in acid–base titrations and describe how these should be used to avoid errors in titration data	☐	☐	☐	Go back to Topic 15.2
• explain the purpose of indicators in acid–base titrations	☐	☐	☐	Go back to Topic 15.2
• use acid–base titrations to determine the concentration of an acid or base in a water sample	☐	☐	☐	Go back to Topic 15.2
• interpret titration curves, including identifying the end point and equivalence point in a titration.	☐	☐	☐	Go back to Topic 15.2

Revision questions

Multiple choice

1 Sulfuric acid reacts with potassium hydroxide according to the equation:

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

Which of the following is incorrect?

A $n(H_2SO_4) = \frac{1}{2} \times n(H_2O)$

B $n(KOH) = \frac{1}{2} \times n(H_2SO_4)$

C $n(H_2SO_4) = \frac{1}{2} \times n(KOH)$

D $n(H_2SO_4) = n(K_2SO_4)$

2 A student titrates 20.00 mL of 0.1106 M potassium hydroxide with sulfuric acid and finds the titre required to be 22.74 mL.

$$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

The concentration of the sulfuric acid is:

A 0.048 64 M

B 0.062 90 M

C 0.097 27 M

D 0.1946 M

3 Analyse the titration curve in Figure 1.

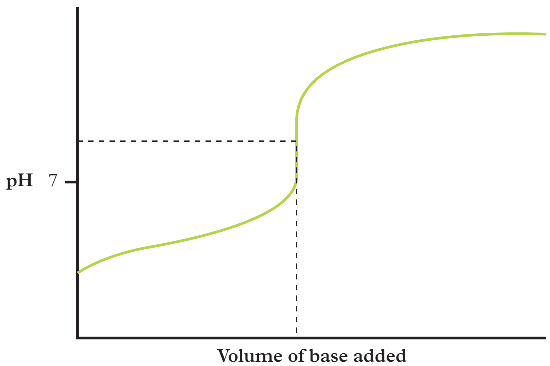


FIGURE 1 A titration curve

Select the reaction it represents.

A $CH_3COOH + NaOH$

B $NaOH + H_2SO_4$

C $NH_4OH + HCl$

D $H_3PO_4 + H_2SO_4$

4 Potassium hydrogen phthalate ($C_8H_5KO_4$) is a primary standard that is used in volumetric analysis of bases. Select the property that does not make it useful as a primary standard.

A Molar mass is 204.23 g mol^{-1}

B Low reactivity

C Has high purity

D Is a solid at room temperature

5 Steps i–v are undertaken during an acid–base titration.

i Fill burette with HCl

ii Record titre

iii Perform titration

iv Rinse burette with solution to be used in it (HCl)

v Rinse burette with water

What is the correct order for the steps?

A i, ii, iii, v, iv

B iv, v, i, iii, ii

C v, iv, i, iii, ii

D v, iv, i, ii, iii

6 A student completes a titration using 0.15 M hydrochloric acid (HCl) against 25.00 mL aliquots of sodium hydroxide (NaOH). The average titre is 18.29 mL. What is the concentration of the NaOH?

A 0.06 M

B 0.11 M

C 0.21 M

D 0.22 M

7 8.42 g of a substance is used to make up a 500 mL standard solution that has a concentration of 0.30 M. Select the substance that is the primary standard.

A Sodium hydroxide (NaOH)

B Sodium carbonate (Na_2CO_3)

C Sulfuric acid (H_2SO_4)

D Potassium hydroxide (KOH)

8 Select the correct average titre for the following titres: 23.63, 23.57, 23.73, 23.68, 23.79 mL.

A 23.68 mL

B 23.63 mL

C 23.73 mL

D 23.62 mL

9 150 mL of 1.2 M potassium hydroxide (KOH) solution undergoes a dilution. What is the final volume of a solution with the concentration 0.380 M?

A 47.4 mL

B 473.7 mL

C 68.4 mL

D 328.95 mL

10 A student completes a titration using 0.62 M potassium hydroxide (KOH) against 15.00 mL aliquots of phosphoric acid (H_3PO_4). The average titre is 22.87 mL. What is the concentration of the H_3PO_4?

A 0.32 M

B 0.95 M

C 1.2 M

D 2.8 M

Describe and explain

11 Explain the purpose of having concordant titres.

12 Explain the random and systematic errors associated with the following glassware.
  a Burette
  b Pipette
  c Volumetric flask
13 Explain how an acid–base reaction can be analysed by volumetric analysis.
14 Explain the reason for using a standard solution in a titration.
15 Describe the steps in selecting a primary standard.
16 A titration was completed, and the following five titres were produced:
  12.57 mL, 12.36 mL, 12.49 mL, 12.34 mL, 12.28 mL
  a Explain which titres are concordant.
  b Calculate the average titre.
17 A student is going to complete a titration using a standard solution of 0.5 M sodium hydroxide (NaOH) against a hydrochloric acid (HCl) solution. Identify what the student will need to rinse the following with.
  a Conical flask
  b Burette
  c Pipette
18 An acid–base titration has an equivalence point of pH 4. Suggest a suitable indicator for the titration and justify your choice.
19 Explain why sodium hydroxide (NaOH) is not suitable to use as a primary standard.

Apply, analyse and compare

20 Determine the final volume of a 65.00 mL solution of 8.0 M hydrochloric acid (HCl) that is diluted to a concentration of 0.12 M.
21 Calculate the volume of 0.77 M sulfuric acid (H_2SO_4) it would take to neutralise a 30.00 mL solution of 0.83 M sodium hydroxide (NaOH).
22 2.483 g of anhydrous sodium carbonate (Na_2CO_3) was dissolved in 200.0 mL of water. 25.00 mL aliquots of this solution were titrated with nitric acid (HNO_3). An average titre of 21.95 mL was needed. Find the concentration of the HNO_3.
23 Analyse the titration curve in Figure 2.
  a Explain whether it is for strong or weak acids or bases.
  b Explain what the pH of the equivalence point is for this titration.
  c Suggest a suitable indicator for this titration and justify your choice.

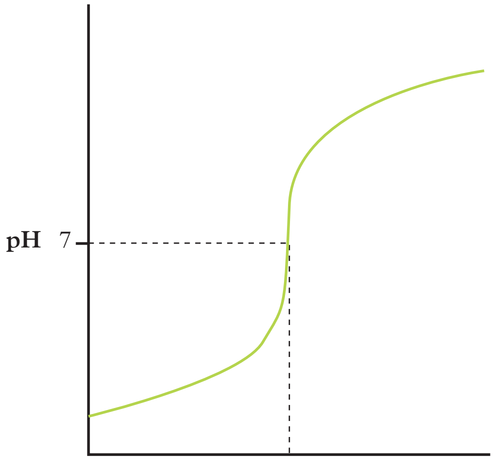


FIGURE 2 A titration curve

24 Analyse the titration curve in Figure 3.
  a Explain whether it is for strong or weak acids or bases.
  b Explain what the pH of equivalence point is for this titration.
  c Suggest a suitable indicator to be used for this titration and justify your choice.

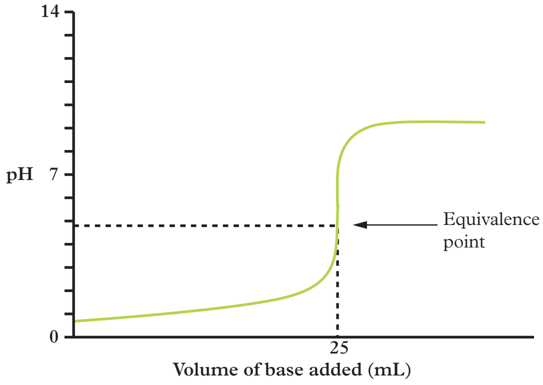


FIGURE 3 A titration curve

25 A 50.00 mL solution of 6 M sulfuric acid (H_2SO_4) is diluted to 1 L. Determine the new concentration of H_2SO_4 in the 1 L volumetric flask.

26 An acid–base reaction takes place in which 50.00 mL of 0.380 M potassium hydroxide (KOH) is reacted with an unknown concentration of 45.00 mL sulfuric acid (H_2SO_4).

a Write a balanced equation for the reaction.

b Calculate the concentration of H_2SO_4 required to completely neutralise the KOH.

27 A student rinses out a burette with deionised water, then fills it with 0.5 M hydrochloric acid (HCl). They then perform four titrations into 10.00 mL aliquots of sodium hydroxide (NaOH) with an average titre of 8.64 mL.

a Write a balanced chemical equation for the reaction.

b Calculate the concentration of the NaOH.

c Explain the errors the student made before starting this titration.

d Discuss what the actual concentration of the NaOH solution would be closer to without the error.

28 Calculate the mass of sodium carbonate (Na_2CO_3) required to make up a 0.580 M standard solution in a 1.00 L volumetric flask.

29 Two students decide to complete a volumetric analysis to check that the concentration of 1.00 M sulfuric acid (H_2SO_4) is correct. They have 100.00 mL of 2.00 M sodium hydroxide (NaOH) solution; they will be using 20.00 mL aliquots of H_2SO_4 and titrating the NaOH against them. One student says that they have enough NaOH to get three concordant titres and the other says they will not have enough and they should dilute the H_2SO_4 solution before completing the titration.

Analyse the titration the students are going to undertake, identify which student you agree with and justify your reasoning.

30 A 50.00 mL solution of 0.37 M potassium hydroxide (KOH) is reacted with 85.00 mL of 0.5 M nitric acid (HNO_3).

a Identify the limiting reagent.

b Calculate the amount (in mol) of reactant that is in excess after the reaction.

Design and discuss

31 Two students performed a titration to determine the concentration of a solution of sodium hydroxide (NaOH). They need your help to consider the data they should use in their calculations and the types of errors that may be present during their practical work.

They used phenolphthalein indicator, and 20.00 mL aliquots of 0.1 M hydrochloric acid (HCl) were titrated against the NaOH. The titres from repeated titrations were recorded.

Their results are shown in the following table.

Titration	Titre of NaOH during experiment		
	Initial burette reading (mL)	Final burette reading (mL)	Titre (mL)
1	0.50	17.60	17.10
2	5.70	22.70	17.00
3	22.60	39.85	17.25
4	2.75	19.80	17.05
Average titre			17.10

a Explain if the students correctly calculated the average titre.

b Calculate the concentration of the NaOH.

c Discuss two errors (not mistakes) that could be present during the third titration.

d A third student looked at the results of all of their titres and suggested the results were accurate. Discuss if they were correct with this statement.

32 Design a method for a volumetric analysis titration with the following parameters.

- Enough standard solution of 0.4 M sodium hydroxide (NaOH) is required to titrate against an unknown concentration of 20.00 mL aliquots of sulfuric acid (H_2SO_4),
- Three concordant results are required, but you should have enough solution so that you can perform five titrations.

In your method, include all masses, volumes and steps required to make the standard solution, all volumes of glassware required and all steps to set up and complete the titration correctly.

33 Discuss some ways in which volumetric analysis could be used for quantitative analysis of substances.

34 To confirm the concentration of a solution of phosphoric acid (H_3PO_4) as 0.50 M, a student performs a titration using 0.90 M potassium hydroxide (KOH) standard solution. They use 20.00 mL aliquots of H_3PO_4 and the titres of KOH are 32.90, 33.35, 33.41, 33.68 and 33.33 mL.

a Calculate the average titre.

b Write the balanced chemical reaction.

c Calculate the experimental concentration of the H_3PO_4.

d Compare the experimental value with the theoretical value. Discuss if the student was accurate and precise in their volumetric analysis.

e Discuss any errors they may have made during the experiment.

FIGURE 4 Standard solutions are required for titrations.

35 A 25.00 mL solution of 0.85 M sulfuric acid (H_2SO_4) is reacted with 120.00 mL of 0.16 M barium hydroxide ($Ba(OH)_2$).

a Identify the limiting reagent.

b Calculate the amount of reagent that is in excess after the reaction.

c The salt precipitate formed is filtered out of the solution after the reaction is complete. Calculate the mass of the precipitate produced in this reaction.

d Calculate the exact volume of the excess reagent that should have been used so that no excess was added.

e Discuss how this is an example of volumetric analysis, but not an example of a titration.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER 16

Measuring gases

KEY KNOWLEDGE

- → CO_2, CH_4 and H_2O as three of the major gases that contribute to the natural and enhanced greenhouse effects due to their ability to absorb infrared radiation
- → the definitions of gas pressure and standard laboratory conditions (SLC) at 25°C and 100 kPa
- → calculations using the ideal gas equation ($PV = nRT$), limited to the units kPa, Pa, atm, mL, L, °C and K (including unit conversions)
- → the use of stoichiometry to solve calculations related to chemical reactions involving gases (including moles, mass and volume of gases)
- → calculations of the molar volume or molar mass of a gas produced by a chemical reaction

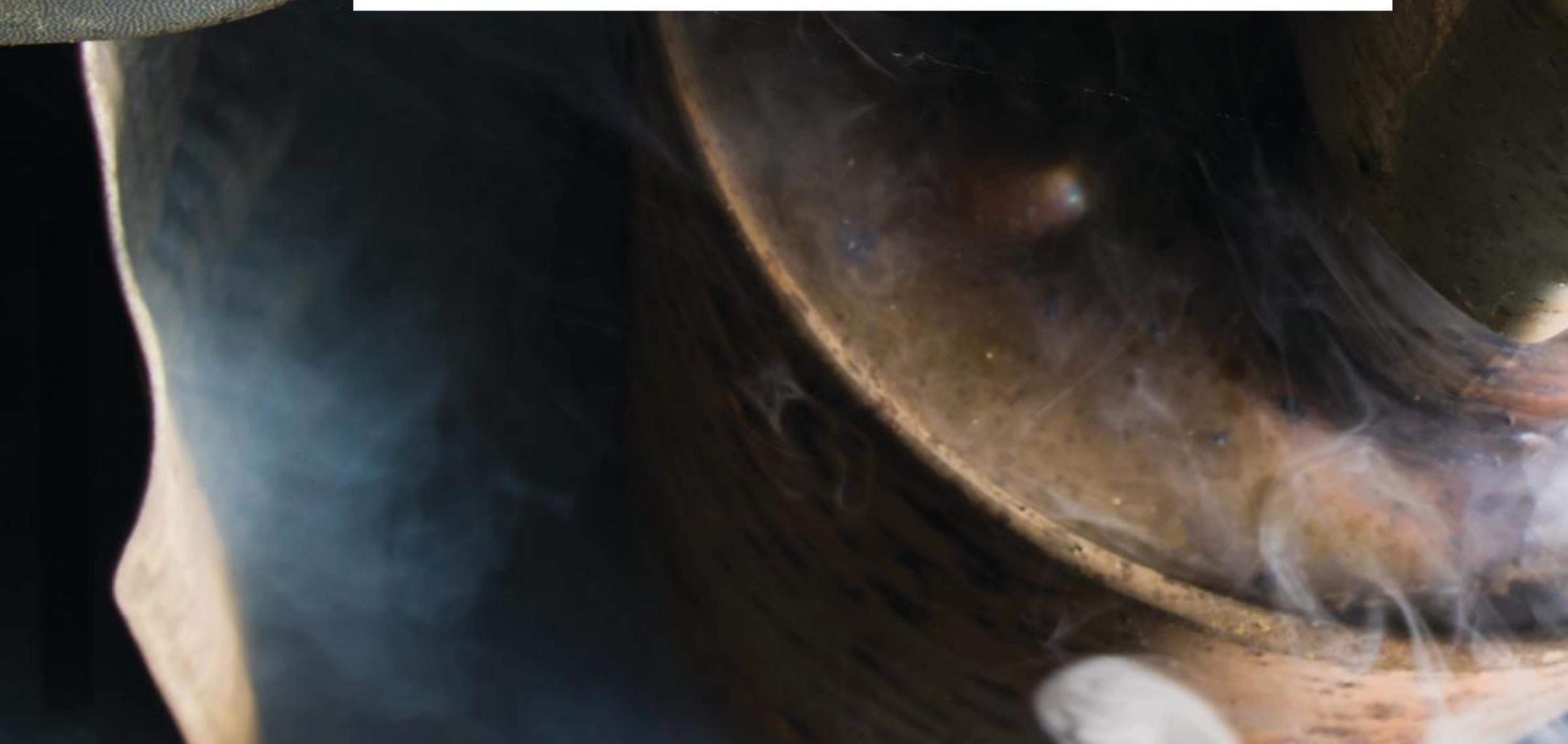


FIGURE 1 Smoke produced from the combustion of fuel is a mixture of carbon dioxide and water vapour.

GROUNDWORK

In Chapter 16, you will learn about the role of atmospheric gases in the greenhouse effect, how kinetic molecular theory explains the behaviour of gases, and how to solve calculations involving gases.

This chapter will build on concepts you have already learnt in Chapters 7 and 13, and Year 7 Science. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

16A Write a balanced equation for the combustion of propane.

16A Groundwork resource
Writing balanced chemical equations

16B Calculate the amount of substance (in mol) in 432 g of carbon dioxide.

16B Groundwork resource
Moles

16C What happens to the kinetic energy of particles when temperature increases?

16C Groundwork resource
Kinetic molecular theory

16D Explain what compressibility is in terms of gases.

16D Groundwork resource
Compressibility

PRACTICALS

16.2A PRACTICAL: CASE STUDY	How do hot-air balloons use our understanding of gas behaviour?	Page 522
16.2B PRACTICAL: CONTROLLED EXPERIMENT	How are pressure and volume related in a variable volume system?	pro
16.3 PRACTICAL: SIMULATION	How do ideal gases behave in a fixed volume system?	pro
16.4 PRACTICAL: CONTROLLED EXPERIMENT	How can we measure the volume of gas produced in a reaction?	pro

16.1 Gases contributing to greenhouse effect

KEY IDEAS

In this topic, you will learn that:

+ greenhouse gases in the atmosphere absorb and reflect infrared radiation
+ the greenhouse gas effect is a natural process in which infrared radiation is trapped by greenhouse gases, resulting in the warming of Earth
+ human activity can increase the amount of greenhouse gases in the atmosphere, accelerating global warming.

Atmospheric gases

atmosphere
the mixture of gases that surrounds Earth

The Earth's **atmosphere** is composed of various gases. Their quantities can vary, but currently, the composition of the atmosphere is:

- 78% nitrogen (N_2)
- 21% oxygen (O_2)
- 0.93% argon (Ar)
- 0.04% carbon dioxide (CO_2)
- 0.00017% or 1.7 ppm of methane (CH_4)
- a variable amount of water vapour (H_2O) depending on temperature.

This composition is especially important when we talk about the greenhouse effect and climate change. Therefore, it is very important to be able to measure and monitor the gases in the atmosphere.

Natural greenhouse gas effect

solar radiation
sunlight or waves of energy emitted from the Sun

infrared radiation
waves of energy emitted from Earth

greenhouse gases
the gases responsible for trapping the Sun's UV rays

The Sun naturally emits **solar radiation** (or, simply, sunlight) into Earth's atmosphere. This radiation is either absorbed by the land and oceans or reflected into space. Solar radiation absorbed by Earth can be re-radiated as **infrared radiation**. This can re-enter the atmosphere, where it can become trapped, or it can move beyond the atmosphere and back into space.

Atmospheric gases can absorb and reflect infrared radiation. They can prevent radiation from reaching Earth or trap radiation that has been reflected from the surface of Earth.

Several atmospheric gases are responsible for preventing radiation from escaping Earth's atmosphere. These are called **greenhouse gases** because they act like the glass exterior of a greenhouse. They are transparent. However, they trap infrared radiation and prevent it from leaving the atmosphere. Earth becomes warmer, like a greenhouse, so this is called the greenhouse effect.

The three major gases that contribute to the greenhouse effect are:

- carbon dioxide (CO_2)
- methane (CH_4)
- water vapour (H_2O).

The greenhouse gas effect is an important natural process. Without greenhouse gases, Earth's average temperature would be about −18°C. You can imagine that this would make the environment very difficult to live in. It would also kill a significant amount of plant and animal life. Greenhouse gases are also important for living things. For example, carbon dioxide is required as a reactant for photosynthesis, which generates energy in the form of glucose (sugar):

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

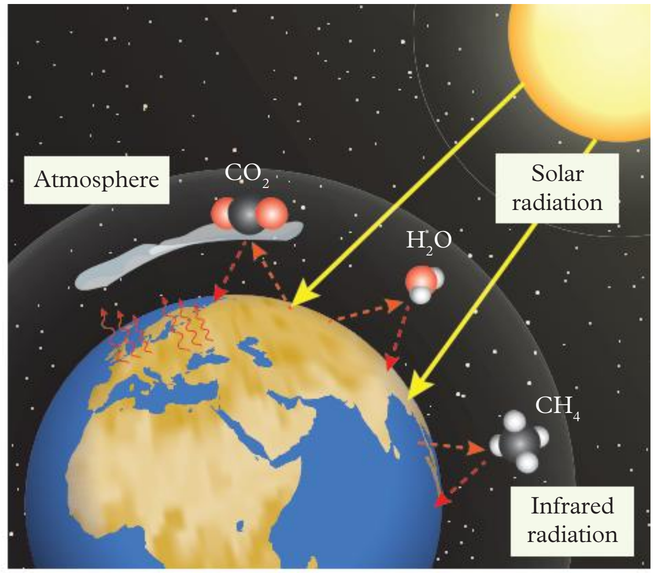


FIGURE 1 The absorption of infrared radiation by carbon dioxide (CO_2), water vapour (H_2O) and methane (CH_4)

Enhanced greenhouse gas effect

The greenhouse effect can be increased by human activity. In fact, humans are the greatest contributor to the increased amounts of greenhouse gases in the atmosphere present today. This is called the **enhanced greenhouse effect**.

Water vapour (gaseous H_2O) is Earth's most abundant greenhouse gas. It is responsible for clouds, rain and humidity. Temperature and the concentration of water vapour in the air are closely related. This is a positive feedback loop where the warming effect of water vapour keeps increasing (Figure 2).

enhanced greenhouse effect the accelerated warming of Earth due to human activities increasing the amount of greenhouse gases in the atmosphere

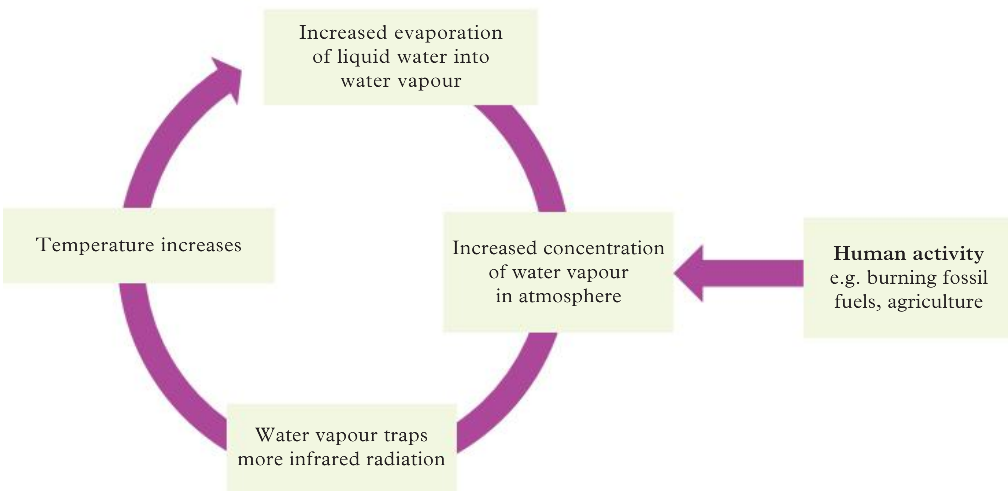


FIGURE 2 An increase in temperature increases the concentration of water vapour in the air. More water vapour can trap more infrared radiation, further warming the atmosphere.

Carbon dioxide and methane concentrations have also increased because of human activity. The concentration of carbon dioxide has almost doubled from pre-1800 levels (approximately a 1.5 times increase). Methane concentrations have more than doubled. In all of history, the concentrations of greenhouse gases within the atmosphere have never been so high.

Humans have contributed to the increases in atmospheric water vapour, carbon dioxide and methane through activities such as burning fossil fuels and agriculture. This enhances the global warming effect.

16.1 REAL-WORLD CHEMISTRY

Measuring atmospheric gases using ice cores

The first consistent measurements of the carbon dioxide concentration in the atmosphere began in 1958. However, data could not be obtained about atmospheric carbon dioxide before this time. To do so, scientists needed a new source of data to determine trends over longer periods of time. One key method, discovered in the 1950s, was drilling ice cores. These are cylinders of ice extracted from polar ice sheets and glaciers.

Polar ice forms in layers when snow deposits in the Artic and Antarctic. As temperatures increase in summer, these snow deposits melt, compact down and then freeze again (Figure 3). This cycle repeats annually to form layers in the ice. Air can become trapped between the layers.

Atmospheric scientists can determine the composition of the atmosphere by analysing the air bubbles trapped in the different layers. These form from gases in the atmosphere at the time the layer was frozen. The temperature of Earth's surface can also be determined when the layer of ice was frozen. This is because the atmospheric temperature is closely linked with concentrations of carbon dioxide (CO_2), methane (CH_4) and other greenhouse gases.

Ice cores have provided scientists with thousands of years' worth of climate data.

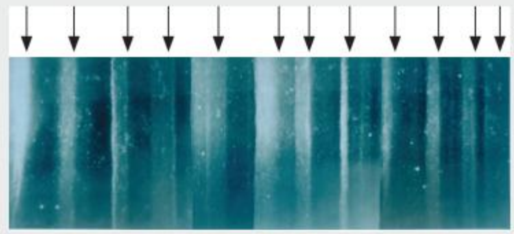


FIGURE 3 A close-up image of an ice core with layers formed in seasonal melting and freezing events. The summer layers are depicted with arrows.

Apply your understanding

1 Explain how ice cores can be used in climate science.

2 Using your understanding of the greenhouse effect, predict how atmospheric temperature is related to concentrations of greenhouse gases.

3 Discuss how analysing data from ice cores may help with addressing the United Nations Sustainable Development Goals.

16.1 CHECK YOUR LEARNING

Describe and explain

1 Explain what the greenhouse effect is and how it compares to a greenhouse.

2 Identify the most abundant greenhouse gases.

3 Explain how infrared radiation is trapped in the atmosphere.

Apply, analyse and compare

4 Describe the difference between the natural greenhouse effect and the enhanced greenhouse effect.

5 From trends in greenhouse gas emissions, explain what may happen to the climate if we do not reduce the amount of greenhouse gases in Earth's atmosphere.

Design and discuss

6 Discuss the importance of being able to measure gases in the atmosphere.

16.2 Gas pressure and standard laboratory conditions

KEY IDEAS

In this topic, you will learn that:

+ gases behave according to kinetic molecular theory
+ gases occupy the same volume under standard laboratory conditions (SLC)
+ standard laboratory conditions are 25°C and 100 kPa.

kinetic molecular theory
the theory that states that all particles are in constant random motion

kinetic energy
the energy associated with objects based on their movement or motion

You might remember the particle model of matter from Year 7 Science. This can be used to explain the behaviour of gases. In the gaseous state, particles do not stay together. They have a high level of energy, so they separate and move freely. The behaviour of gases can be explained by the **kinetic molecular theory**. In this topic, you will learn about the behaviour of gases and factors that affect this.

Kinetic molecular theory

The kinetic molecular theory explains the movement and behaviour of particles in solids, liquids and gases. **Kinetic energy** is the energy associated with the motion of particles. The more kinetic energy they have, the more they will move.

The kinetic molecular theory of gases states that gas particles:

- are in constant, random, straight-line motion
- have very little attraction to one another
- will spread out to occupy a container, meaning that they have a low density and can be compressed
- do not lose kinetic energy when they collide
- have more kinetic energy at higher temperatures.

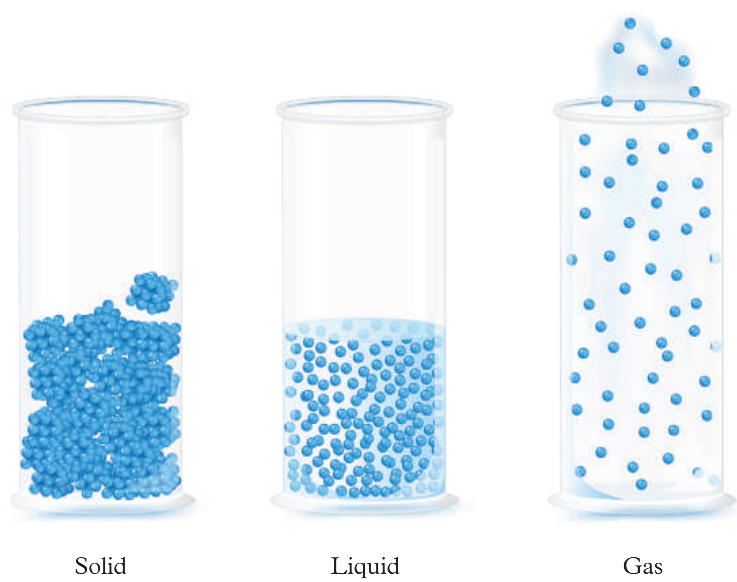


FIGURE 1 Unlike solids and liquids, gases expand to occupy the whole space of a container.

16.2A
PRACTICAL: CASE STUDY
How do hot-air balloons use our understanding of gas behaviour? Go to page **522**.

16.2B
PRACTICAL: CONTROLLED EXPERIMENT
How are pressure and volume related in a variable volume system? Go to your **Student obook pro**.

Factors affecting gas behaviour

A **system** includes the gas particles and the container that they occupy. There are three key factors that affect the kinetic energy of gases within a system:

- pressure
- temperature
- volume.

FIGURE 2 Pressure is generated when gas molecules collide with the internal walls of a balloon.

system
the gas particles and the container they occupy

gas pressure
the force exerted, per unit area, by a gas particle on another gas particle or the wall of a container

pascal
a unit of pressure that is the force of one newton per square metre of a substance; 1 kilopascal (kPa) is equal to 1000 pascals (Pa)

temperature
a measure of the average heat energy of the particles within a system

Pressure

Gas pressure is the force that gas particles exert when they collide with other gas particles and the walls of a container. This relationship is given by the equation:

$$P = \frac{F}{A}$$

where P is pressure, F is force and A is area.

Gas particles that have more kinetic energy will move more. This increases their chances of colliding with the walls of the container and the average force of each collision. If the area of the container is restricted, pressure increases. Pressure is measured in the unit **pascal** (Pa). This is the force of one newton per square metre of substance ($N\ m^{-2}$).

Temperature

All particles in a container have different amounts of heat energy. As the **temperature** of the system increases, the average heat energy of the system also increases. This causes the kinetic energy and movement of the particles to increase. The relationship between temperature and kinetic energy can be represented using the Maxwell–Boltzmann distribution (Figure 3). When the system is cold, there are more lower-energy particles present. When the system is hot, there are more higher-energy particles. The kinetic energy of any gaseous system is the same at a specified temperature.

Study tip

Increasing the number of gas molecules (concentration of the gas) in a fixed volume system will increase the number of collisions with the walls of the container. This increases the pressure of the system.

Study tip

Increasing the number of gas molecules (concentration of the gas) in a variable volume system will increase the collisions with the walls of the system and the volume of the container. Therefore, the pressure of the system remains constant.

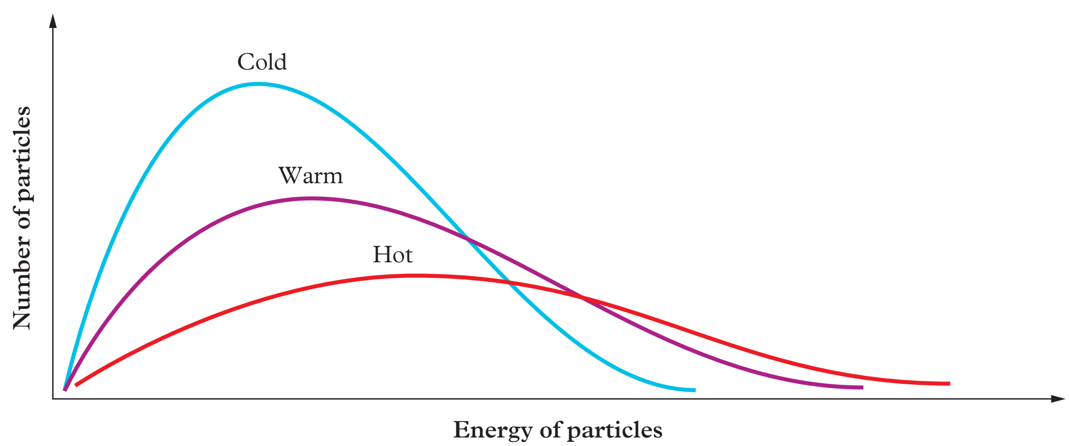


FIGURE 3 The Maxwell–Boltzmann representation of the amount of energy distributed in a reaction system at different temperatures

Volume

Gases occupy the whole container. Therefore, the **volume** of a container is equal to the volume of the gas inside that container. Volumes can occupy:

- **fixed volume or rigid wall systems** – such as gas cylinders, where the volume of the gas remains constant
- **variable volume systems** – such as a balloon, where the volume changes to accommodate the gas.

As volume decreases, there is less distance between gas particles. Because the frequency of collisions increases, the system will have more kinetic energy.

volume
a measure of the space occupied by a substance

fixed volume or rigid wall system
a system where the volume of the container does not change

variable volume system
a system where the volume of a container can change to accommodate the gas

Standard laboratory conditions (SLC)

To compare gases, scientists keep the systems at the same standard conditions. There are two types of standard conditions:

- **standard temperature and pressure (STP)**: 0°C (or 273 K) and 100 kPa, where one mole of any gas occupies 22.4 L
- **standard laboratory conditions (SLC)**: 25°C (or 298 K) and 100 kPa, where one mole of any gas occupies 24.8 L.

At the same temperature and pressure, all gases have the same average kinetic energy. Therefore, by keeping these factors the same (e.g. by using SLC), we can analyse differences in the number of moles and volumes occupied by a gas. In the next topic, we will look at calculating the number of moles and volumes of gases at SLC.

The pressure, temperature and volume of a gaseous system are related to each other by the ideal gas equation. Before you learn to use this equation, you need to be able to express pressure and temperature in the correct units.

standard temperature and pressure (STP)
standard conditions for temperature and pressure; gases are measured at 1 atm and 0°C or 273 K

standard laboratory conditions (SLC)
standard conditions for temperature and pressure under laboratory conditions; gases are measured at 1 atm and 25°C or 298 K

16.2 CHECK YOUR LEARNING

Describe and explain

1 Define 'gas pressure'.

2 Explain what standard laboratory conditions are.

3 Describe how kinetic molecular theory is used to explain the behaviour of gases.

Apply, analyse and compare

4 Use the kinetic molecular theory to explain why:

a there are spaces between gas particles

b gas particles have more energy than particles in a solid or liquid

c gases occupy the space of a container.

5 Use kinetic molecular theory to explain what happens to the kinetic energy of a gaseous system when:

a temperature is decreased

b pressure is increased

c volume is kept the same.

Design and discuss

6 When a scented candle is lit at one end of a room, the scent slowly diffuses to the opposite end of the room. Discuss how kinetic molecular theory explains this phenomenon.

Study tip

In VCE Chemistry, you will only need to remember and use SLC.

Study tip

At SLC, water is a liquid, so its state is (l).

16.3 The ideal gas equation

KEY IDEAS

In this topic, you will learn that:

+ pressure can be expressed in a range of different units, including Pa, kPa, atm and mmHg
+ pressure, temperature and volume are related by the ideal gas equation.

Unit conversions for gas calculations

kelvin
a unit of temperature that is based on levels of kinetic energy; one kelvin is equal to one degree Celsius

absolute zero
the temperature, in kelvin, at which kinetic energy is the lowest; 0 K

16.3A Worked example
Find me in your obook pro

16.3A Worked example
Video demonstration

atmosphere (atm)
a unit of pressure that is based on the atmospheric pressure at sea level; 1 atm is approximately 100 kPa, 100 000 Pa or 760 mmHg

millimetres of mercury (mmHg)
a unit of pressure that is based on the pressure needed to raise the volume of mercury by 1 mm in a manometer

So far, we have talked about SLC in terms of temperature in degrees Celsius (°C) and pressure in kilopascals (kPa). To use the ideal gas equation, we need to convert degrees Celsius to kelvin (K). You may also encounter pressure in the units: pascals (Pa), atmosphere (atm) and millimetres of mercury (mmHg).

Converting units of temperature

The **kelvin** is a unit of temperature commonly used in chemistry. The kelvin scale is based on the temperature at which kinetic energy is the lowest, which is at –273°C. This point is called **absolute zero** or 0 K.

The conversion between degrees Celsius and kelvin is simple. One degree on the Celsius scale is equal to one kelvin. This means that to get from degrees Celsius to kelvin, we simply add 273. To get from kelvin to degrees Celsius, we minus 273. See how this is done in Worked example 16.3A.

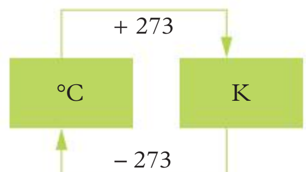


FIGURE 1 The conversion tool for converting between units of temperature

Converting units of pressure

At sea level, atmospheric pressure is measured as one **atmosphere** (1 atm). This is the equivalent to 101 325 Pa or 101.3 kPa. 1 atm is also the same as 760 mmHg (**millimetres of mercury**). A pressure of 1 mmHg is what is needed to increase the volume of mercury so that it travels 1 mm higher up the column of a manometer (a device for measuring pressure). You can use Figure 2 to help you with converting between Pa, kPa, atm and mmHg. Worked example 16.3B walks you through how to convert between units of pressure.

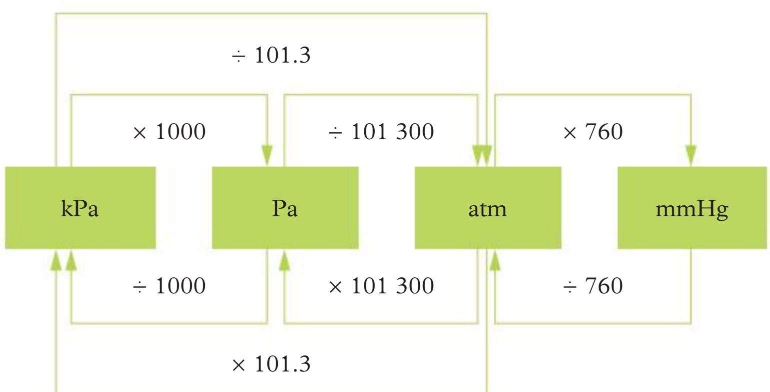


FIGURE 2 The conversion tool for converting between units of pressure

FIGURE 3 A manometer uses mercury to measure pressure.

16.3 PRACTICAL: SIMULATION

How do ideal gases behave in a fixed volume system? Go to your **Student obook pro**.

16.3B WORKED EXAMPLE

CONVERTING BETWEEN UNITS OF PRESSURE

Convert the following units.

a 92 kPa to Pa

b 635 kPa to atm

c 0.872 atm to kPa

d 7362 mmHg to atm

Solution

Think	Do
Step 1: Use the unit conversion tool to determine how much you need to multiply or divide the value by.	**a** To convert from kPa to Pa, we need to multiply by 1000. **b** To convert from kPa to atm, we need to divide by 101.3. **c** To convert from atm to kPa, we need to multiply by 101.3. **d** To convert from mmHg to atm, we need to multiply by 760.
Step 2: Convert the value.	**a** 92 kPa 1000 = 92 000 Pa **b** 635 kPa ÷ 101.3 = 6.27 atm **c** 0.872 atm × 101.3 = 88.3 kPa **d** 7362 mmHg ÷ 760 = 9.64 atm

Study tip

Although mmHg is not used in the ideal gas equation, it is still very handy to know.

Study tip

For units of volume, you only need to know how to convert between millilitres and litres.

Ideal gases

Most gases will behave like an **ideal gas** at SLC. In theory, this means that the gas particles do not collide with each other and do not occupy any volume. Assuming gases behave ideally at SLC makes it more straightforward to predict the behaviour of gases under different conditions.

Ideal gas equation

The **ideal gas equation** (also called the universal gas equation) makes it simpler for chemists to determine experimental values in gas experiments. It removes the need to do multiple calculations. The ideal gas equation takes three **gas laws**, which describe the relationships between temperature, pressure and volume, to form the equation:

$$PV = nRT$$

where P is pressure (in kPa), V is volume (in L), n is the amount of substance (in mol), T is temperature (in K) and the **universal gas constant** is $R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$.

ideal gas
a gas in which particles do not interact with one another and occupy no volume

ideal gas equation
an equation that relates temperature, pressure, volume and moles of a gas; also called the universal gas equation

gas laws
the laws that explain the properties of gases and how they behave at different temperatures, volumes and pressures

universal gas constant (R)
the proportionality constant (8.314 $\text{J mol}^{-1} \text{ K}^{-1}$) that is used to define gas behaviour under ideal conditions

16.3A CHALLENGE

Calculating the universal gas constant

The universal gas constant, 8.31 $\text{J mol}^{-1} \text{ K}^{-1}$, is used to define how gases behave under ideal conditions of temperature, pressure and volume. Calculate R at STP (1 atm, 0°C), by rearranging the ideal gas equation. To get your answer in $\text{J mol}^{-1} \text{ K}^{-1}$, you will need to convert the units first. Pressure needs to be in N m^{-2} ($1 \text{ N m}^{-2} = 1$ Pa), volume needs to be in m^3 ($1 \text{ m}^3 = 1000$ L), and temperature needs to be in kelvin. Show all your working.

Worked example 16.3C shows you how to use the ideal gas equation to calculate volume when you are given the amount of a gas in moles.

Study tip

For calculations using the ideal gas equation, pressure must be in kPa, volume must be in L, temperature must be in K and *R* must be in J mol^{-1} K^{-1}.

Study tip

Although the universal gas constant is calculated at STP, the study design only needs you to be able to complete calculations at SLC.

16.3C WORKED EXAMPLE

CALCULATING VOLUME FROM MOLES USING THE IDEAL GAS EQUATION

Calculate the volume that 0.94 mol of nitrogen (N_2) gas occupies at SLC.

Solution

Think	Do
Step 1: Rearrange the ideal gas equation to find the unknown (volume).	$V = \frac{nRT}{P}$
Step 2: Identify the values you have.	n = 0.94 mol R = 8.31 J mol^{-1} K^{-1} At SLC, T = 293 K and P = 100 kPa
Step 3: Plug in the values to find the unknown.	$V(N_2) = \frac{0.94 \times 8.31 \times 298}{100}$ = 23.28 L = 23 L (2 sig fig)

Moles can be expressed as:

$$n = \frac{m}{M}$$

This means that the ideal gas equation can be expanded to form the equation:

$$PV = \frac{mRT}{M}$$

See how this modified equation is used in Worked examples 16.3D and 16.3E.

Study tip

There can be many ways of answering these types of questions. As long as you are using the correct ideal gas equation, the precise method is not important.

16.3E Worked example
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16.3E Worked example
Video demonstration

16.3D WORKED EXAMPLE

CALCULATING VOLUME FROM MASS, USING THE IDEAL GAS EQUATION

Calculate the volume that 6.00 g of oxygen (O_2) gas occupies at SLC.

Solution

Think	Do
Step 1: Rearrange the ideal gas equation to find the unknown (volume).	$V = \frac{mRT}{PM}$
Step 2: Identify the values you have.	m = 6.00 g M = 32 g mol^{-1} R = 8.31 J mol^{-1} K^{-1} At SLC, T = 293 K and P = 100 kPa
Step 3: Plug in the values to find the unknown.	$V(O_2) = \frac{6.00 \times 8.31 \times 298}{100 \times 32}$ = 4.64 L (3 sig fig)

Have a look at Challenge 16.3B and Skill drill 16.3 to test your understanding. Don't forget to check out Real-world chemistry 16.3.

16.3B CHALLENGE

Who am I?

0.938 g of a hydrocarbon gas (containing hydrogen and carbon atoms only) occupies 764 mL at SLC. Identify the gas.

16.3 Real-world chemistry
Find me in your obook pro

16.3 SKILL DRILL

Evaluating the investigation method and validity of measurements

Key science skill: Analyse and evaluate data and investigation methods

A scientist sets up an experiment to determine the effect of decreasing the volume of a gaseous system (from 4 L to 2 L) on the pressure of the system. After the experiment, the scientist discovers that the temperature gauge had malfunctioned. The scientist started the experiment with 4 L on a 35°C day, before a power outage resulted in them having to pause the experiment and continue the following day. The next day was a cooler 25°C when the experiment was conducted using the 2 L system.

Practise your skills

1 Identify the independent variable.
2 Identify the dependent variable.
3 Identify only the controlled variables, which, if not controlled, will affect the dependent variable.
4 Determine whether the experiment is valid and justify your response.

Need help analysing and evaluating investigation methods? See Topic 1.8 (page 24).

16.3 CHECK YOUR LEARNING

Describe and explain

1 Identify the two requirements of an 'ideal' gas.
2 Using the ideal gas equation, explain what happens to:
   a volume when temperature is doubled
   b pressure when temperature is halved
   c volume when pressure is doubled.

Apply, analyse and compare

3 Convert the following units.
   a 315 K to °C
   b 74°C to K
   c 273 mmHg to kPa
   d 50 kPa to atm
   e 25 mL to L
4 Calculate the volume of:
   a 3.2 mol of water vapour at SLC
   b 2.9 g of oxygen gas at SLC
   c 9.37×10^{21} nitrogen molecules at SLC.
5 Calculate the:
   a mass of 120 mL of CO_2 at SLC
   b number of H atoms in 525 mL of CH_4 at SLC.

Design and discuss

6 Research the following three gas laws and describe how they are combined to form the ideal gas equation.
   a Gay–Lussac's law, which describes the relationship between temperature and pressure
   b Charles' law, which describes the relationship between temperature and volume
   c Boyle's law, which describes the relationship between pressure and volume

16.4 Chemical reactions involving gases

KEY IDEAS

In this topic, you will learn that:

- gases can be produced and consumed in chemical reactions
- stoichiometry can be used to determine the mass or volume of gases in chemical reactions.

Chemical reactions can both consume and produce gases. In this topic, we will explore a few types of chemical reactions involving gases, and then practise using stoichiometry to solve calculations related to these reactions.

Chemical reactions that consume and produce gases

Reaction types that consume and produce gases include:

- combustion
- synthesis and decomposition
- reactions of acids.

Combustion reactions

One important type of gas reaction is the combustion reaction. Combustion reactions involve oxygen. They are used to produce energy for transport, in machinery, batteries and within the body. The two main types of combustion reaction are combustion of metals to form metal oxides, and combustion of non-metals to form covalent compounds. Combustion of hydrocarbons, including many of the fuels we use, is a specific type of non-metal combustion.

Combustion of hydrocarbons

The (complete) combustion of a hydrocarbon typically forms carbon dioxide and water. This is a significant contributor to the greenhouse effect, especially if the hydrocarbon fuel used in the combustion reaction does not come from a renewable source. If the fuel is sourced from renewable plant material, such as ethanol, the carbon dioxide produced in the combustion of the fuel is theoretically consumed when the plants are regrown to produce more fuel (by photosynthesis):

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

The glucose produced in photosynthesis can be fermented to produce ethanol and carbon dioxide:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

Ethanol is one example of a fuel that can be sourced from plant material and therefore is renewable. When it combusts, ethanol produces carbon dioxide and water:

$$2C_5H_{10}OH(l) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$$

FIGURE 1 Ethanol is very commonly sourced from corn.

How can we measure the volume of gas produced in a reaction? Go to your **Student obook pro**.

Incomplete combustion of hydrocarbons

In some cases, carbon dioxide will not form from the combustion of a hydrocarbon. This occurs when the amount of oxygen available is low or limiting. In these cases, hydrocarbons undergo **incomplete combustion**. Less available oxygen will result in fewer oxygen–carbon bonds. This may lead to the formation of carbon monoxide (CO) or carbon (C).

Formation of CO: $2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(l)$
Formation of C: $C_3H_8(g) + 2O_2(g) \rightarrow 3C(s) + 4H_2O(l)$

incomplete combustion
a chemical reaction in which limiting oxygen reacts with a hydrocarbon to form either carbon monoxide or carbon and water

Synthesis and decomposition reactions

Synthesis and addition

A combination reaction, sometimes called **synthesis** or addition, involves reacting two substances to form one product. One example is the two-step reaction between sulfur dioxide, oxygen and water vapour to form sulfuric acid:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$
$$SO_3(g) + H_2O(g) \rightarrow H_2SO_4(aq)$$

Pollutants in the atmosphere that contain gaseous sulfur dioxide (SO_2), sulfur trioxide (SO_3), nitric oxide (NO) or nitrogen dioxide (NO_2) can therefore react with oxygen and water vapour to form acids. These precipitate from the atmosphere as acid rain.

synthesis
a chemical reaction in which two atoms, elements or molecules react to form a new molecule; also known as addition

decomposition
the breakdown of a molecule into smaller molecules

Decomposition

Decomposition reactions occur when a molecule breaks down into smaller molecules. This can happen when heat or electricity is applied. Carbonates, such as calcium carbonate or limestone, will commonly decompose when heated to produce carbon dioxide gas:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Reactions of acids

Acids also react to form gaseous products. One reaction is the corrosion of metals that you learnt about in Chapter 12. For example, iron corrodes when exposed to acid rain to form a salt and the very flammable hydrogen gas.

$$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$$

Metal carbonates can also react with acids to form a salt, water and carbon dioxide. For example, the reaction between limestone and acid rain produces carbon dioxide, a big contributor to the greenhouse effect.

$$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

FIGURE 2 Pollution generated by factories can form acid rain.

Using stoichiometry for gases

Stoichiometry can be used to determine the mass or volume of gases consumed or produced in chemical reactions.

Solving stoichiometry problems

Substances are often considered in quantities of mass, volume or density. Petrol is pumped into a car as a liquid and measured in units of volume. The gas in a gas cylinder is also measured by volume. However, wood – an impure solid – is measured in units of mass. All of these can undergo reactions that consume or form gases. For these reactions, we need to be able to calculate the amount (in mol) of a substance from mass or volume.

FIGURE 3 Acid rain is harmful to the environment.

Revision of stoichiometry

Stoichiometric problems can be solved in four steps.

1 Write a balanced chemical equation for the reaction, identifying the known (given) and unknown (required) quantities of substance. The known quantity is the one where you can calculate the number of moles. The unknown quantity is the substance of interest in the question.

2 Calculate the amount (in mol) of the known quantity of substance present. This may involve one calculation or multiple calculations to determine the number of moles of a substance.

3 From the equation, find the molar ratio that states the proportion of known to unknown quantities in the reaction and use it to calculate the number of moles of the required substance. This is called a **ratio statement**.

4 Calculate the quantity of the unknown substance. This may involve one or multiple calculations.

See how stoichiometry is applied to reactions involving gases in Worked examples 16.4A, 16.4B and 16.4C. You can also practise applying your understanding of the sustainability concepts to reactions of gases in Skill drill 16.4.

Study tip

If the formula is not in the data book, you can develop your own from the units. For example, if the unit of density (*d*) is g mL^{-1}, the formula would be density = mass (g) ÷ volume (mL).

ratio statement an equation used to determine the quantity of an unknown substance (usually in moles) relative to another substance in the same chemical reaction

Study tip

Remember to divide down to 1 if you struggle with mole ratios in chemical equations. This is demonstrated in part a of Worked example 16.4A.

16.4 SKILL DRILL

Applying sustainability concepts to reactions that release greenhouse gases

Key science skill: Analyse, evaluate and communicate scientific ideas

There is a heavy focus on sustainability in VCE Chemistry. Therefore, it is important that you practise drawing links between the chemistry content and green chemistry principles, the United Nations Sustainable Development Goals and linear versus circular economies. Measuring gases is very relevant to sustainability.

FIGURE 4 Green chemistry is a big focus in the new study design.

Practise your skills

1 Identify which green chemistry principles relate to having a good understanding of stoichiometry and calculations involving greenhouse gases.

2 Discuss how the ability to measure greenhouse gases helps with addressing the United Nations Sustainable Development Goals.

3 Discuss how understanding greenhouse gas emissions is significant for shifting from a linear economy to a circular one.

4 List three reliable scientific or media sources where you could find information about climate challenges. Justify your choices.

Need help applying sustainability concepts? See Topic 1.11 (page 32).

16.4A Worked example
Find me in your obook pro

16.4A Worked example
Video demonstration

Study tip

Make sure you have the right units before you do your calculations.

16.4C WORKED EXAMPLE

CALCULATING THE AMOUNT (IN MOL) OF CO_2 FORMED FROM ACID-BASE AND DECOMPOSITION REACTIONS

CO_2 is formed when the products of an acid–base reaction further decompose:

$2HCl(aq) + Na_2CO_3(aq) \rightarrow H_2CO_3(aq) + 2NaCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$

What quantity (in mol) of CO_2 is formed if 4.3 g of Na_2CO_3 is reacted?

Solution

Think	Do
Step 1: Calculate the number of moles of Na_2CO_3.	$n(Na_2CO_3) = \frac{m}{M} = \frac{4.3}{106} = 0.0406$ mol
Step 2: Use a ratio statement to determine the number of moles of CO_2 formed.	$n(CO_2) = n(Na_2CO_3) = 0.0406$ mol $= 0.041$ mol (2 sig fig)

16.4B Worked example
Find me in your obook pro

16.4B Worked example
Video demonstration

16.4 CHECK YOUR LEARNING

Describe and explain

1 Describe what a combustion reaction is.

2 Explain how chemical reactions can contribute to the greenhouse effect.

3 Explain why not all gas-producing reactions contribute to the greenhouse effect.

Apply, analyse and compare

4 Compare complete and incomplete combustion.

5 Write a ratio statement, and then determine the amount (in mol) of CO_2 in the following reactions.

a 5 moles of HCl are consumed in an acid–base reaction.

$2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

b 7.4 moles of CH_4 are combusted.

$CH_4(g) + O_2(g) \rightarrow H_2O(l) + CO_2(g)$

c 12 moles of glucose are consumed in photosynthesis.

$C_6H_{12}O_6(aq) + O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

6 Calculate the:

a amount (in mol) of CO_2 produced when 5 L of propane (density = 1.97 g L^{-1}) are completely combusted

b mass (in grams) of CO_2 required to generate 3.6 moles of oxygen from photosynthesis

c mass (in grams) of O_2 required to react with glucose in cellular respiration to form 253 g of water

d the volume (in litres) of octane (density = 703.3 g L^{-1}) required to produce 400 g of CO_2 by complete combustion.

7 Calculate the mass (in grams) of the following chemicals in the reactions shown:

a CO produced when 6.45 moles of oxygen are available to react:

$2C_3H_8(g) + 9O_2(g) \rightarrow 4CO_2(g) + 2CO(g) + 8H_2O(l)$

b SO_3 consumed to produce 23.1 moles of sulfuric acid:

$SO_3(g) + H_2O(g) \rightarrow H_2SO_4(aq)$

c Mg required to produce 300 g of hydrogen gas:

$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(aq)$

Design and discuss

8 Discuss the benefits or impacts of using fuels that are sourced from plant materials in terms of their:

a renewability

b carbon neutrality

c environmental impact.

9 Discuss the negative impacts of gaseous pollutants in the atmosphere, using balanced chemical equations to support your answer.

16.5 Calculating molar volume or mass of a gas

KEY IDEAS

In this topic, you will learn that:

+ at the same temperature and pressure, all gases have the same molar volume.

molar volume
the volume occupied by one mole of a gas at a specified temperature and pressure

The relationship between pressure, volume and temperature can be extended to the number of atoms or molecules in a gaseous system.

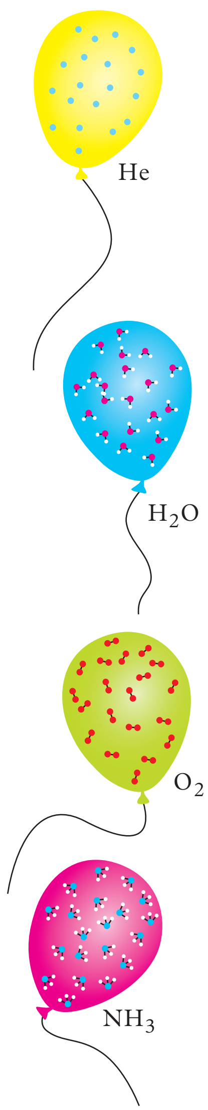

FIGURE 1 Four balloons of the same volume, at the same temperature and pressure, have the same number of gas particles.

Molar volume

Avogadro hypothesised that under identical conditions of temperature and pressure, equal volumes of gases contain an equal number of gas particles. Therefore, one mole of any gas (6.02×10^{23} molecules), at the same temperature and pressure, will occupy a specific volume (Figure 1). This is called the **molar volume**.

The relationship between moles and volume is represented by the equation:

$$n = \frac{V}{V_m}$$

where n is the amount of substance (in mol), V is the volume (L) and V_m is the molar volume (L mol^{-1}).

You might remember from Topic 16.2 that one mole of gas occupies 22.4 L at STP and 24.8 L at SLC. These values are the molar volumes.

This means that at STP, the equation is: $n_{STP} = \frac{V}{22.4}$

At SLC, the equation becomes: $n_{SLC} = \frac{V}{24.8}$

Worked examples 16.5A–16.5D show you a number of ways that molar volume can be used in calculations.

16.5A WORKED EXAMPLE

CALCULATING THE MASS OF GASES AT SLC

What is the mass (in g) of helium in a 20.0 L cylinder at SLC?

Solution

Think	Do
Step 1: Substitute the values into the equation.	$n(\text{He}) = \frac{V}{V_m}$ $= \frac{20.0}{24.8}$ $= 0.806$ mol
Step 2: Calculate the mass of the gas.	$m(\text{He}) = n \times M$ $= 0.806 \times 4.0$ $= 3.23$ g (3 sig fig)

16.5B WORKED EXAMPLE

CALCULATING THE VOLUME OF CO_2 RELEASED BY COMBUSTION AT SLC

What volume of carbon dioxide gas is generated from the combustion of a 4.0 L butane gas cooker at SLC?

Solution

Think	Do
Step 1: Write a balanced equation for the combustion of butane gas.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$ Note that at SLC, water is a liquid.
Step 2: Calculate the number of moles of butane.	$n(C_4H_{10}) = \frac{V}{V_m} = \frac{4.0}{24.8} = 0.161$ mol
Step 3: Calculate the moles of carbon dioxide using a ratio statement.	$n(CO_2) = n(C_4H_{10}) \div 2 \times 8 = 0.645$ mol
Step 4: Calculate the volume of carbon dioxide.	$V(CO_2) = n \times V_m = 0.645 \times 24.8 = 16$ L (2 sig fig)

16.5C Worked example
Find me in your obook pro

16.5C Worked example
Video demonstration

16.5D Worked example
Find me in your obook pro

16.5D Worked example
Video demonstration

16.5 CHECK YOUR LEARNING

Describe and explain

1 Explain what molar volume is.

Apply, analyse and compare

2 Calculate the volume at SLV of:
- **a** 3.2 mol of water vapour
- **b** 2.9 g of oxygen gas
- **c** 9.37×10^{21} nitrogen molecules.

3 Calculate the:
- **a** mass of 120 mL of CO_2 at SLC
- **b** number of H atoms in 525 mL of CH_4 at SLC.

4 Calculate the volume of oxygen required to fully combust 1.50 L of methane at SLC.

5 Calculate the volume of carbon dioxide generated from the combustion of the contents of a 20.0 L propane tank at SLC.

6 A car has a petrol tank that holds 40 L of octane (density = 0.703 g mL^{-1}). When the octane combusts, it reacts completely with an excess of oxygen. A scientist wishes to determine the volume of ethanol that will react with the same amount of oxygen that the 40 L octane tank has consumed.
- **a** Calculate the mass of the octane.
- **b** Write a balanced chemical equation for the combustion of octane.
- **c** Calculate the mass of oxygen consumed in the reaction.
- **d** Write a balanced chemical equation for the combustion of ethanol.
- **e** Assuming the mass of oxygen calculated in part **c** is the same mass of oxygen that will react with ethanol, calculate the mass of ethanol.
- **f** Calculate the volume of ethanol if it has a density of 0.789 g mL^{-1}.

7 Determine the mass of xenon gas that can be used to fill a 1.4 L cylinder at SLC.

Design and discuss

8 Using an example, prove that increasing the temperature of a gas will increase its molar volume.

9 Predict the impact of decreasing pressure on the molar volume of a gas.

Study tip

The study design only requires you to be able to complete calculations at SLC.

Study tip

The molar volume of any gas at 25°C and 100 kPa (SLC) is 24.8 L mol^{-1} If you forget, you can find these values in the data book.

CHAPTER

16 Review

Chapter summary

16.1

- Greenhouse gases absorb infrared radiation and warm Earth. This is called the greenhouse effect.
- The three major greenhouse gases are carbon dioxide, methane and water vapour.
- The greenhouse effect is a natural process. Human activity has contributed to an enhanced greenhouse effect. This is due to an increase in greenhouse gas emissions caused by the burning of fossil fuels and an increase in agriculture.

16.2

- The behaviour of gases can be explained by the kinetic molecular theory. The three key factors that affect gas behaviour are pressure, temperature and volume.
- Gas pressure is measured as the force exerted, per unit of area, on the walls of a container.
- Standard laboratory conditions (SLC) are used to compare the behaviour of different gases. At SLC, temperature is 25°C, pressure is 100 kPa, and one mole of any gas occupies 24.8 L.

16.3

- Gas pressure can be expressed in pascals (Pa), kilopascals (kPa), atmospheres (atm) and millimetres of mercury (mmHg). 1 atm = 100 kPa = 760 mmHg.
- Temperature can be expressed in degrees Celsius (°C) or kelvin (K). 0 K = –273°C.
- The ideal gas equation is used to calculate the pressure, volume or temperature of a gas using the universal gas constant, R, which is equal to 8.314 J mol^{-1} K^{-1}.

16.4

- Combustion reactions are the main contributors to greenhouse gas emissions because they produce carbon dioxide and water vapour.
- Gases can also be consumed or produced in synthesis, decomposition and corrosion reactions.
- Stoichiometry can be used to predict the volume, mass or moles of gases produced or consumed in chemical reactions.

16.5

- Molar volume is the volume that one mole of a gas occupies at the same pressure and temperature. It can be used to calculate the volume or mass of a gas produced in a chemical reaction.

Key formulas

Pressure	$P = \frac{F}{A}$
Moles	$n = \frac{m}{M}$
Ideal gas equation	$PV = nRT$ or $PV = \frac{mRT}{M}$
Relationship between number of moles and volume	$n = \frac{V}{V_m}$ $n_{SLC} = \frac{V}{24.8}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as 'Confidently', 'Mostly' or 'Not really'. If you're not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can...	Confidently	Mostly	Not really	Revision link
• explain what the natural greenhouse effect and enhanced greenhouse effect are, and identify the major greenhouse gases	☐	☐	☐	Go back to Topic 16.1
• define gas pressure and standard laboratory conditions (SLC)	☐	☐	☐	Go back to Topic 16.2
• convert between units of pressure, including Pa, kPa and atm	☐	☐	☐	Go back to Topic 16.3
• convert between units of temperature, including °C and K	☐	☐	☐	Go back to Topic 16.3
• use stoichiometry to solve calculations related to chemical reactions involving gases, including moles, mass and volume of gases	☐	☐	☐	Go back to Topic 16.4
• use molar volume to calculate the volume and molar mass of a gas produced by a chemical reaction	☐	☐	☐	Go back to Topic 16.5

Revision questions

Multiple choice

1 Kinetic molecular theory does *not* state that:

A particles move in straight lines.
B particle movement is random.
C particle movement is circular.
D particle collisions are elastic and conserve energy.

2 What are standard laboratory conditions?

A 0°C and 100 kPa
B 25°C and 100 atm
C 0°C and 101.3 kPa
D 25°C and 0.987 atm

3 Which of the following are the major greenhouse gases?

A Carbon monoxide, water vapour and methane
B Carbon dioxide, water vapour and methane
C Oxygen, carbon dioxide and water vapour
D Methane, oxygen and carbon dioxide

4 The greenhouse effect occurs because gases trap:

A electromagnetic radiation.
B solar radiation.
C UV radiation.
D infrared radiation.

5 The pressure of a gaseous system will increase when:

A temperature increases.
B volume is reduced.
C moles of gas are increased.
D all of the above.

6 Which of the following includes the correct units for the ideal gas equation?

A Pressure in kPa, volume in L, temperature in K
B Pressure in Pa, volume in L, temperature in K
C Pressure in kPa, volume in mL, temperature in K
D Pressure in kPa, volume in mL, temperature in °C

7 73 K is equal to:

A –200°C.

B 0°C.

C 346°C.

D 200°C.

8 Which of the following statements about chemical reactions involving gases is *false*?

A Complete and incomplete combustion of hydrocarbon gases can form different products.

B Stoichiometry can be used to calculate the mass of gas formed from all reactions apart from decomposition reactions.

C Synthesis reactions are also known as addition reactions.

D Incomplete combustion occurs when oxygen availability is low.

9 The mass of CO_2 produced in the combustion of 2.0 L of butane at SLC is:

A 14.2g

B 3.5 g

C 0.9 g

D 2.0 g

10 What is the amount (in mol) of oxygen gas in a 32.0 L cylinder at SLC?

A 1.43 mol

B 0.775 mol

C 0.700 mol

D 1.29 mol

Short answer

Describe and explain

11 Define 'gas pressure'.

12 Explain why standard laboratory conditions are used.

13 Explain why the natural greenhouse effect is important for life on Earth.

14 Describe the effect of an increase in temperature on the pressure and volume of a:

a fixed volume system

b variable volume system.

15 Describe the effect of an increase in pressure on the temperature of a fixed volume system.

16 Using kinetic molecular theory, explain why the molar volume of a gas is higher at SLC than at STP.

Apply, analyse and compare

17 A scientist wishes to inflate four balloons with the same number of molecules of four different gases. Determine the volume (in L) needed to occupy each balloon if they are filled with hydrogen (H_2), oxygen (O_2), helium (He) and argon (Ar).

18 Calculate the volume (in L) of propane gas occupied by 106.45 g at SLC.

19 Calculate the mass (in g) of butane that occupies a 20.0 L gas cylinder at SLC.

20 Calculate the mass (in g) of carbon dioxide produced from the complete combustion of a 60.0 L tank of octane at SLC.

21 3.00 L of propane undergoes incomplete combustion at SLC according to the equation:

$$2C_3H_8(g) + 9O_2(g) \rightarrow 4CO_2(g) + 2CO(g) + 8H_2O(l)$$

Calculate the:

a amount (in mol) of CO_2 produced

b amount (in mol) of oxygen consumed

c mass of CO_2 (in g) produced

d mass of CO (in g) produced

e volume of water (in mL) produced (density = 1 g mL^{-1}).

22 When an excess of oxygen is available, 3.00 L of propane can undergo complete combustion at SLC according to the following equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

a Calculate the mass of CO_2 (in g) produced.

b Determine how much more CO_2 (in g) is produced by complete combustion of propane than by incomplete combustion of propane.

23 Which of the following chemical reactions produces a greater mass of carbon dioxide?

I 5 moles of glucose are consumed in cellular respiration according to the equation:

$$C_6H_{12}O_6(aq) + 6\,O_2(g) \rightarrow 6\,CO_2(g) + 6\,H_2O(l)$$

II 5 moles of glucose are fermented according to the equation:

$$C_6H_{12}O_6(aq) \rightarrow 2\,C_2H_5OH(l) + 2\,CO_2(g)$$

24 Calculate the mass (in g) of carbon dioxide produced from the complete combustion of a 20.0 L cylinder of gas containing a mixture of 40% propane and 60% butane at SLC.

25 Calculate the mass (in g) of 300 mL of CH_4 at SLC.

26 Calculate the mass (in g) of CO_2 produced when 3.9 L of butane is completely combusted.

27 Calculate the mass (in g) of O_2 generated from 5.34 moles of carbon dioxide in photosynthesis.

28 Calculate the volume (in L) of octane (density = 703.3 g L^{-1}) required to produce 6.4 moles of CO_2.

29 Calculate the volume (in L) of:

a 5.3 g of water vapour at SLC

b 4.44 moles of oxygen gas at SLC

c 19.0 g of neon gas at SLC

d 4.522×10^{22} nitrogen molecules at SLC

e 3.2 g of carbon dioxide at SLC

f 5.2 g argon gas at SLC.

30 Use the following information to determine the identity of the monoatomic gases.

a 8.21 g of an unknown monoatomic gas fills a balloon to 925 mL at SLC.

b 380 mg of an unknown monoatomic gas fills a balloon to 236 mL at SLC.

c 1.29 g of an unknown monoatomic gas fills a balloon to 800 mL at SLC.

Design and discuss

31 The incomplete combustion of a hydrocarbon fuel generates less carbon dioxide than incomplete combustion does. Therefore, it may appear to contribute less to the greenhouse effect. However, incomplete combustion also produces carbon monoxide. Research the effects of carbon monoxide on humans and the environment.

32 A student takes a plastic water bottle to school on a snowy winter's day. The student travels to school by car and the car heater is on. It takes almost an hour to get to school. On arriving at school, the student drinks a mouthful of water. A small amount of gas exits the bottle when it is opened. At lunchtime, the student takes the water bottle outside and accidentally leaves it there. The sides of the bottle become crumpled in. Discuss what has happened to the water bottle to cause both effects.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

CHAPTER

17 Analysis for salts

KEY KNOWLEDGE

→ sources of salts found in water or soil (which may include minerals, heavy metals, organometallic substances) and the use of electrical conductivity to assess the salinity and quality of water or soil samples

→ quantitative analysis of salts:

→ molar ratio of water of hydration for an ionic compound

→ the application of mass–mass stoichiometry to determine the mass present of an ionic compound

→ the application of colorimetry and/or UV–visible spectroscopy, including the use of a calibration curve to determine the concentration of ions or complexes in a water or soil sample

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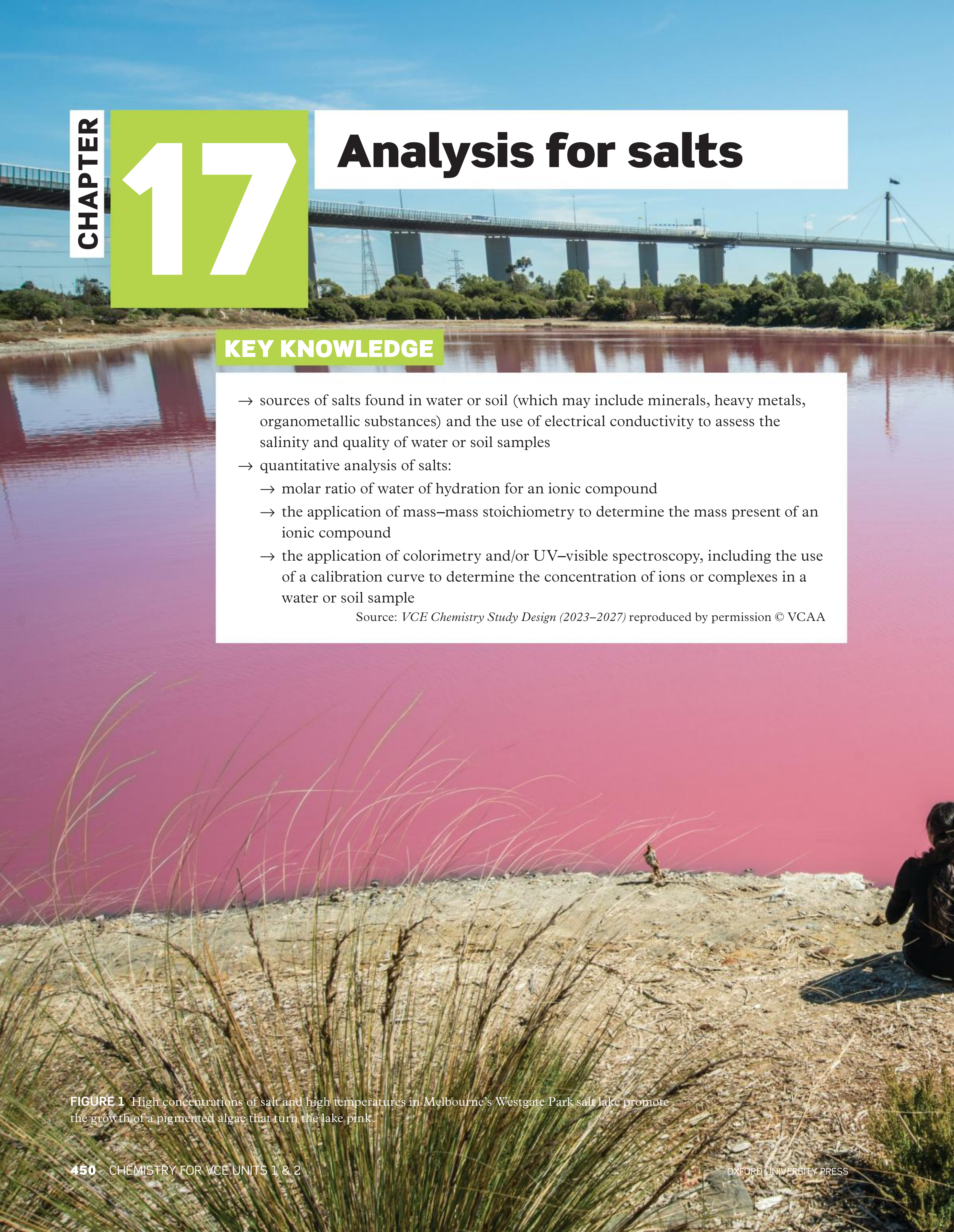
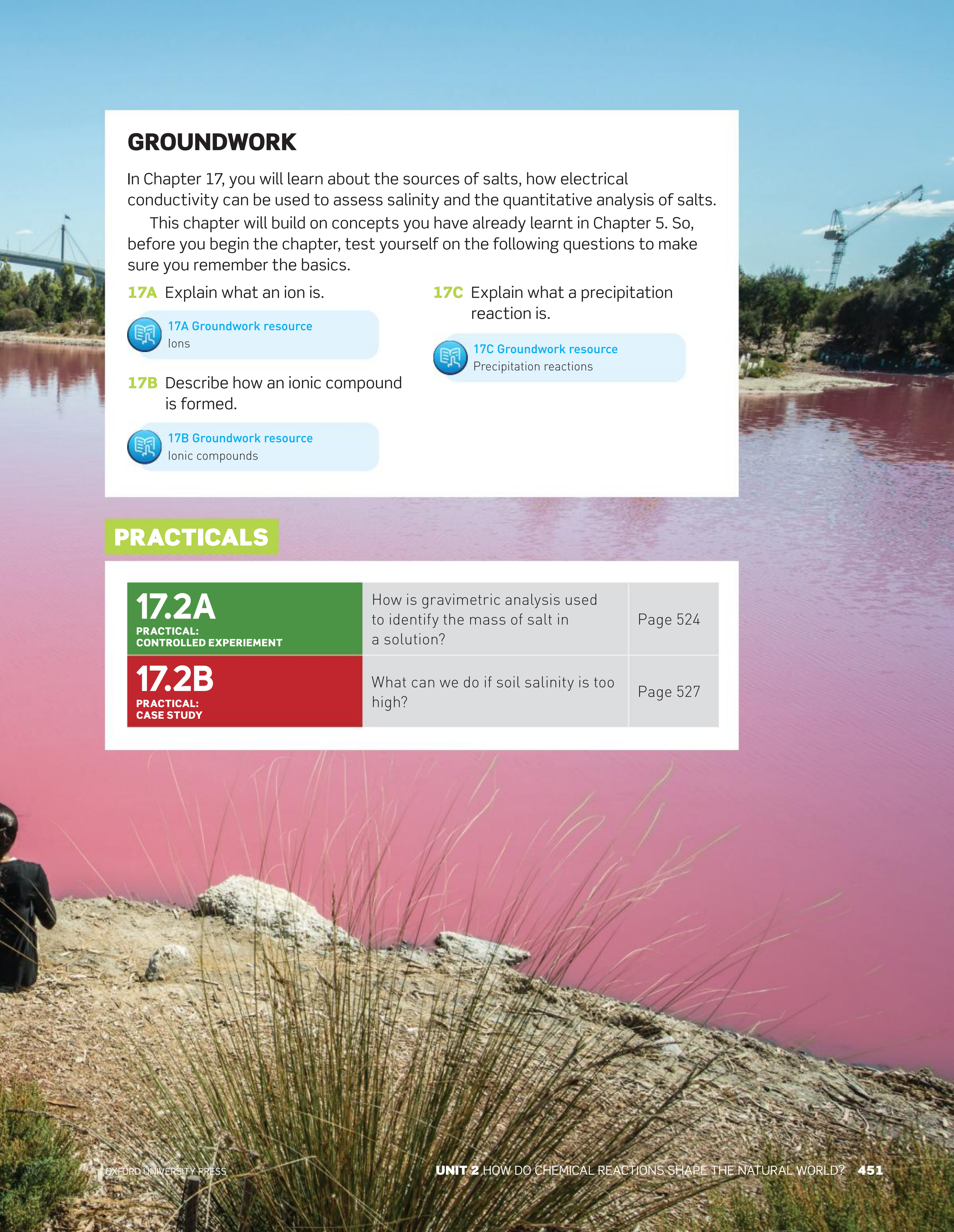


FIGURE 1 High concentrations of salt and high temperatures in Melbourne's Westgate Park salt lake promote the growth of a pigmented algae that turn the lake pink.



GROUNDWORK

In Chapter 17, you will learn about the sources of salts, how electrical conductivity can be used to assess salinity and the quantitative analysis of salts.

This chapter will build on concepts you have already learnt in Chapter 5. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

17A Explain what an ion is.

17A Groundwork resource
Ions

17B Describe how an ionic compound is formed.

17B Groundwork resource
Ionic compounds

17C Explain what a precipitation reaction is.

17C Groundwork resource
Precipitation reactions

PRACTICALS

17.2A PRACTICAL: CONTROLLED EXPERIEMENT	How is gravimetric analysis used to identify the mass of salt in a solution?	Page 524
17.2B PRACTICAL: CASE STUDY	What can we do if soil salinity is too high?	Page 527

17.1 Sources of salts in water and soil

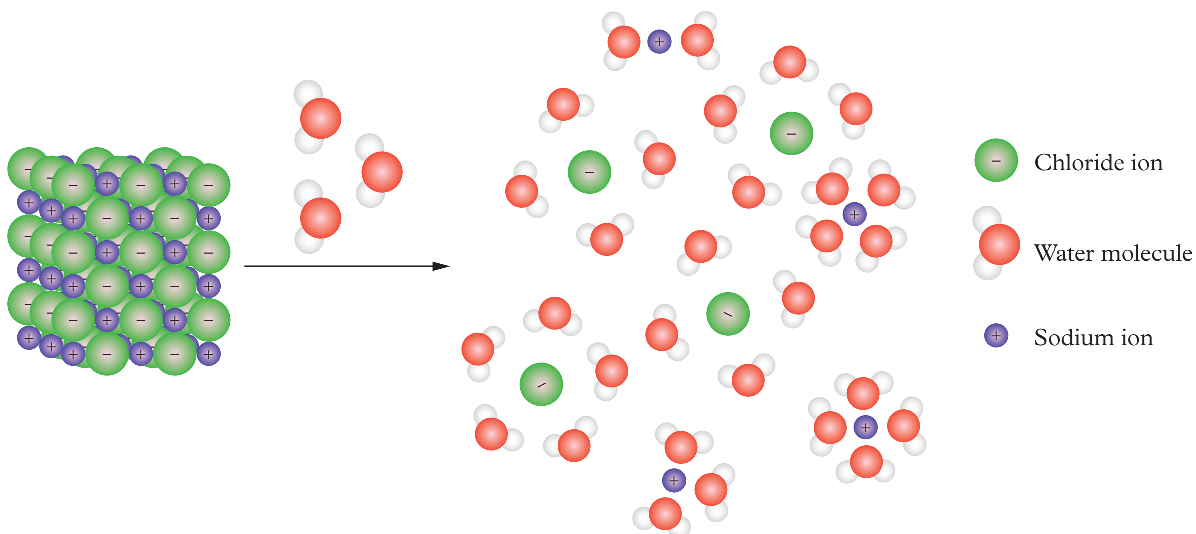
KEY IDEAS

In this topic, you will learn that:

- salinity is the measure of concentration of salts
- there are many sources of salts in water and soils
- electrical conductivity can be used to assess the salinity of a water or soil sample
- the level of salinity can affect water quality for specific uses, such as drinking or irrigation.

salinity the concentration of salt in water or soil

Salinity is the measure of the concentration of salts in water or soils. As you learnt in Chapter 5, salts are ionic compounds made up of a metal cation (positive ion) and a non-metal anion (negative ion), as shown in Figure 1. Salts are highly soluble in water, which means they are easily transported with the movement of water.



$$NaCl \xrightarrow{H_2O} Na^+ + Cl^-$$

FIGURE 1 Salts are formed from anions and cations and are highly soluble in water.

There are two main types of salinity:

- **primary salinity** – salts occurring naturally in the landscape; for example. salt deposits that are stored deep in soils, surface salt deposits, salt marshes or salt lakes
- **secondary salinity** – additional salt added to the soil or water through human activity; for example, urbanisation, agriculture, vegetation clearing, irrigation and industrial practices.

groundwater the water beneath the surface in soil, in rocks or soil pore spaces

Under normal conditions, the deep roots of native plants absorb water before the water makes it through the salts stored in deep soils. This prevents excess salt from entering the **groundwater** and increasing the concentration of salt in the water cycle. However, urbanisation, land clearing and poorly managed irrigation can lead to salinity increasing to excessive levels that can affect drinking water, agriculture and ecosystem health.

Sources of salts found in water

Salts can come from a variety of different sources. Primary salinity can result from natural processes such as weathering of rocks, ocean spray, or wind and rain depositing salt over thousands of years.

Secondary salinity is more concerning, and is often a result of clearing native, deep-rooted trees and altered land use. Secondary salinity can be categorised into two types:

- **Dryland salinity** occurs when deep-rooted native plants are removed or replaced with shallow-rooted plants that require less water, leading to excess water entering the groundwater.
- **Irrigation-induced salinity** occurs when excess water is added to crops and it travels past the root zone and into the groundwater.

Both dryland and irrigation-induced salinity allow increased amounts of water to pass through the deep soils that have high concentrations of salts. The water absorbs the salt as it is added into the groundwater system, or excess water goes into the groundwater stores and causes the **water table** to rise, bringing excess salts to the surface where it can be left behind when the water evaporates.

water table
the underground boundary between the soil and the area where groundwater saturates spaces between sediments and cracks in rock

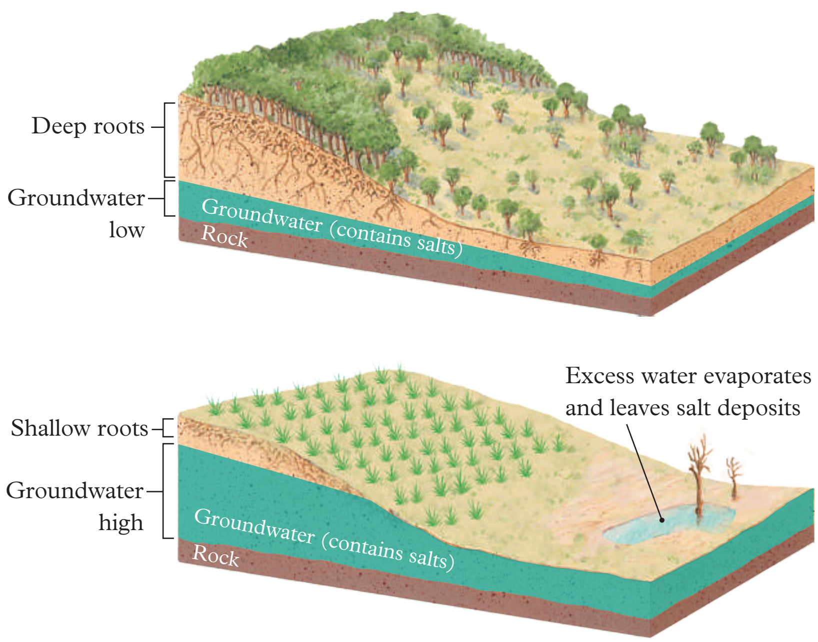


FIGURE 2 Plants with shallow roots don't absorb as much water as deep-rooted plants, causing excess water to be added into the water table with higher concentrations of salts.

FIGURE 3 Irrigation-induced salinity can negatively affect agricultural activities, reducing production and the productivity of land.

Salts can also be introduced into the water system from human activity such as:

- **mining** – large quantities of water is used to process the materials extracted from the earth. Some of the water containing various ions can be released back into waterways
- **agriculture** – the use of fertilisers such as ammonium nitrate (NH_4NO_3), urea ($CO(NH_2)_2$), and ammonium sulfate ($(NH_4)_2SO_4$) on crops is common. But when it rains, fertilisers can dissolve and enter the waterways, which contributes to excess build-up of these ions in the waterways
- **cleaning** – until recently, phosphates were added to detergents as softening agents. Wastewater from washing machines and sinks containing these detergents would enter the water system and cause a build-up of the phosphate anion (PO_4^{3-}).

Types of salts

When we hear the term 'salt', we often think of sodium chloride (NaCl). However, salt can be any ionic compound, such as mineral salts, heavy metals and organometallic substances.

Mineral salts

Mineral salts can enter the waterways naturally when rain and wind dissolve rocks and salt sediments in soils and carry them into the water table and rivers, lakes and creeks.

In our drinking water, minerals and ions, such as Mg^{2+}, Na^{+}, Cl^{-}, NO_3^{-}, K^{+} and Ca^{+}, occur at different levels, depending on factors such as the source of the drinking water; levels of water reservoirs; time of year (summer creates more demand for water, so water travels faster through pipes); and the temperature of the water.

There are many places in Victoria where you can see these naturally occurring minerals, such as:

- **Peninsula Hot Springs –** the hot springs contain mineral ions from ancient mineral aquifers, such as Mg^{2+}, Cl^{-}, HCO_3^{-} and K^{+}. Visitors can bathe in the geothermal hot springs, which contain high levels of these ions that many believe have healing and rejuvenating properties
- **Hepburn Mineral Springs –** the mineral springs also have high concentrations of mineral ions, such as Na^{+}, K^{+}, Ca^{2+}, Mg^{+2}, Fe^{2+}, Cl^{-} and HCO_3^{-}. People visit the area to bathe in the hot springs, or drink from the old-fashioned pumps that deliver the water from the mineral spring aquifers.

FIGURE 4 The Peninsula Hot Springs in Mornington contain naturally occuring mineral ions.

Heavy metal salts

Heavy metals are usually described as high-density metals that can have a harmful or toxic effect on living organisms.

Australian drinking-water guidelines outline the sources and reasons for some of these heavy metals to be in our water systems, and the levels they are allowed to occur to. You can see these in Table 1.

17.1 Real-world chemistry
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TABLE 1 Sources of heavy metals and allowable levels in water systems

Heavy metal	Source	Allowable levels in drinking water (mg L^{-1})
Mercury (Hg)	• Industry emissions or spills • Used batteries, electrical components, fungicides and preservatives	<0.001
Cadmium (Cd)	• Impurities in the zinc of galvanised pipes or in fittings in water heaters, water coolers and taps • Wastewater, industry and fertilisers	<0.002
Arsenic (As)	• Natural sources – minerals and ores • Run-off from gold mining, residues from burning fossil fuels and industrial waste	<0.01
Lead (Pb)	• Natural sources • Household plumbing sources and industrial waste	<0.01
Chromium (Cr)	• Natural sources such as rocks and soils, • Industrial sources, e.g. production of catalysts, ceramic and glass manufacturing and paints	<0.05
Copper (Cu)	• Natural sources • Copper pipes and fittings, and industry	<1 for aesthetic considerations, but <2 for health considerations
Tin (Sn)	• Coatings in food containers, ceramics and textile industries	No guideline value, as concentrations are likely to be much lower than the level that can cause health effects (<0.000005 in rivers, estuaries, oceans)
Aluminium (Al) (considered a light metal)	• Natural leaching from soil and rock • Aluminium salts in coagulants in water treatment • Domestic and industrial sources	<0.2

Source: *Australian Drinking Water Guidelines* (2011), National Health and Medical Research Council.

Heavy metals do not biodegrade. They stay in the environment and bioaccumulate in the food chain (Figure 5 on the next page). The presence of heavy metals in the food chain can cause health problems, which is why governments monitor and provide guidelines for the levels of these heavy metals in drinking water.

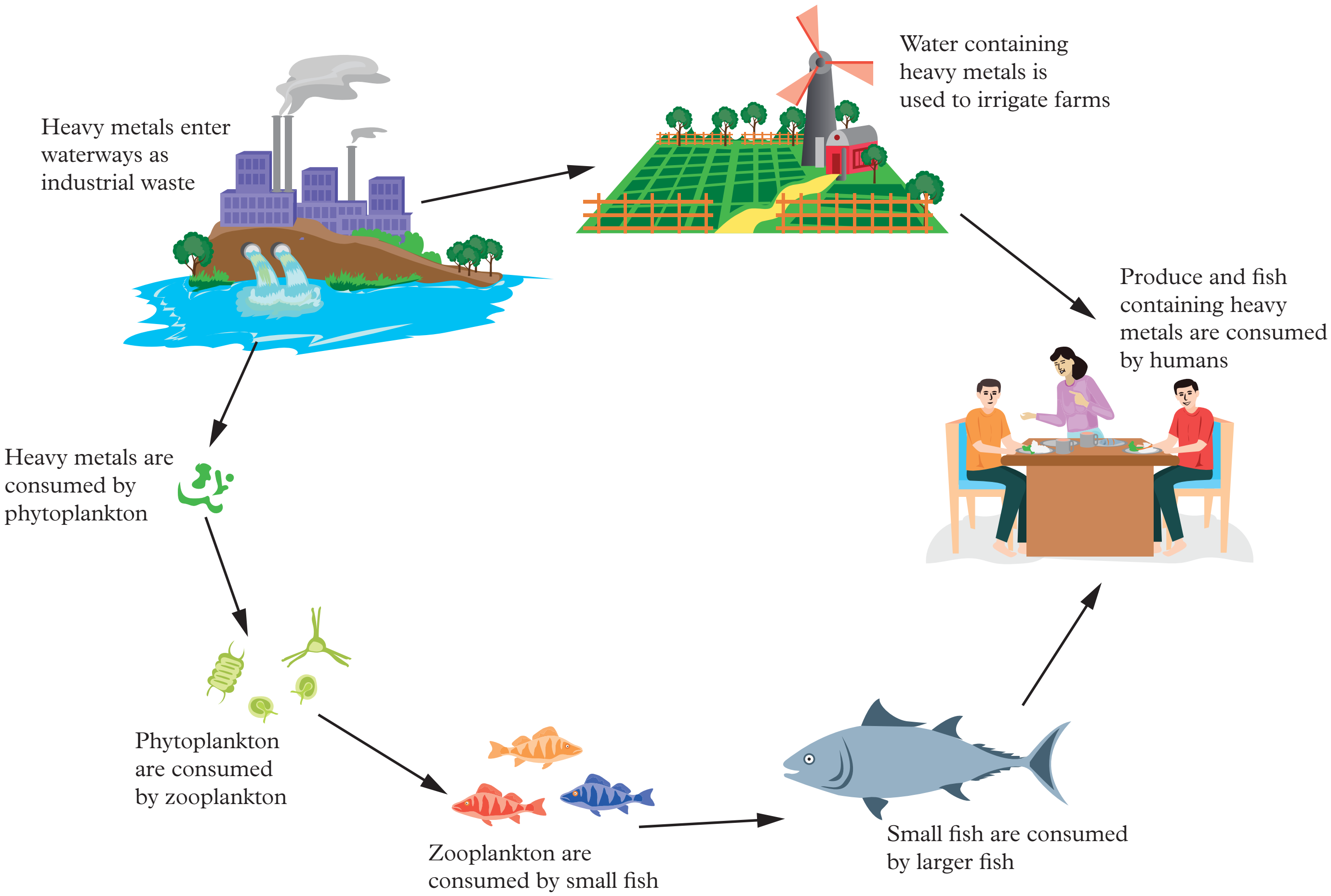

FIGURE 5 Bioaccumulation of heavy metals in the food chain

Organometallic substances

organometallic compound
a chemical compound that has at least one bond between a metal and a carbon that is part of an organic group

Organometallic compounds are chemical compounds in which there is at least one bond between a metal atom and a carbon atom. Organometallic compounds are another source of heavy metals that pollute water systems. Many organometallic compounds are toxic to humans.

Organometallic compounds are used in various industrial processes, including the production of light-emitting diodes (LEDs); margarine; some organic compounds; and as catalysts in some commercial chemical reactions.

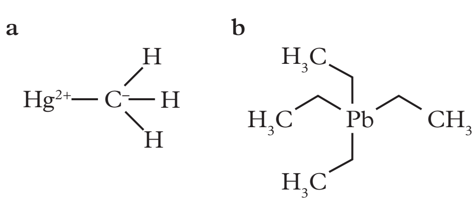

FIGURE 6 **a** Methyl mercury and **b** tetraethyl lead are organometallic compounds.

Examples of organometallic compounds are methyl mercury and tetraethyl lead (Figure 6). Each contain a carbon–metal bond, making them organometallic compounds.

Tetraethyl lead was once added to petrol, but it was banned in Australia because it causes a build-up of lead in roadside soil and houses and accumulates in waterways.

Methyl mercury is formed when compounds containing mercury are combusted. They are more toxic than metal mercury because they are soluble in water and can be transported around the body in the same ways as required nutrients are.

Electrical conductivity to assess salinity and water quality

Salinity levels (concentrations of salts in water and soil) can be both qualitatively and quantitatively analysed. We will look at the quantitative analysis techniques for salts in the next topic. But first we will explore how electrical conductivity can be used to assess the concentration of salts, and water and soil quality.

Electrical conductivity

Electrical conductivity is one of the most common tests for salinity. Pure water does not conduct electricity well. But when salts are dissolved in water, the separated ions carry a negative or positive charge. Salt water is therefore much more conductive than pure water.

There is a direct relationship between electrical conductivity and concentration of salt in the water, which you can see in Figure 7.

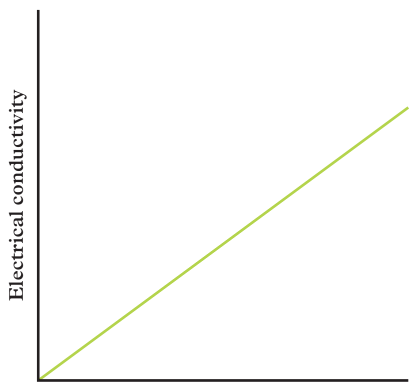

FIGURE 7 The linear relationship between electrical conductivity and salt concentration

As the concentration of salts in the solution increases, the electrical conductivity increases because there are more ions in the solution to carry more electrical charge between the electrodes.

A simple qualitative analysis of salt in water, one which you may have already performed during your study of VCE Chemistry, is shown in Figure 9 on the next page. It is a simple set-up using a power supply attached to a light globe and two electrodes. If the light globe lights up when you put it in the water sample, the sample must contain dissolved salts. Pure water does not conduct electricity well, so will not enable the globe to light up.

FIGURE 8 Salts dissolved in water are able to carry an electric charge.

However, this test doesn't give you a concentration or an understanding of the amount of salt in the samples of water. For this, you can use an electrical conductivity meter, which will give you a reading of the electrical conductivity of the sample in mS cm^{-1} (milliseimens per centimetre) or μS cm^{-1} (microsiemens per centimetre), where:

$$1 \text{ mS} = 1000 \text{ μS}$$
$$1 \text{ μS} = 0.001 \text{ mS}$$

You can assume that a low conductivity means a low salt concentration, and a high electrical conductivity means a high salt concentration in a solution. However, the electrical conductivity doesn't give you the exact concentrations of salts in a sample; nor does it give you the types of salts in a solution. For these, you need quantitative analytical methods, which we will cover in Topic 17.2.

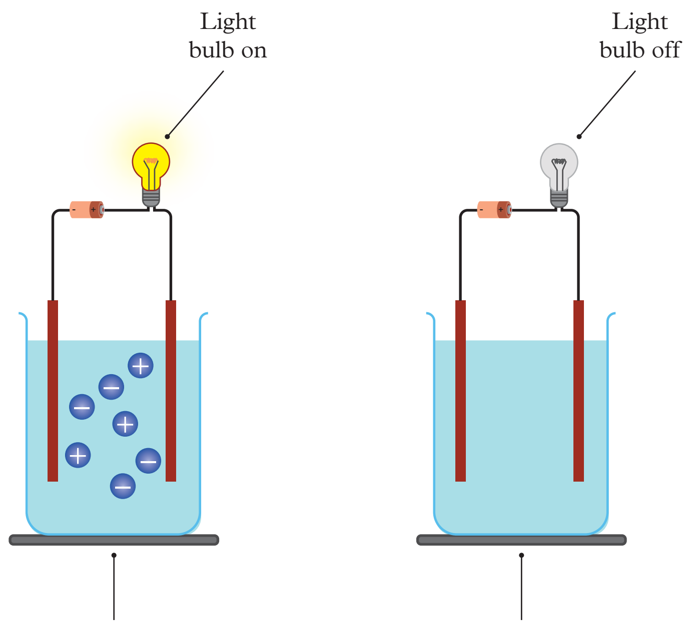

FIGURE 9 Testing the electrical conductivity of water: pure water does not conduct electricity.

Assessing water quality

Low levels of salt ions such as sodium and chloride are important for the functioning of plants and animals. For example, humans need a small amount of sodium and chloride for nerves and muscles to work. If water contains too much salt, it can damage aquatic ecosystems and might not be suitable for human consumption.

Table 2 outlines the levels of salinity in water and how this can affect its suitability for different uses.

TABLE 2 Effect of salinity levels on water use

Salinity ($\mu S\ cm^{-1}$)	Use
0–800	Human consumption, irrigation, livestock
800–2500	Human consumption (although the lower part of the range is better, if available), irrigation (if properly managed), livestock
2500–10 000	Not recommended for human consumption (although water of $\leq 3000\ \mu S\ cm^{-1}$ can be consumed), irrigation of salt-tolerant crops (if $\leq 6000\ \mu S\ cm^{-1}$), some livestock ($\leq 6000\ \mu S\ cm^{-1}$ for poultry and pigs, $\leq 10\ 000\ \mu S\ cm^{-1}$ for all other livestock)
Greater than 10 000	Not suitable for humans, irrigation nor for livestock

Source: Adapted from Murray River Catchment Committee, Water Quality Salinity Standards, 2013

Therefore, testing the electrical conductivity of water is important for understanding water quality in terms of salinity and whether a water system is suitable for its intended purpose.

Drinking water is tested regularly to ensure it meets government standards and is safe for human consumption. Other tests that are performed on drinking water examine:

- clarity or **turbidity**, which is a measure of the cloudiness of water due to suspended solids
- concentration and presence of minerals and ions (e.g. Fe^{2+}, Fe^{3+} and Mn^{2+}) and common ions (e.g. Na^+, Ca^{2+}, K^+, Cl^-, SO_4^{2-} and HCO_3^-) or heavy metals (e.g. cadmium, arsenic and lead)
- pH levels
- concentration of chlorine, chloride ions, fluorine and fluoride ions
- levels of contaminants such as fuels and pesticides
- levels of bacteria.

FIGURE 10 An electrical conductivity meter

17.1 Skill drill
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turbidity
the cloudiness (or opacity) of a liquid due to of suspended solids

17.1 CHECK YOUR LEARNING

Describe and explain

1 Explain the difference between primary and secondary salinity.

2 Describe the ways in which heavy metals enter waterways.

3 Explain why electrical conductivity is considered qualitative analysis rather than quantitative analysis.

Apply, analyse and compare

4 Analyse the heavy metals listed in Table 3 and compare them with the allowable levels in drinking water in Table 1 (page 455).

a Explain the levels of each metal allowed in drinkable water samples.

b Identify where the sample fails or passes for each heavy metal.

c Identify if the water sample would be safe to drink.

TABLE 3 Concentrations of some heavy metals in water samples

	$[Pb^{2+}]$ (mg L^{-1})	$[Cd^{2+}]$ (mg L^{-1})	$[As^{3+}]$ (mg L^{-1})	$[Hg^+]$ (mg L^{-1})
Sample 1	0.0073	0.0016	0.0099	0.0062
Sample 2	0.0085	0.0021	0.006	0.0003
Sample 3	0.0001	0.0004	0.01	0.001

5 Analyse the following ions and classify them as relatively safe or toxic.

Mg^{2+}, Na^+, Hg^+ Cl^-, NO_3^-, Cu^{2+}, K^+, Ca^+, Pb^{2+}

Design and discuss

6 Discuss why replacing native deep-rooted plants with shallow-rooted plants can affect the salinity of groundwater.

7 Discuss why heavy metal levels are monitored in drinking water.

8 Discuss how tests for salinity and quality of water are relevant to the United Nations Sustainable Development Goal 6 ‘Clean water and sanitation’.

17.2 Quantitative analysis of salts

KEY IDEAS

In this topic, you will learn that:

+ gravimetric analysis is a quantitative analysis technique
+ ionic salts are often hydrated, and you can determine their molar ratio
+ mass–mass stoichiometry can be used to find the mass of unknown ionic compounds
+ a concentration curve can be plotted from the analysis of ionic compounds using colorimetry and UV–visible spectroscopy.

In this topic, we will explore techniques that can be used to qualitatively analyse the amount and types of salts found in water and soil samples, including:

- gravimetric analysis to determine the:
  - molar ratio of water hydration for an ionic compound
  - mass of an ionic compound using mass–mass stoichiometry
- colorimetry to determine the concentration of ionic compounds or complexes
- UV spectroscopy to determine the concentration of ionic compounds or complexes.

Gravimetric analysis

gravimetric analysis the quantitative determination of the mass of a substance

Gravimetric analysis is a set of methods for determining an unknown mass of a substance in a sample. In this topic, we will explore two of these in methods:

- heating a hydrated sample to remove water and calculating the molar ratio of water in the sample
- precipitating a component of the sample into a solid form and determining the composition of a sample.

Molar ratio of water hydration for an ionic compound

Salts can exist as hydrates, meaning there are some water molecules incorporated into the lattice structure. For example, $CaCl_2 \cdot 6H_2O$ is called calcium chloride hexahydrate. Every mole of the ionic salt contains 6 moles of water within the lattice structure.

Once the water is removed from the salt, by using the technique outlined in Figure 1, an **anhydrous salt** is produced.

anhydrous salt an ionic compound from which the water has been removed

Heating a sample to remove water

Heating a sample to remove water is a technique used to determine the molar ratio of water hydration for an ionic compound.

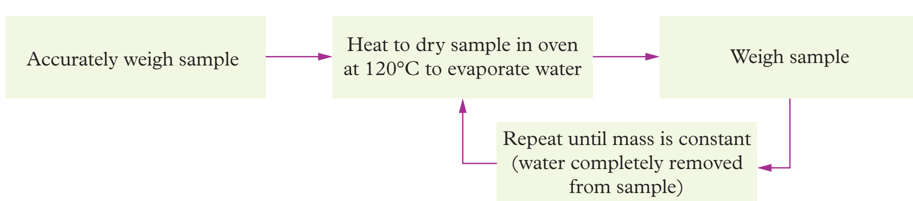


FIGURE 1 The technique used to remove water from a sample to produce an anhydrous salt

hydrated salt an ionic compound that contains water molecules in a set mole ratio within the ionic lattice

Once you have a completely dried sample, you can calculate the molar ratio of the **hydrated salt**.

17.2A PRACTICAL: CONTROLLED EXPERIEMENT
How is gravimetric analysis used to identify the mass of salt in a solution? Go to page **524**.

17.2B PRACTICAL: CASE STUDY
What can we do if soil salinity is too high?. Go to page **527**.

Calculating the molar ratio of a hydrated salt

To calculate the molar ratio of a hydrated salt:

1 Use the formula:
mass of water (in hydrated salt) = hydrated salt – anhydrous salt
(change in mass) = (initial mass) – (final mass)

2 Calculate the number of moles of each. Divide each substance, the mass of water and the anhydrous salt by their molar mass (M) to find the number of moles (n) of each.

3 Calculate the molar ratio. Divide each mole by the smallest of the two numbers and you have the molar ratio of the hydrated salt.
You can see this set of calculations used in Worked example 17.2A.

17.2A WORKED EXAMPLE

DETERMINING THE MOLAR RATIO OF A HYDRATED SALT

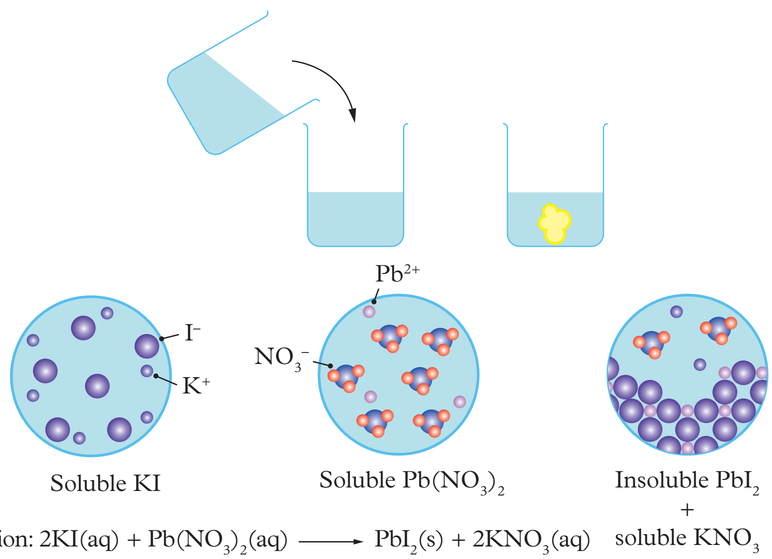
You weigh out a sample of 15.67 g of hydrated magnesium carbonate, $MgCO_3{\bullet}xH_2O$. After heating the sample to a constant mass, the anhydrous salt that remains is 7.58 g. Calculate the formula of the anhydrous salt.

Solution

Think	Do
Step 1: Substitute the values provided in the question into the formula: mass of water (in hydrated salt) = hydrated salt – anhydrous salt	mass of water (in hydrated salt) = 15.67 – 7.58 = 8.09 g
Step 2: Calculate the number of moles of each. Divide each substance, the mass of water and the anhydrous salt by their molar mass (M) to find the mole (n) of each.	Anhydrous salt / Water mass (m): 7.58 / 8.09 molar mass (M): 84.3 / 18.0 n (mol): 0.09 / 0.45
Step 3: Then calculate the molar ratio. Divide each mole by the smallest of the two numbers.	Molar ratio = 1:5, meaning for every 1 magnesium carbonate, there are 5 water molecules. The formula of the hydrated salt is $MgCO_3{\bullet}5H_2O$

Mass–mass stoichiometry to determine an ionic compound

In Chapter 5, you learnt that precipitation reactions can be used to form a precipitate (an insoluble salt) by adding together aqueous salt solutions. The reaction and equation are summarised in Figure 2.



Full equation: $2KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$

Ionic equation: $Pb^{2+}(aq) + 2I^-(eq) \longrightarrow PbI_2(s)$

FIGURE 2 Precipitation reaction and equation

Using a precipitation reaction to determine mass of an ionic compound

By forming a precipitate, you can conduct gravimetric analysis, as shown in Figure 3. Then, you can use stoichiometry to calculate the mass of the unknown ionic compound in the sample.

TABLE 1 Common precipitates formed for use in gravimetric analysis

Precipitate	Used for
$AgCl$	Analysis of chloride ions in foods
$Mg_2P_2O_7$	Analysis of phosphate in water samples
Fe_2O_3	Analysis of iron in ores
$BaSO_4$	Analysis of sulfur in fertilisers

The precipitate must be stable when heated, so some precipitates are not suitable for gravimetric analysis.

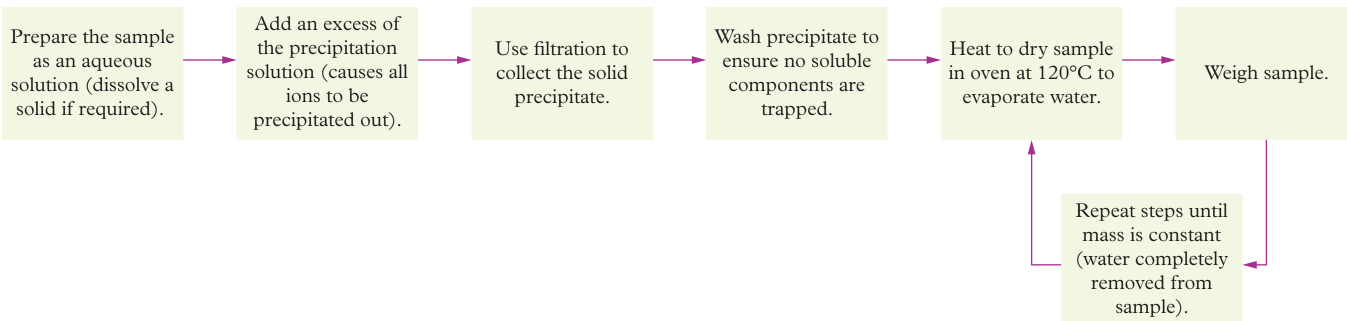


FIGURE 3 The gravimetric analysis technique used to produce a precipitate to identify the mass of an unknown ionic compound

Sometimes, the filtrate can be reacted with the precipitation solution again and the process repeated to ensure all the unknown compound is reacted. Once you know the mass of the sample, you can use mass–mass stoichiometry to determine the mass of the salt in the sample.

Mass-mass stoichiometry to determine the mass (concentration) of ionic compounds

In previous chapters, you have used stoichiometry to calculate the number of moles of substances using the mole ratio. Mass–mass stoichiometry works on the same principle but uses $n = \frac{m}{M}$ to calculate the number of moles of one substance from the original mass, and the mole ratio to calculate the mass of the unknown substance. The relationship for mass–mass stoichiometry is shown in Figure 4.

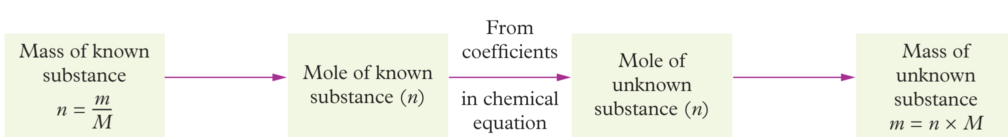


FIGURE 4 The process of mass–mass stoichiometry

In a balanced chemical equation, the coefficients give the ratio in which the reactants will react with each other. For example:

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

According to the balanced equation, 2 mol of potassium iodide (KI) will react with 1 mol of lead nitrate ($Pb(NO_3)_2$) to produce 1 mol of the precipitate, lead iodide (PbI_2) and 2 mol of potassium nitrate (KNO_3).

Using the ratios from a balanced chemical equation, you can then calculate an unknown mass by transposing the equation

$$n = \frac{m}{M}$$

where n is amount (mol), m is mass (g) and M is molar mass (g mol^{-1}).

There are four main steps to performing calculations for mass–mass stoichiometry:

1. Write a balanced chemical equation, define the variables and identify the known substance.
2. Calculate the amount, in mol, of the substance with the known concentration and volume using $n = \frac{m}{M}$.
3. Use the mole ratio for the equation to calculate the unknown mol of the other reactant; the mole ratio is:
$$\text{mole ratio} = \frac{\text{coefficientof unknown}}{\text{known coefficient}}$$
4. Calculate the unknown mass using $n = \frac{m}{M}$.

Worked example 17.2B shows these steps being used to calculate known and unknown number of moles and masses using stoichiometry.

17.2B Worked example
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17.2B Worked example
Video demonstration

17.2C WORKED EXAMPLE

MASS–MASS STOICHIOMETRY FOR CALCULATING THE PERCENTAGE OF LEAD IN A SOIL SAMPLE

A precipitation reaction was undertaken to find the mass of lead ions (Pb^{2+}) in a 400.00 g soil sample. The soil was dissolved in water, the aqueous solution was precipitated using excess aluminium sulfate ($Al_2(SO_4)_3$) and was then filtered. The final mass of the precipitate ($PbSO_4$) was 17.41 g.

a Calculate the mass of the lead ions in the soil sample.

b Calculate the % mass of lead ions in the soil sample.

Think	Do
Step 1: Write the balanced equation for the reaction, define the variables and identify the known substance.	Known substance ↓ $Al_2(SO_4)_3(aq) + 3Pb^{2+}(aq) \rightarrow 3PbSO_4(s) + 2Al^{3+}(aq)$ Pb^{2+}: m = ? ; M = 207.2 g mol^{-1} $PbSO_4$: **Precipitate**; m = 17.41 g; M = 303.3 g mol^{-1}

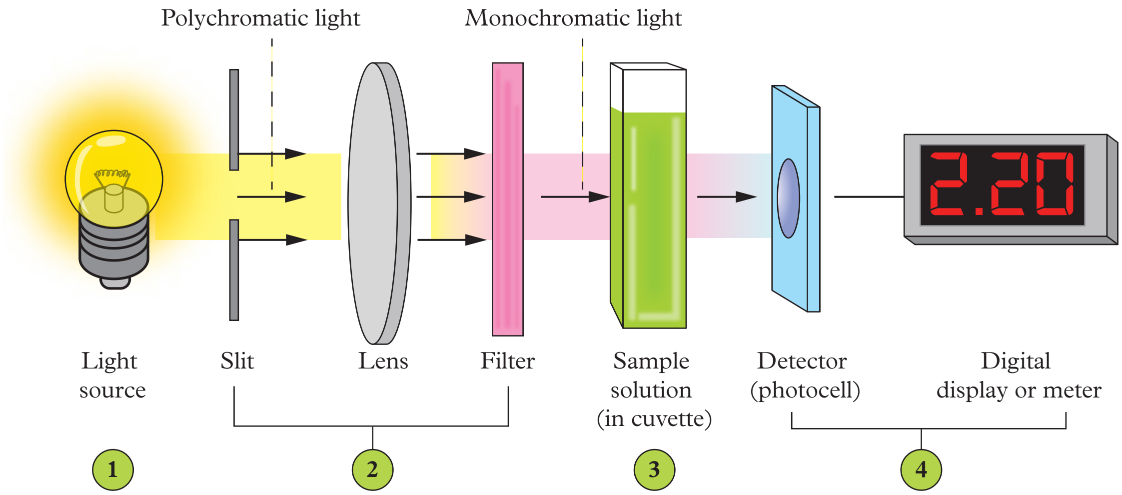
Think	Do
Step 2: Calculate the amount, in mol, of the substance with the known mass.	We have identified the known substance as $PbSO_4$, so calculate the number of mol of $PbSO_4$ by using $n = \frac{m}{M}$. $m(PbSO_4) = 17.41$ g $M(PbSO_4) = 207.2 + 32.1 = (16 \times 4) = 303.3$ g mol^{-1} $n(PbSO_4) = \frac{m}{M}$ $= \frac{17.41}{303.3}$ $= 0.0574$ mol
Step 3: Use the mole ratio for the equation to calculate the unknown number of mol of the other reactant.	Mole ratio $= \frac{\text{unknown coefficient}}{\text{known coefficient}}$ $\frac{n(Pb^{2+})}{n(PbSO_4)} = \frac{3}{3}$ To calculate the unknown number of mol of Pb^{2+}, multiply the known number of mol of $PbSO_4$ by the mole ratio. $n(Pb^{2+}) = n(PbSO_4) \times$ mol ratio $= 0.0574 \times \frac{3}{3}$ $= 0.0574$ mol
Step 4: Calculate the unknown mass using $m = n \times M$.	**a** $n(Pb^{2+}) = 0.0574$ mol $M(Pb^{2+}) = 207.2$ g mol^{-1} $m(Pb^{2+}) = n \times M$ $= 0.0574 \times 207.2$ $= 11.89$ g (4 sig fig)
Step 5: Calculate the % mass of the lead ions in the soil sample.	**b** $Pb^{2+} = \frac{\text{mass of } Pb^{2+}}{\text{mass of original soil sample}} \times \frac{100}{1}$ $= \frac{11.89}{400.00} \times \frac{100}{1}$ $= 2.973\%$ (4 sig fig)

Colorimetry

colorimetry
a method for determining the concentration of a coloured compound in solution by measuring absorption of a particular wavelength of light

Colorimetry is used to determine the concentration of coloured compounds in a solution. It is a device that allows specific solutions to absorb a particular wavelength of light. The process is summarised in Figure 6.

FIGURE 5 Cuvettes are used in colorimetry.



1 A light source produces a light ray.

2 The light ray passes through a series of lenses and filters, which splits the light ray, allowing only a specific wavelength to pass through.

3 The specific light ray reaches the cuvette, where it is absorbed by the solution.

4 The transmitted rays fall on the photocell, which measures the transmitted light and converts it into electrical signals that are measured and displayed.

FIGURE 6 The basic structure of a colorimeter

A colorimeter uses a specifically selected colour of light to identify the intensity of coloured compounds in a solution. Sometimes, specific compounds are added to a sample to produce a coloured compound so that it can be analysed by colorimetry.

The wavelength is selected to ensure that the sample absorbs as much of the coloured light as possible, so a complementary coloured filter needs to be selected. For example, copper sulfate transmits blue light, and it absorbs red-orange light, so a red or orange light filter of around 580–680 nm needs to be used. You can see in Table 2 the colours of visible light and their complementary colours, or the colour of filter that needs to be used in colorimetry.

TABLE 2 The colours of visible light and their complementary colours, or the filter colour to be used

Colour observed/colour of sample	Colour absorbed (colour of filter)	Wavelength of filter colour/ colour absorbed (nm)
Green-yellow	Violet	380-420
Yellow	Violet-blue	420–440
Orange	Blue	440–470
Red	Blue-green	470–500
Purple	Green	500–520
Violet	Green-yellow	520–550
Violet-blue	Yellow	550–580
Blue	Orange	580–620
Blue-green	Red	620–680
Green	Purple	680–780

When using colorimetry to analyse an unknown concentration of sample, you must also use another solution with known concentrations of the same ions. This way you can construct a **calibration curve** to determine the unknown concentration. We will look at developing and using calibration curves later in this chapter.

calibration curve a graph of absorbance versus concentration of standard solutions so that an unknown concentration can be determined

UV–visible spectroscopy

UV–visible spectroscopy is an instrumental technique that uses light from the ultraviolet (UV) and visible light regions of the electromagnetic spectrum, as shown in Figure 7.

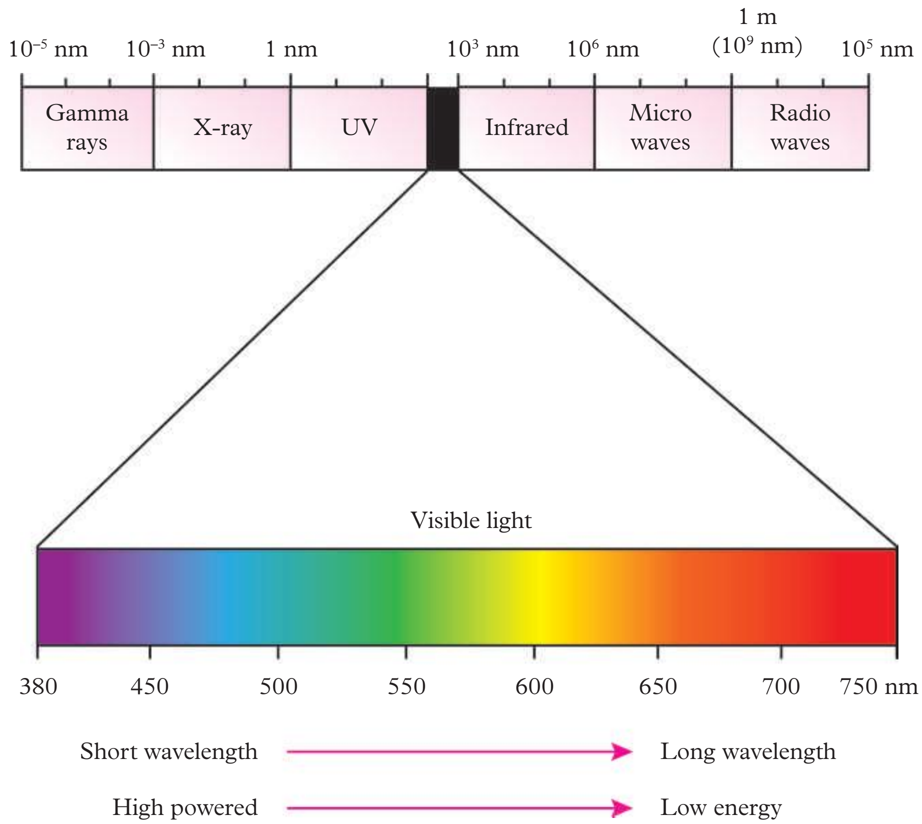


FIGURE 7 The electromagnetic spectrum

A UV–visible spectrophotometer is similar to a colorimeter, but it uses a monochromator rather than a filter (Figure 8). This monochromator selects light that has the correct wavelength to be used in an analysis.

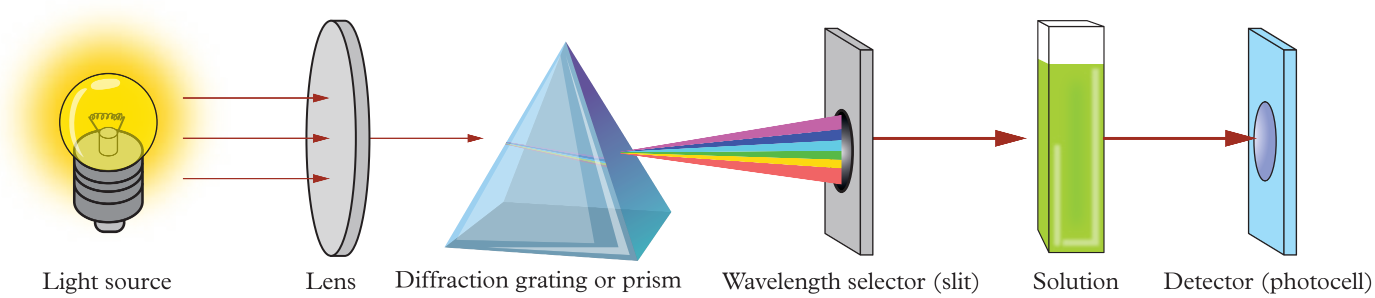


FIGURE 8 The processes in a UV–visible spectrophotometer

UV–visible spectrometry uses the same principles as colorimetry, although it is far more selective because it analyses the full range of UV and visible light wavelengths and selects the best one to be used. It is also less prone to interference by compounds with similar colours. As you can see in its UV absorption spectrum (Figure 9), chlorophyll absorbs mostly blue (450 nm) and red (650 nm). In the absence of blue and red, the transmitted radiation appears green.

The measurement of the **absorbance** of the samples and standard solutions will then take place at one of these two high peaks – blue or red.

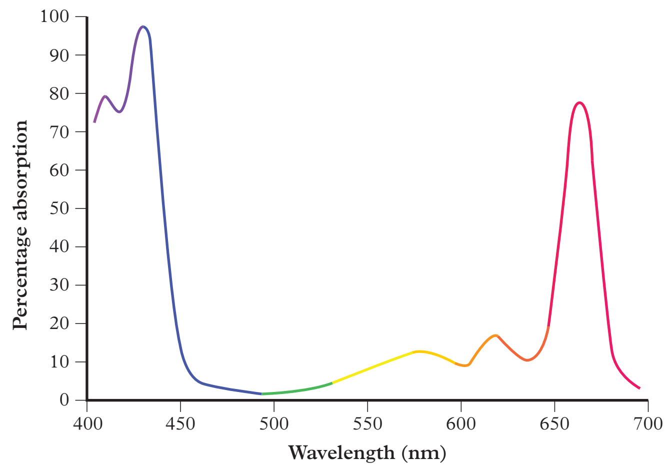


FIGURE 9 The absorption spectrum of chlorophyll

Complexes

Sometimes a metal cation won't have a strong colour and cannot be analysed by colorimetry or UV–visible spectroscopy. In these cases, an oxidant can be added to oxidise the solution and change the colour of the solution. For example, Fe^{3+} is normally a clear solution. It can be oxidised using acidic potassium permanganate solution, and it forms the complex $FeSCN^{2+}$, which has a vibrant red colour. This coloured metal complex can then be analysed using colorimetry or UV–visible spectroscopy.

absorbance
a measure of the amount of light that is absorbed by a sample during colorimetry of UV–visible spectroscopy

Using a calibration curve to determine unknown concentrations in water and soil samples

As discussed earlier in the chapter, the concentration of ions and complexes can be determined by plotting a calibration curve. To do this, you need to analyse known concentrations of a standard solution along with the unknown sample. Standard solutions contain the same ions under the same conditions. This can be done by either colorimetry, for simple coloured samples, or UV–visible spectroscopy for more complex compounds.

The concentration versus absorbance of the unknown and standard solutions are plotted on a graph. Let's examine the absorbance from the colorimetry of copper sulfate, in different blue solutions, as shown in Figure 10.

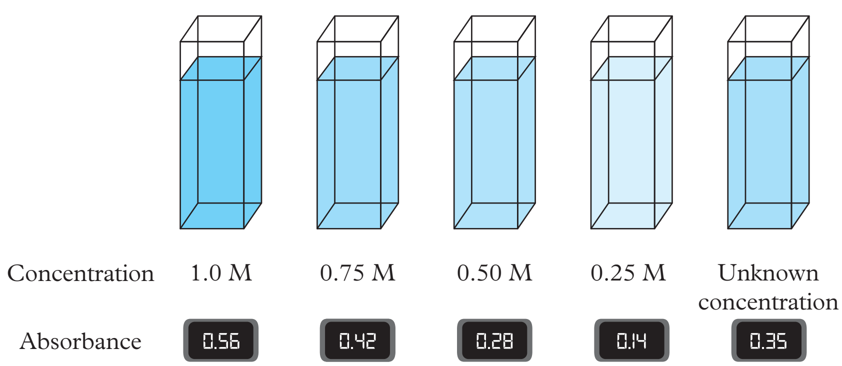


FIGURE 10 Concentrations and absorbances from colorimetry of copper sulfate solutions

TABLE 3 Known concentrations and measured absorbances of copper sulfate solutions

Concentration (M)	Absorbance
0.25	0.14
0.50	0.28
0.75	0.42
1.0	0.56
Unknown	0.35

- **Step 1:** Run known concentrations of copper sulfate solutions through the colorimeter and measure the absorbances. You can see these in Figure 10 and Table 3.
- **Step 2:** Plot the concentration (*x*-axis) versus absorbance (*y*-axis) to create a calibration curve using a line of best fit (see Figure 11).

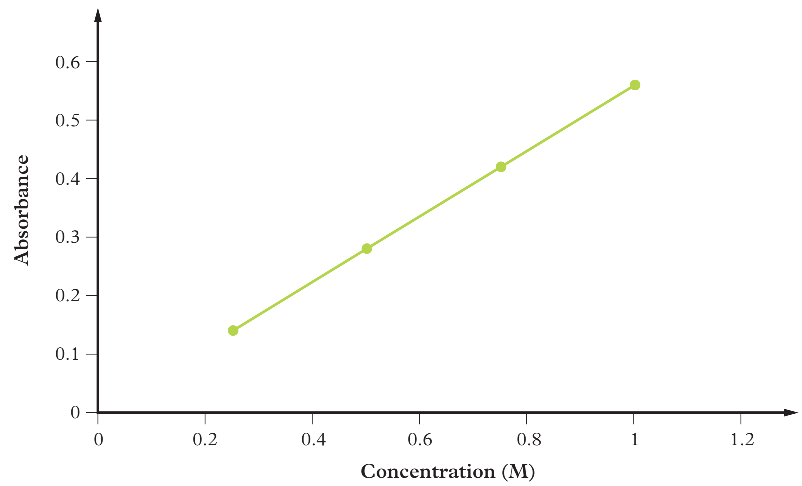


FIGURE 11 A calibration curve of known concentrations of copper sulfate and absorbance from colorimetry

- **Step 3:** Find the absorbance of the unknown copper solution measured in the colorimeter on the calibration curve, by following across from the absorbance on the *y*-axis, and then follow down to the concentration on the *x*-axis and determine the concentration in M, as shown in Figure 12.

Study tip

To ensure accuracy when creating a calibration curve, you need to have an idea of the concentration of your unknown sample, so that you can have two data points above and below, if possible. If you are significantly outside of your range, you won't be able to determine the concentration.

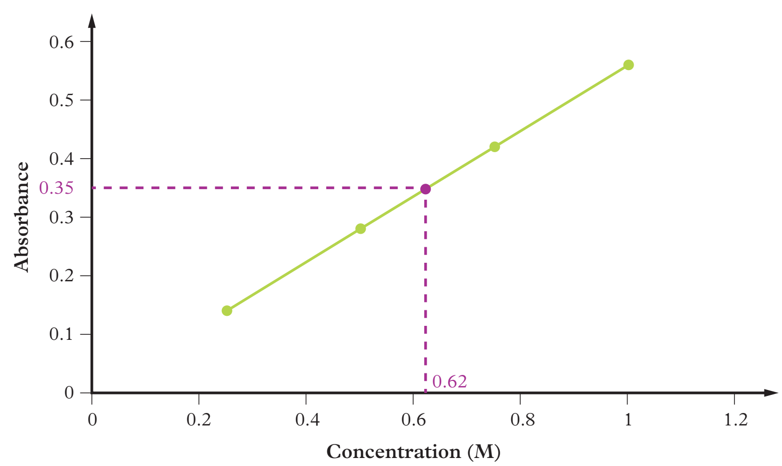


FIGURE 12 The concentration of a copper sulfate solution can be determined from the calibration curve of known concentrations.

17.2 Challenge
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From the calibration curve, we can interpret that the concentration of copper sulfate in the sample is 0.62 M.

The process of determining the concentration of ions or complexes is the same for UV–visible spectroscopy as shown here for colorimetry.

17.2 CHECK YOUR LEARNING

Describe and explain

1 What is a hydrated salt?

2 Describe the process of gravimetric analysis and explain how it differs for calculating the molar ratio of hydration of a salt and for determining mass of a precipitate.

3 Explain the function of a calibration curve.

Apply, analyse and compare

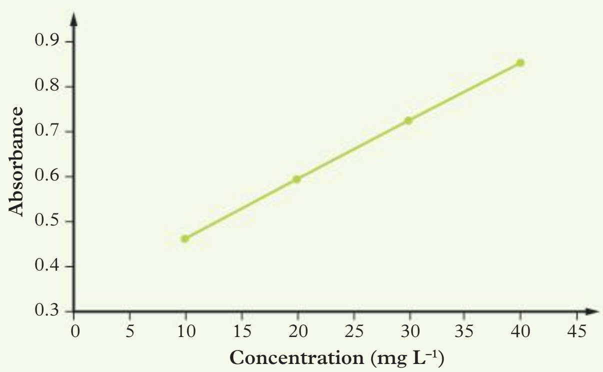
4 A 0.470 g sample of hydrated nickel(II) chloride is heated and weighed until a constant mass is found. The final mass is 0.256 g.

- **a** Calculate the molar mass of the hydrated nickel(II) chloride.
- **b** Determine the formula of the hydrated nickel(II) chloride.

5 Compare colorimetry and UV–visible spectroscopy.

6 When solutions of sodium chloride and lead(II) nitrate are added together, a precipitate of lead(II) chloride is formed. Calculate the mass of sodium chloride that would be required to form 3.62 g of the precipitate.

7 Water from a household water tank was analysed for Cr^{3+} ions by UV–visible spectroscopy. A calibration curve was created from four known concentrations of chromium ions, as shown.



A 20 mL sample was diluted to 100 mL and the absorbance was measured as 0.66.

- **a** What is the concentration of Cr^{3+} ions in the diluted sample?
- **b** Calculate the concentration of Cr^{3+} ions in the water tank.
- **c** Identify if this water is safe to consume with this level of chromium in it.

Chromium ions can be oxidised to form potassium dichromate, an orange solution.

- **d** Describe the process you could use to analyse potassium dichromate by colorimetry.

8 Compare the sample preparation techniques you would use for analysing a sample for chromium ion concentration by UV–visible spectroscopy and colorimetry.

9 Magnesium sulfate is a hydrate salt with the formula $MgSO_4{\cdot}6H_2O$. A measured amount of the salt was heated and the water evaporated. The final mass of the dried anhydrous salt was 13.2 g. Calculate the:

- **a** mass of the water that would have been evaporated off
- **b** total mass of the hydrated salt before evaporation occurred.

10 An unknown hydrated salt has the hydration ratio $X_3(PO_4)_2{\cdot}4H_2O$.
You start with 7.739 g of the hydrated salt and evaporate all the water off to achieve a final mass of anhydrous salt of 6.28 g.

- **a** Calculate the mass of water evaporated off.
- **b** Calculate the molar mass of the salt.
- **c** Determine the metal ion X that is in the salt.

Design and discuss

11 Students wanted to analyse the mass of chloride ions in a 100 mL tap water sample. They added excess silver nitrate solution to the sample, then filtered, dried and weighed the precipitate that was formed until a constant mass of 4.56 g was achieved.

- **a** Write a balanced ionic equation for the precipitation reaction.
- **b** Calculate the mass of the chloride ions in the tap water.
- **c** Calculate the percentage of ions in the 100 mL tap water.
- **d** Discuss the effect on the results if the students didn't correctly dry the precipitate.

CHAPTER

17 Review

Chapter summary

17.1

- Primary salinity is the natural occurrence of salts in the environment.
- Secondary salinity is the addition of salts to the environment through human activity.
- Sources of secondary salinity can be dryland salinity or irrigation-induced salinity and human activity such as mining, agriculture and domestic sources.
- There are different types of salts – they include mineral salts, heavy metal salts, organometallic substances.
- Electrical conductivity can be used to qualitatively analyse water and soil samples for the presence of salts.
- As the concentration of salt in a sample increases, electrical conductivity also increases.

17.2

- Gravimetric analysis can be used to determine an unknown mass of a substance in a sample by two techniques:
  - heating a hydrated sample to remove water and calculate the molar ratio of water in the sample
  - precipitating a component of the sample into a solid form and determining the composition of a sample by mass–mass stoichiometry to calculate the original sample mass.
- Colorimetry is used to determine the concentration of coloured compounds in a solution.
- UV–visible spectroscopy can be used to determine concentrations of ions in solutions and is more accurate than colorimetry.
- A calibration curve is prepared by measuring the absorbance of a series of standard solutions with known concentrations, with the absorbance plotted against the concentration.
- The concentration of a sample can then be determined by plotting its absorbance on the calibration curve and reading the concentration.

Key formulas

Amount, in mol, of a substance	$n = \frac{m}{M}$

Chapter checklist

Use the success criteria in the table below to rate how well you understand each concept as ‘Confidently’, ‘Mostly’ or ‘Not really’. If you’re not feeling confident about any of these skills or ideas, use the revision links to revisit them.

I can ...	Confidently	Mostly	Not really	Revision link
• describe the sources of salts found in water (including minerals, heavy metals and organometallic substances)	☐	☐	☐	Go back to Topic 17.1
• explain how electrical conductivity to assess the salinity and quality of water or soil samples	☐	☐	☐	Go back to Topic 17.1
• use mass–mass stoichiometry to determine the mass present of an ionic compound	☐	☐	☐	Go back to Topic 17.2
• use calibration curves to determine the concentration of ions or complexes in water or soil samples	☐	☐	☐	Go back to Topic 17.2

Revision questions

Multiple choice

1 Identify the list that contains heavy metals that are all found in water and soil.

A Hg, Au, As, Ag
B Cu, Cr, As, Sn
C Pb, Mn, Cr, Sn
D Pb, H, Ca, Hg

2 Identify the incorrect statement. Dryland salinity occurs when:

A excess water is added to crops and enters groundwater.
B deep-rooted plants are removed.
C shallow-rooted plants replace deep-rooted plants.
D less water is required by plants, and excess water enters groundwater.

3 A water sample was tested and had an electrical conductivity of 6500 $\mu S\ cm^{-1}$. Identify whether it is suitable for:

A humans.
B cows.
C pigs.
D poultry.

4 A hydrated salt of mass 21.90 g was heated and dried until a final mass of 12.67 g was maintained. Identify the salt.

A $CaCO_3{\bullet}6H_2O$
B $Na_2CO_3{\bullet}6H_2O$
C $Mg(NO_3)_2{\bullet}6H_2O$
D $CaSO_4{\bullet}6H_2O$

5 Identify the correct solution to use to precipitate chloride ions from a solution.

A $AgCl$
B KNO_3
C $AgNO_3$
D KCl

6 If you were analysing a blue-green sample by colorimetry, identify the filter colour you would use.

A Violet
B Yellow
C Blue-green
D Red

7 Identify which is not a source of heavy metal pollution in soil and water.

A Potassium from the breakdown of organic waste

B Arsenic from mining run-off

C Cadmium from batteries

D Lead from paints

8 A calibration curve for Fe^{2+} is shown in Figure 1.

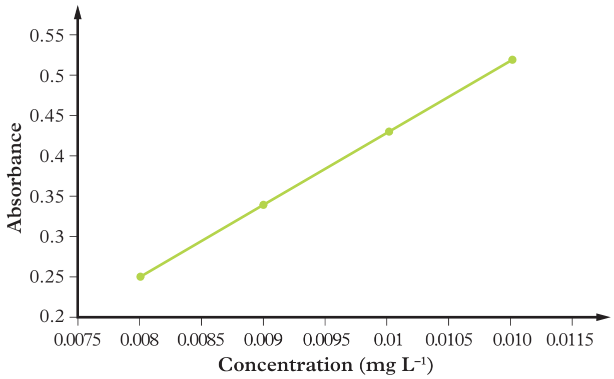


FIGURE 1 A calibration curve for Fe^{2+}

If the absorbance of a water sample was 0.46, identify the correct concentration of Fe^{2+} ions in the water sample.

A 0.0096 mg L^{-1}

B 0.010 mg L^{-1}

C 0.0103 mg L^{-1}

D 0.011 mg L^{-1}

The following information relates to Questions 9 and 10.

A gravimetric analysis was completed to find the concentration of copper(II) sulfate in a water sample. The sample was precipitated with sodium hydroxide and a mass of 14.36 g of the precipitate was formed.

9 Identify the correct balanced precipitation reaction.

A $Cu_2SO_4(aq) + 2NaOH(aq) \rightarrow 2CuOH(s) + Na_2SO_4(aq)$

B $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(aq) + Na_2SO_4(s)$

C $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$

D $CuSO_4(aq) + NaOH(aq) \rightarrow CuOH(s) + NaSO_4(aq)$

10 Identify the mass of copper sulfate in the water sample.

A 8.10 g

B 11.70 g

C 16.13 g

D 23.50 g

Short answer

Describe and explain

11 Explain the reason for using electrical conductivity testing for salts in water and soil.

12 Explain the relationship between concentration of salt in a sample and electrical conductivity.

13 You have a water sample you suspect to be just within the acceptable drinking level range of mercury ions, <0.001 mg L^{-1} Identify the four concentrations of Hg^{2+} ions you would use if you were to analyse the sample by UV–visible spectroscopy.

14 Explain why it is important to test drinking water to see if it complies with a set of standards before it is made available to consumers.

FIGURE 2 A scientist testing drinking water

15 Explain why in colorimetry you use light of a complementary colour to analyse a sample.

16 Describe the steps you would need to take to analyse a sample of hydrated nickel(II) chloride to be able to determine its formula.

17 When using UV–visible spectroscopy, explain if you could use a pre-existing calibration curve, done on a different machine to analyse an unknown sample's concentration.

18 Explain why it wouldn't be ideal to analyse samples of water and/or soil for concentrations of potassium ions or sodium ions using gravimetric analysis.

Apply, analyse and compare

19 An 8.62 g sample of hydrated sodium carbonate is heated and weighed until a constant mass is found. The final mass is 6.44 g.

  a Calculate the molar ratio of the hydrated sodium carbonate.

  b Determine the formula of the hydrated sodium carbonate.

20 A 37.45 g sample of hydrated magnesium carbonate is heated and weighed until a constant mass is found. The final mass of the anhydrous salt is 20.2 g.

  a Calculate the molar ratio of the hydrated magnesium carbonate.

  b Determine the formula of the hydrated magnesium carbonate.

21 Calcium carbonate hexahydrate is a hydrated salt of formula $CaCO_3•6H_2O$. A measured amount of the salt was heated and the water evaporated. If the final mass of the dried anhydrous salt was 8.37 g, calculate the:

  a mass of the water that would have been evaporated off

  b total mass of the hydrated salt before evaporation occurred.

22 An unknown hydrated salt has the hydration ratio $X_2SO_4•5H_2O$.
You start with 37.83 g of the hydrated salt and evaporate all the water off to achieve a final mass of anhydrous salt of 23.16 g.

  a Calculate the mass of water evaporated off.

  b The amount of anhydrous salt in the sample is 0.163 mol. Calculate the molar mass of the salt.

  c Determine the metal ion X that is in the salt.

23 If 14.5 g of $MgSO_4•7H_2O$ is thoroughly heated, what mass of anhydrous magnesium sulfate will remain?

24 When solutions of silver nitrate and barium chloride are added together, a precipitate is formed.

  a Write the balanced precipitation equation for this reaction.

  b Identify the compound that will be the precipitate.

  c Calculate the mass required to form 1.90 g of the precipitate.

25 A soil sample near a major road was tested for lead ions (Pb^{2+}) by UV–visible spectroscopy. The sample had an absorbance of 0.39. Four known concentrations of Pb^{2+} ions were analysed to create the calibration curve shown in Figure 3.

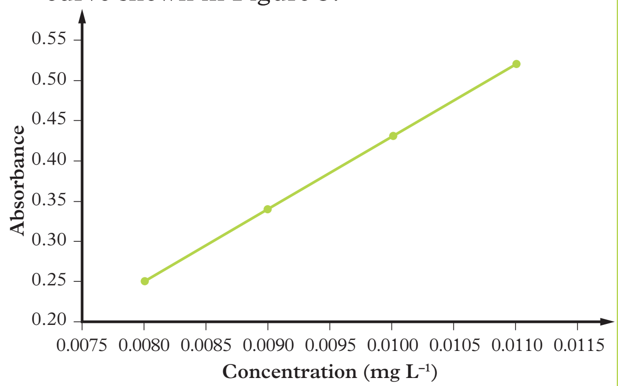


FIGURE 3 A calibration curve for Pb^{2+} ions

  a Determine the concentration of Pb^{2+} ions in the soil sample

  b Propose a method for determining the concentration of Pb^{2+} ions by gravimetric analysis instead of UV–visible spectroscopy.

26 The concentration of Cu^{2+} ions in a sample of dam water was analysed by colorimetry. The standard solutions and samples' absorbances are shown in Table 1.

TABLE 1 The concentration of Cu^{2+} ions in a sample of dam water

Concentration (mg L^{-1})	Absorbance
0.2	0.34
0.4	0.50
0.6	0.66
0.8	0.82
Sample	0.57

a Draw a calibration curve by using the data from the standard solutions.

b Use the calibration curve to determine the concentration Cu^{2+} in the water sample.

Design and discuss

27 Some students decided to analyse the difference in copper in tap water between old houses and new houses in the same street.

FIGURE 4 Taps from **a** an older house and **b** a new house

Their hypothesis was 'If the pipes are older, then there will be more copper present in the water sample because old pipes are more likely to transfer copper into the water passing through them'.

They researched and found that copper ions mainly exist as copper carbonate in water, so they decided to perform precipitation reactions with the same volume of each sample and added excess sodium hydroxide solution to each of the four samples.

They filtered off the precipitates and dried and weighed each of the samples by the same method. The results are in Table 2.

TABLE 2 Student results

Sample	Mass of precipitates (g)
New house 1	3.89
New house 2	2.67
Old house 1	4.32
Old house 2	3.29

a Write a balanced ionic equation for the precipitation reactions that the students performed.

b Calculate the mass of the copper carbonate in each of the samples.

c Analyse the students' hypothesis and explain if they were correct or not.

d Discuss if the students' calculating the mass of copper carbonate in the water is equivalent to copper ions in the water.

e If the students didn't correctly dry the precipitate, discuss how it would affect their results.

28 A 20 mL water sample needs to be analysed for the concentration chloride ions.

a Design a method to analyse this sample, by gravimetric analysis, in which you would end up with a precipitate of AgCl.

b Write a balanced equation for the precipitation reaction.

c Calculate the concentration of chloride ions, in mg L^{-1}, if the final mass of precipitate is 1.06 g.

d If you had five different samples to analyse for chloride ion concentration, discuss another technique you could use to analyse multiple samples.

29 The concentrations of cadmium ions (Cd^{2+}) were analysed in two water samples by UV–visible spectroscopy. A set of four known concentrations was analysed along with two water samples. The data is in Table 3.

TABLE 3 Cd^{2+} concentrations in two different samples

Concentration (mg L^{-1})	Absorbance
0.001	0.33
0.0015	0.51
0.002	0.69
0.0025	0.87
Sample 1	0.53
Sample 2	0.94

a Draw a calibration curve for the known concentrations.

b Determine the concentrations of cadmium ions of the two water samples.

c Explain whether the samples are safe to drink with this concentration of cadmium ions in them.

d Discuss the effectiveness of using this set of known concentrations for both samples.

e Discuss what you would change if you were to complete this analysis again.

You can find the following resources for this section in your obook pro:

Quizlet
Compete in teams or against yourself to test your knowledge.

Chapter quiz
Test your understanding of Key Knowledge in this chapter.

UNIT 2

AREA OF STUDY 2 Checkpoint

Multiple choice

Question 1

Gases in the atmosphere include:

I methane

II water vapour

III nitrogen

IV carbon dioxide

Which of these can cause enhanced greenhouse effects?

A I and II only

B II, III and IV only

C I, II and IV only

D I, II and III only

Question 2

The mass of $CuSO_4{\bullet}5H_2O$ needed to make 1.5 L of a 0.08 M solution is closest to:

A 3.0 g

B 5.0 g

C 30 g

D 50 g

Question 3

Sulfur burns according to the equation:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

When 4 g of sulfur is burned at SLC, the volume of sulfur dioxide gas is approximately:

A 1.53 L

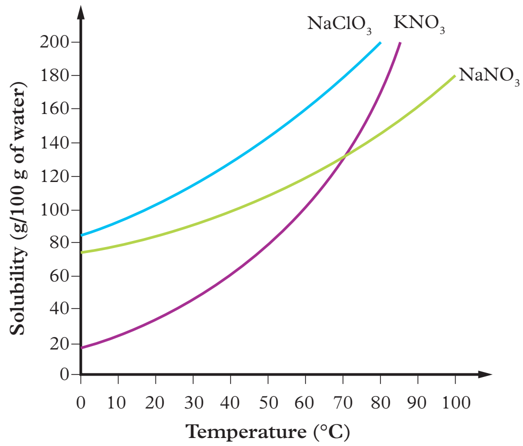
B 3.09 L

C 6.1 L

D 12.3 L

The following information relates to Questions 4 and 5.

The graph shows solubility curves for three metal salts: potassium nitrate (KNO_3), sodium nitrate ($NaNO_3$) and sodium chlorate ($NaClO_3$).



Question 4

100 g of a saturated solution of KNO_3 was cooled from 70°C to 40°C. What mass of KNO_3 would be expected to have crystallised from the solution?

A 56 g

B 60 g

C 70 g

D 80 g

Question 5

40.0 g samples of each salt were placed into separate beakers. 50.0 g of water was added to each beaker, and the beakers were heated to 40°C. In which beaker(s) would you expect all of the salt to dissolve?

A None of the beakers

B The KNO_3 beaker only

C The $NaNO_3$ and $NaClO_3$ beakers only

D All three beakers

Question 6

A particular brand of vodka has an ethanol content of 40.0% v/v. Given that the density of ethanol is 0.785 g mL^{-1}, the mass of ethanol in a 40.0 mL glass of vodka is:

A 12.6 g
B 16.0 g
C 20.4 g
D 31.4 g

Question 7

Anhydrous sodium carbonate (Na_2CO_3) is a very useful substance in an analytical chemistry laboratory. The mass (in g) of sodium carbonate needed to make 500 mL of a 0.050 M aqueous solution is:

A 0.265
B 0.530
C 1.06
D 2.65

Question 8

What volume of water must be added to 7.0 mL of 10 M nitric acid to make a 3.5 M solution of HNO_3?

A 7 mL
B 13 mL
C 20 mL
D 27 mL

Question 9

The process for making sulfuric acid involves several stages. A key reaction is the conversion of sulfur dioxide to sulfur trioxide according to the equation:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

The maximum amount of sulfur trioxide (in L) that can be prepared from 100 L of SO_2 and 100 L of O_2, if all gases are the same temperature and pressure, is:

A 50
B 100
C 150
D 200

Question 10

What is the concentration (in mol L^{-1}) of a sodium hydroxide solution if a 20.00 mL aliquot is titrated against a 0.05 M sulfuric acid solution and the average titre was found to be 17.5 mL?

A 0.009
B 0.04
C 8.0
D 0.09

Short answer

Question 1 (3 marks)

A solution contains 2 g of LiCl dissolved in 250 mL. Calculate the concentration of the solution in the following units.

a mg L^{-1} 1 mark
b % (m/v) 1 mark
c mol L^{-1} 1 mark

Question 2 (9 marks)

Windex™ is a brand of window cleaner that contains ammonia.

A Chemistry class carried out a titration of aliquots of Windex™, using a 20 mL pipette, against a 0.102 M solution of hydrochloric acid. Phenolphthalein was the indicator.

A student obtained the following titres of hydrochloric acid: 22.37, 21.21, 21.28, 21.25 and 21.19 mL.

a Construct the balanced chemical equation for the reaction, including states. 1 mark
b What is the average titre? 1 mark
c What is the amount (in mol) of hydrochloric acid in the average titre? 1 mark
d What is the amount (in mol) of ammonia in each aliquot of Windex? 1 mark
e What is the concentration (molarity) of ammonia in the Windex? 2 marks
f During each titration, when would the student stop adding the acid? 1 mark
g Another student obtained titres of 21.2, 21.4, 21.3 and 21.6 mL. Discuss their precision and accuracy. 2 marks

Question 3 (8 marks)

A farm dam was tested for water quality. Three separate 100.0 mL samples were taken for analysis.

a To determine the amount of solids dissolved in the water, one 100 mL sample was filtered and the resultant liquid was evaporated to dryness in an evaporating basin. The results are shown below.

Mass of evaporating basin (g)	42.19
Mass of evaporating basin and solid (g)	42.67

i Calculate the percentage by mass of the total dissolved solids in the water from the dam. Assume that the density of the water is 1.0 g mL^{-1}. 2 marks

ii What step could be taken to improve the accuracy of the value in part a(i)? 1 mark

b The second 100 mL sample of water was analysed for phosphate ion (PO_4^{3-}) concentration. The phosphate ion in the water was reacted with a special reagent to produce a coloured solution. The absorbance of the coloured solution was measured to find the phosphate ion concentration. The results of the analysis of the solution and a number of standard solutions containing the phosphate ion are shown in the following table.

Standard	1	2	3	4	Sample
PO_4^{3-}	0.05	0.10	0.15	0.20	?
Absorbance	0.103	0.205	0.308	0.410	0.331

The calibration process produced a straight-line graph with the relationship:

$$\text{absorbance} = 2.05 \times \text{concentration}$$

The recommended maximum concentration of phosphate ion in water is 1.04×10^{-6} M. Does the water from the dam comply with this limit? Show any working. 2 marks

c The remaining 100.0 mL water sample was analysed for the chloride ion (Cl^-) concentration. An excess of silver nitrate solution was added to the water sample. The silver chloride precipitate that formed was isolated, washed and dried. The mass of the precipitate formed was 0.698 g.

i Construct the ionic equation for the precipitation reaction.

ii Calculate the chloride ion concentration in the dam water in parts per million (ppm).

3 marks

Question 4 (2 marks)

The table shows the relationship between properties of various samples of compounds or elements. Complete the table by indicating whether Quantity 1 is greater than (>), less than (<) or equal to (=) Quantity 2. The first row has been completed as an example.

Quantity 1	>, <, or =	Quantity 2
Number of atoms of helium in 1.0 mol of He	=	Number of atoms of Ne in 1.0 mol of Ne
Number of molecules of CO_2 in 2 mol of CO_2		Number of oxygen atoms in 1 mol of CO_2
Volume of 0.20 M HNO_3 (a strong acid) required to completely react with 20.0 mL of 0.20 M KOH solution		Volume of 0.20 M HCN (a weak acid) required to completely react with 20.0 mL of 0.20 M KOH solution

2 marks

Question 5 (5 marks)

You are given a bottle containing small crystals of white solid and three beakers containing solutions of this solid. One of these solutions is an unsaturated solution, one is a saturated solution and the other is a supersaturated solution.

a What is meant by the terms 'unsaturated', 'saturated' and 'supersaturated'? 3 marks

b Describe a simple experiment that would distinguish between the saturated and unsaturated solutions in the beakers. 2 marks

Question 6 (5 marks)

An unknown chemical compound was analysed by a number of techniques.

The elements in the compound were found to be carbon, hydrogen and fluorine. 1.739 g of the compound was then burned and found to contain 0.887 g carbon and 0.149 g hydrogen.

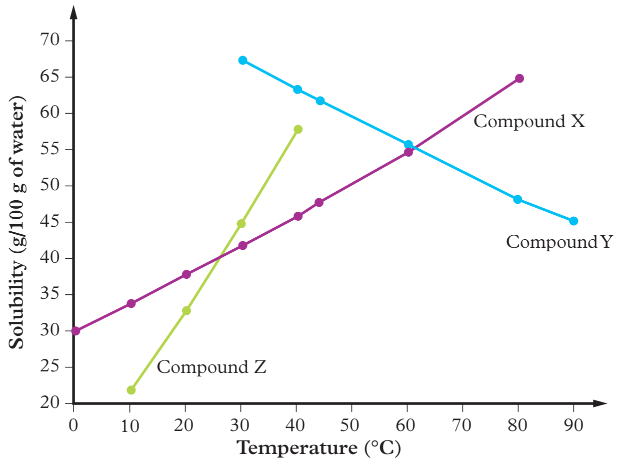
a Use the information to find the empirical formula of the compound. 1 mark

b 2.190 g of the compound was vaporised at SLC, where it had a volume of 0.578 L. Use this information to find the amount (in mol) of the vapour and hence the molar mass of the compound. 3 marks

c Use the answers to parts **a** and **b** to find the molecular formula of the compound. 1 mark

Question 7 (8 marks)

The graph shows how solubility changes with temperature for three compounds, X, Y and Z.



a Which one or more of the compounds (X, Y and Z) is a gas? Give reasons for your answer. 2 marks

b What is the solubility of compound Y at 60°C? 1 mark

c What volume of water (to the nearest mL) at 40°C is needed to make a saturated solution from 50 g of compound X? 2 marks

d If 80 g of compound Z is added to 60 g of water at 40°C, what mass of Z would remain undissolved? 3 marks

TOTAL MARKS

/50 marks

CHAPTER 18

Student-designed investigation

KEY KNOWLEDGE

Investigation design

→ chemical science concepts specific to the selected scientific investigation and their significance, including the definition of key terms

→ scientific methodology relevant to the selected scientific investigation, selected from the following: classification and identification; controlled experiment; fieldwork; modelling; product, process or system development; or simulation

→ techniques of primary qualitative and quantitative data generation relevant to the investigation

→ accuracy, precision, repeatability, reproducibility, resolution, and validity of measurements in relation to the investigation

→ health, safety and ethical guidelines relevant to the selected scientific investigation

Scientific evidence

→ the distinction between an aim, a hypothesis, a model, a theory and a law

→ observations and investigations that are consistent with, or challenge, current scientific models or theories

→ the characteristics of primary data

→ ways of organising, analysing and evaluating generated primary data to identify patterns and relationships, and to identify sources of error

→ the use of a logbook to authenticate generated primary data

→ the limitations of investigation methodologies and methods, and of data generation and/or analysis

Science communication

→ the conventions of scientific report writing, including scientific terminology and representations, standard abbreviations and units of measurement

→ ways of presenting key findings and implications of the selected scientific investigation

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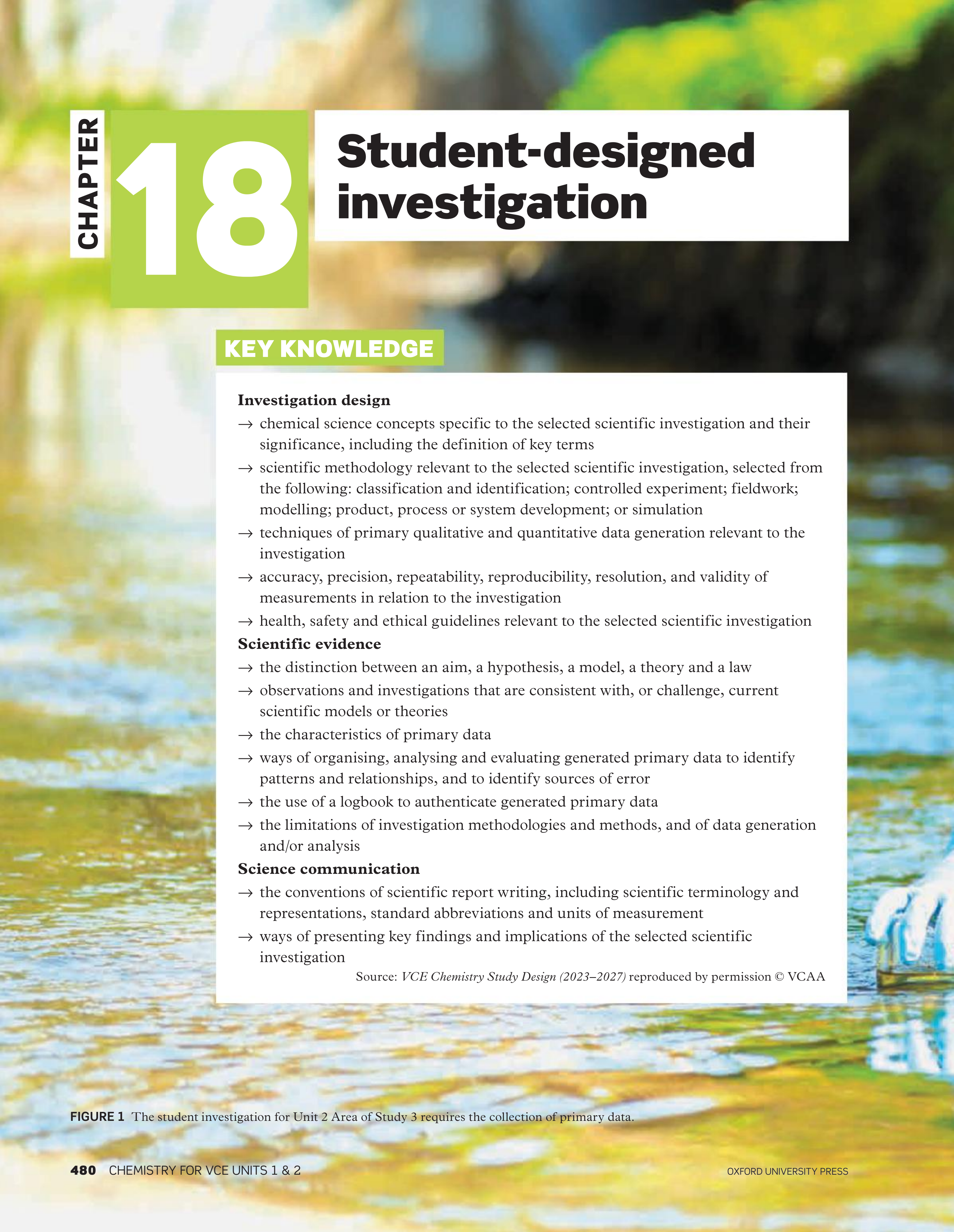


FIGURE 1 The student investigation for Unit 2 Area of Study 3 requires the collection of primary data.

GROUNDWORK

In Chapter 18, you will learn about how you can prepare for your student-designed investigation, required for Unit 2 Area of Study 3.

This chapter will build on concepts you have already learnt in Chapter 10. So, before you begin the chapter, test yourself on the following questions to make sure you remember the basics.

18A What is primary data?

18A Groundwork resource
Primary sources

18B What is the difference between the independent, dependent and controlled variables?

18B Groundwork resource
Variables

18C What are the different graph types you can use to display a relationship between two variables?

18C Groundwork resource
Presenting data

18.1 Topic selection

KEY IDEAS

In this topic, you will:

+ write a scientific question
+ select an experiment design or adapt one
+ conduct research and define key theories
+ understand how to analyse the assessment criteria.

This chapter will guide you through your student-designed investigation. Each topic will focus on the essentials of investigation design, from planning and conducting the investigation, to communicating your results. Goals and to-do lists for each stage of the investigation are also included to help you.

Before you get started, make sure you have the details of this assessment task from your teacher. You should always check what the assessment criteria will be and note any tips or instructions your teacher might provide to help you succeed.

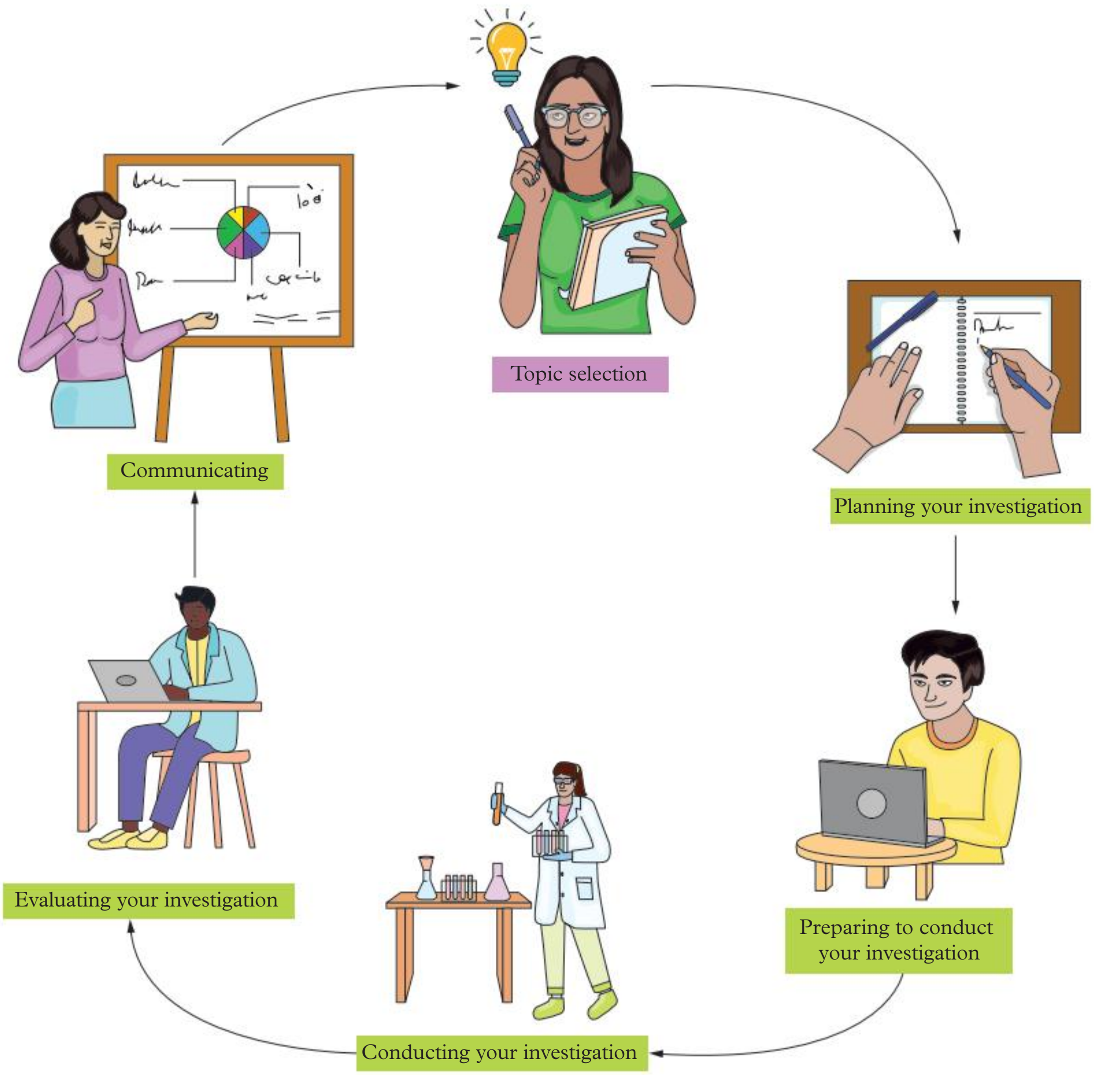


FIGURE 1 The scientific investigation process begins with the topic selection phase.

A scientific investigation is a multistep process. In Figure 1, you can see that the process can cycle, specifically when the final steps of your first investigation prompt a second investigation. Following the scientific process is essential to the success of the investigation.

This chapter will go through each phase of the process so that you can complete your own scientific investigation for Unit 2 Area of Study 3 'How do quantitative scientific investigations develop our understanding of chemical reactions?'

The start of your investigation begins with the topic selection phase. This phase includes:

- deciding on your topic
- deciding to design or adapt an experiment
- choosing and writing an investigation question
- conducting research and defining key theories
- analysing the assessment criteria.

Deciding on your topic

Your teacher may have already assigned you a specific topic or set of topics to select from. However, if you are deciding on your own topic, then you must select one from Unit 2 Area of Study 1 or 2.

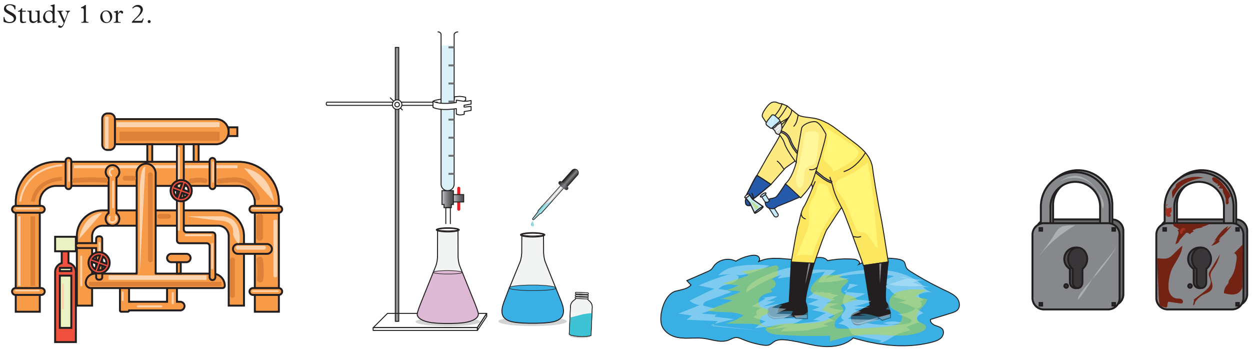


FIGURE 2 Your investigation topic must relate to concepts covered in Unit 2 Area of Study 1 or 2.

Reading through Chapters 11–17 can help refresh your memory of these topics. When deciding on your topic, you may also wish to consider investigating how chemistry concepts from one of these topics relates to the United Nations Sustainability Development Goals. Your investigation could relate to sustainability and access to clean water, air and soil.

FIGURE 3 Your research topic could integrate links between chemistry concepts and the United Nations Sustainability Development Goals, such as analysis of clean water and its distribution.

Topic selection can also be inspired by many other areas. Table 1 presents some potential sources for your topic selection and a thinking prompt you can use to help assess your potential interest in a topic.

TABLE 1 Sources and ideas to help inspire your topic selection

Potential source	Thinking prompt
Direct observation of and curiosity about an object, an event, a phenomenon, a practical problem or a technological development	Have any science concepts, real-world chemistry extracts or scientific articles gained your interest?
Anomalous or surprising investigation results or findings from analysis of qualitative and/or quantitative data	Throughout practicals that you have completed, has anything occurred that particularly spiked your interest or intrigue?
Extension of a previous inquiry	Have any inquiry tasks in practicals led you to ask further questions?
Research involving secondary data	Have you come across something that others have investigated that you might like to expand upon or look into further?

Whatever the source of inspiration for your topic, you should aim to select a topic that interests you. This will make the process more enjoyable and allow you to engage more with it. Topic selection can also occur while choosing and writing a question and/or designing or adapting an experiment.

Deciding to design or adapt an experiment

When you are looking at potential experiments to identify or test a research question, you should consider the following:

- Has your teacher outlined an experiment or set of experiments you have to use to complete your investigation?
- Did you complete an experiment during Unit 2 that left you asking a question, or collect any data that didn't quite align with your hypothesis?

If either of these is the case, then reusing or adapting an experiment that you have completed during Unit 2 will be the best way to complete your investigation.

If you have a topic you intend to investigate and haven't previously completed a practical activity on it, you can look online for experiments to use or adapt. You may also design your own experiment. Table 2 presents some of the different types of inquiry methods you can use to collect primary data.

If you are designing your own experiment, or reusing or adapting one, make sure to consider:

- whether you can complete the experiment in the time you have been given
- the equipment you have access to
- whether the experiment allows you to collect primary data.

Research tip

Always write your investigation question as a question and not as a statement.

TABLE 2 Inquiry methods that can be used to collect primary data

Inquiry type	Inquiry outline	Inquiry question types	Example questions
Controlled experiment	An investigation of the relationship between an independent variable (IV) and a dependent variable (DV), while controlling all other variables (CVs).	*What effect does … have on …?* *Is … related to …?*	What effect does temperature have on the pH of water? Is pH related to the electrical conductivity of water?
Pattern seeking	The investigation of one variable to determine what other variable(s) can affect it, and to what extent other variables may be important in their effects on the variable under investigation. This can include observing natural events or phenomena and identifying patterns and/or relationships, and then proposing a link. It may involve multiple variables, because some variables may be difficult to control.	*What factors affect …?* *What are the optimal conditions for …?*	What factors affect the corrosion of metals? What are the optimal conditions for soils, to grow soybeans in?
Single-variable exploration	The investigation of one variable or factor at a time, usually to see how it changes over time, focusing on observations and identification of a phenomenon. This type of inquiry can lead to questions about the causes of an observed phenomenon and prompt further investigation.	*How does … change over time?* *Do all …?* *When do(es) …?*	How does the pH of water change over time? Do all metals rust in water? When do metals rust?

Choosing and writing an investigation question

A scientific inquiry or investigation involves asking or responding to an investigation question and then performing experiments and reporting on your findings. Your investigation and report should always respond to the question that you initially started with.

Once you have a topic or experiment in mind, you should start to consider questions that you have about that topic. You could also start with a general question that you have come across during Unit 2 and narrow this down into a research question. How you refine your question can vary depending on whether you start off with a topic or idea or with an experiment in mind.

FIGURE 4 Selecting an investigation topic that interests you will make the investigation more worthwhile and enjoyable.

Starting with a topic or idea

Figure 5 shows a process that could help you write an investigation question starting with a topic or idea. This process starts with the topic or idea, considers the theory involved and then structures questions around the relevant theory.

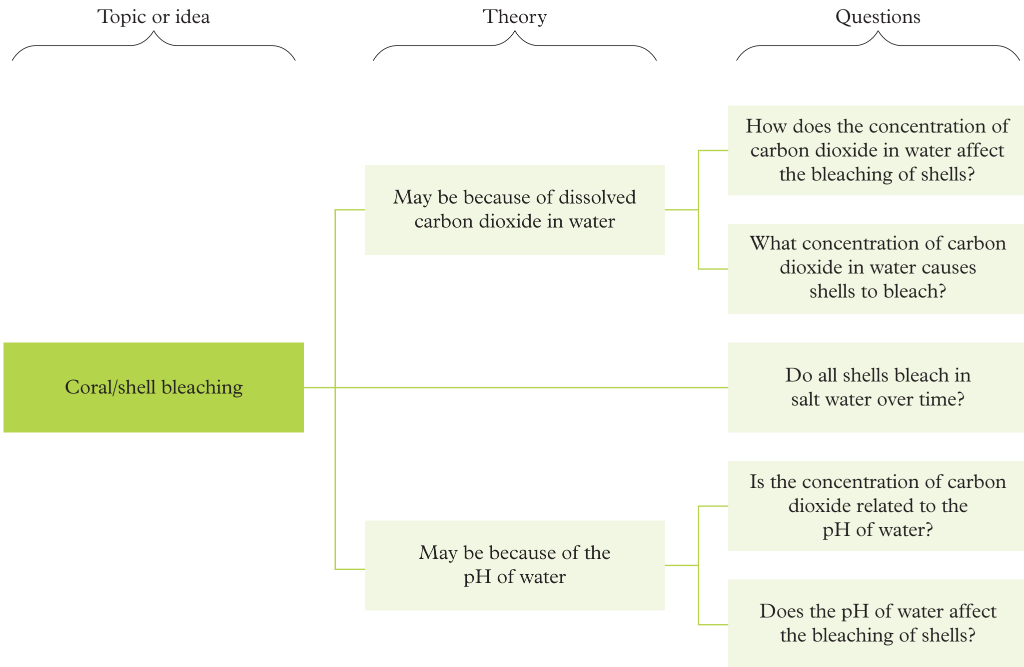


FIGURE 5 An example of the process involved in developing a research question from a topic or idea

Starting with an experiment

If your teacher has given you an experiment or a set of experiments to complete, or you completed an experiment and it left you with questions, you could use the process in Figure 6 to help build your question.

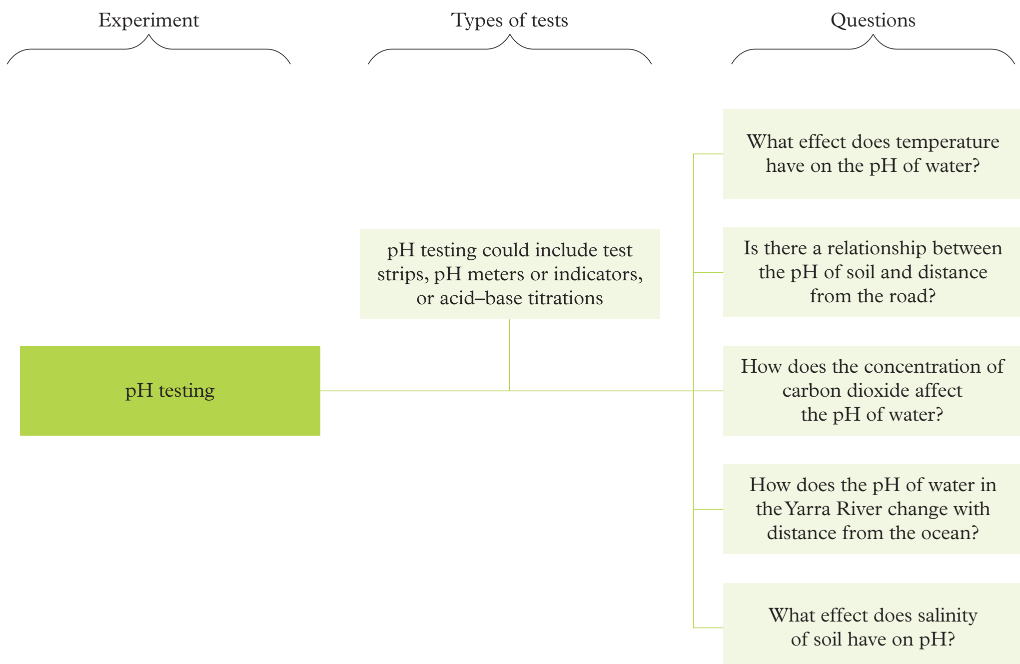


FIGURE 6 An example of the process involved in developing a research question from an experiment

After completing the processes in Figure 5 or 6, you might have a variety of questions to choose from. You might also need to go back a step and look at designing or adapting an experiment that will fit with your question and the time allocated. Whatever investigation question you decide on, it should:

- be clear and focused
- have an appropriate scope (not too vague or too narrow)
- not be too easy or difficult to answer (the question should require more than a simple yes or no answer)
- be researchable
- be analytical rather than descriptive.

Research tip

Read your proposed question to a friend or teacher and ask them if they think it delivers on all five of the key features of an investigation question.

Conducting research and defining key theories

Once you have your topic or experiment and you have a good scientific question, you should conduct some research into the theories surrounding your topic.

Research will help you to:

- understand the key theories or phenomena surrounding your topic
- find information to write your introduction
- determine better experiments or methods for conducting your experiment
- write your hypothesis and understand the reasoning behind your prediction
- better understand the data that you might need to collect
- fill any current gaps in your knowledge.

Before starting to research and compiling your key theories, read the information on sourcing information and analysing your chosen secondary sources by using the CRAAP method on page 10.11 of Chapter 10.

FIGURE 7 After selecting your topic or experiment, you should research the theories surrounding your topic.

Analysing the assessment criteria

When given your assessment criteria for your Unit 2 Area of Study 3 Scientific investigation, you should check to see what is required to achieve the highest marks. You can do this by annotating the rubric or criteria (Figures 8 and 9). You can also use the assessment criteria as a checklist (Figure 10).

HYPOTHESIS
NOT SHOWN
2.1 Makes predictions
2.2 Structures prediction by distinguishing between variables
2.3 Justifies predictions using known theory

Use an 'If ____, then____ because______.' statement where you use the variables and put in why it will happen.

Use an 'If ____, then_____.' statement where you have used the dependent variable to make this statement.

CONCLUSION
NOT SHOWN
12.1 Summarises key findings
12.2 Links key findings to hypothesis
12.3 Identifies limitations in experiment
12.4 Makes recommendations to overcome limitations

Explains the aim and how you achieved it.

Did what you found in the experiment support your hypothesis? Was it correct?

A limitation is a restriction that the experiment had on your conclusion. Can you think of the limitations for this experiment?

Suggest ways in which you could do the experiment where the limitations wouldn't apply.

FIGURE 8 An example of a developmental rubric for a hypothesis and a conclusion, and annotations of what to do to get each mark

Investigation question should be formulated and be:
- **testable**
- **clear and focused**

and have appropriate scope.

Very low	Low	Medium	High	Very high
Some attempt at formulation of an investigable question with very limited outline of investigation design	Mostly appropriate formulation of an investigable question with limited outline of investigation design	Appropriate formulation of an investigable question with sound investigation design	Accurate formulation of an investigable question with well-constructed investigation design	Highly proficient formulation of an investigable question with sophisticated investigation design

This rubric descriptor also includes the overall investigation design, which should link to the:
- **question**
- **topic (soil, water, gases)**

or
- **the practical your teacher set.**

FIGURE 9 An example of a performance descriptor rubric for the investigation question and investigation design with annotations of the breakdown of the top marks

Question:
- **is suitable for volumetric analysis or pH**
- **is original**
- **fits the theme.**

Research question	Question posed is suitable for scientific investigation using either volumetric analysis or pH	/1
	Evident creative effort to generate an original question that falls within one of the set themes	/1
Aims and hypothesis	The purpose of the research is made clear with a concise aim that demonstrates a clear statement of intention	/1
	The hypothesis clearly defines the dependent and independent variables	/1
	The hypothesis uses prior scientific understanding to predict a plausible outcome	/1

Aim:
- **is clear**
- **is concise**
- **includes an intention.**

Hypothesis:
- **uses an 'If ___, then ____ because ____' statement**
- **includes IV**
- **includes DV**
- **predicts theory.**

FIGURE 10 An example of an assessment criteria rubric and annotations with checklists for each step

Whatever the assessment criteria you are given, remember to check through it at the start to help you determine what you must do and include. You should also go through the criteria again at the end and double check that you have included everything.

FIGURE 11 Make sure you check the assessment criteria at the start and at the end of your investigation.

TO-DO LIST

- Write a scientific question that you would like to investigate.
- Make sure your question is posed as a question and not a statement.
- Select an experiment to adapt or design for your investigation.
- Define the key scientific theories and terms that are relevant to your investigation topic.
- Analyse the criteria you will be marked on for this assessment.

18.2

Planning your investigation

KEY IDEAS

In this topic, you will:

- define your variables
- write a suitable hypothesis for your experiment
- construct a methodology for your experiment.

After you have selected your topic, the next phase involves planning your scientific investigation. This will include:

- defining the variables
- writing a hypothesis
- designing a method (including materials and equipment).

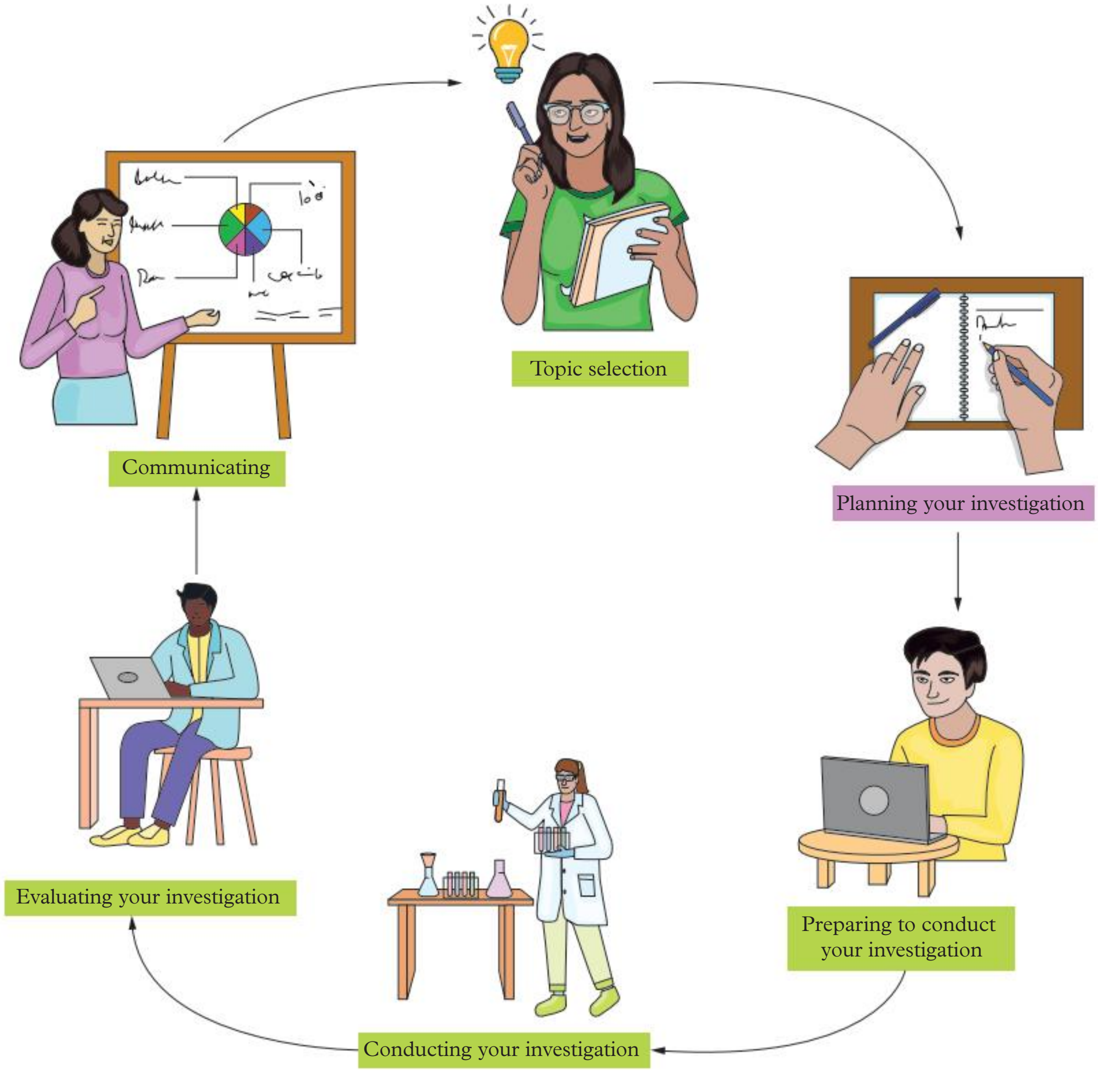


FIGURE 1 The scientific investigation process – this topic is about planning your investigation.

Defining the variables

Before you can write a hypothesis and method for your investigation, you must define your variables. The three types of variables to consider are the:

- independent variable (IV) – the thing you change
- dependent variable (DV) – the thing that you measure
- controlled variable(s) (CV) – the thing(s) you control.

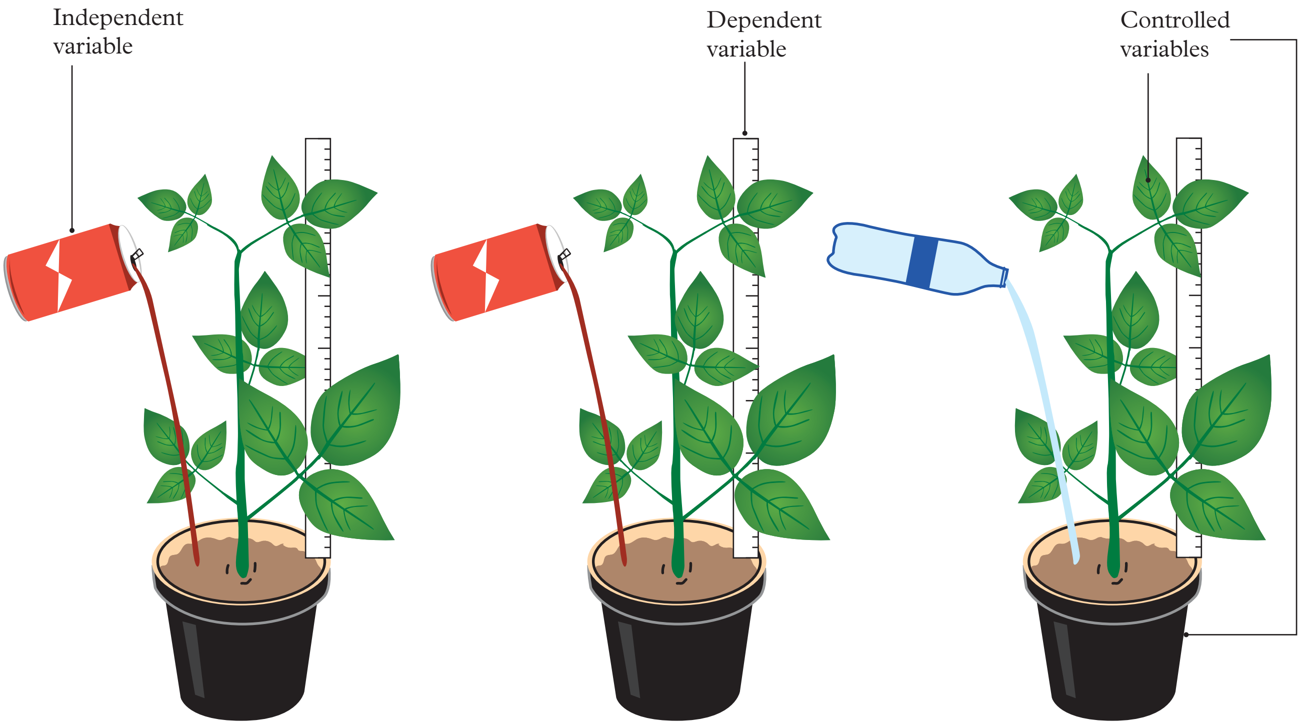


FIGURE 2 When testing the effect of soft drink on plant growth, the IV = soft drink, DV = plant height, and controlled variables could include pot plant size, amount of liquid given, soil type and light exposure.

Writing a testable hypothesis

Chapter 1 Chemistry toolkit showed you how to write a hypothesis. Now it's time to put this skill into practice. To write a hypothesis, you can use an IF/THEN/BECAUSE statement to make sure you cover all the elements you need.

If	Then	Because
If the **independent variable** is *[changed]*	then the **dependent variable** will *[change]*	because of **scientific reasoning**.
E.g. increased, decreased	*E.g. increase/decrease the amount/ rate/height/weight/number*	*A possible explanation for the relationship between the IV and DV*

FIGURE 3 A summary of how to write a testable hypothesis

A useful hypothesis is a testable statement that often includes a prediction. In some instances, a research question may not lend itself to having an accompanying hypothesis. In such cases, students should work directly with their research questions. If this applies to your research question, check with your teacher before moving on, because some criteria might have marks allocated for a testable hypothesis and you may need to change your experiment.

Research tip

To remember the IV and DV roles and how you would graph them you can use DRY MIX:
DRY – **D**ependent (**R**esponding) **Y**-axis
MIX – (**M**anipulated) **I**ndependent **X**-axis

18.2 WORKED EXAMPLE

WRITING A TESTABLE HYPOTHESIS FOR A RESEARCH QUESTION

Research question: How does the pH of soil affect soybean growth?

Think	Do
Step 1: Define your variables.	IV: Soil pH DV: Height of plant (mm) Controlled variables: • Soil type • Amount of soil • Sunlight amount • Watering (time of day and volume of water) • Day/time plants are measured
Step 2: Write your hypothesis.	Hypothesis: *If* the pH of the soil is increased or decreased from neutral, *then* the soybean plant will be shorter *because* soybean optimal growth occurs at a soil pH 7 (neutral).

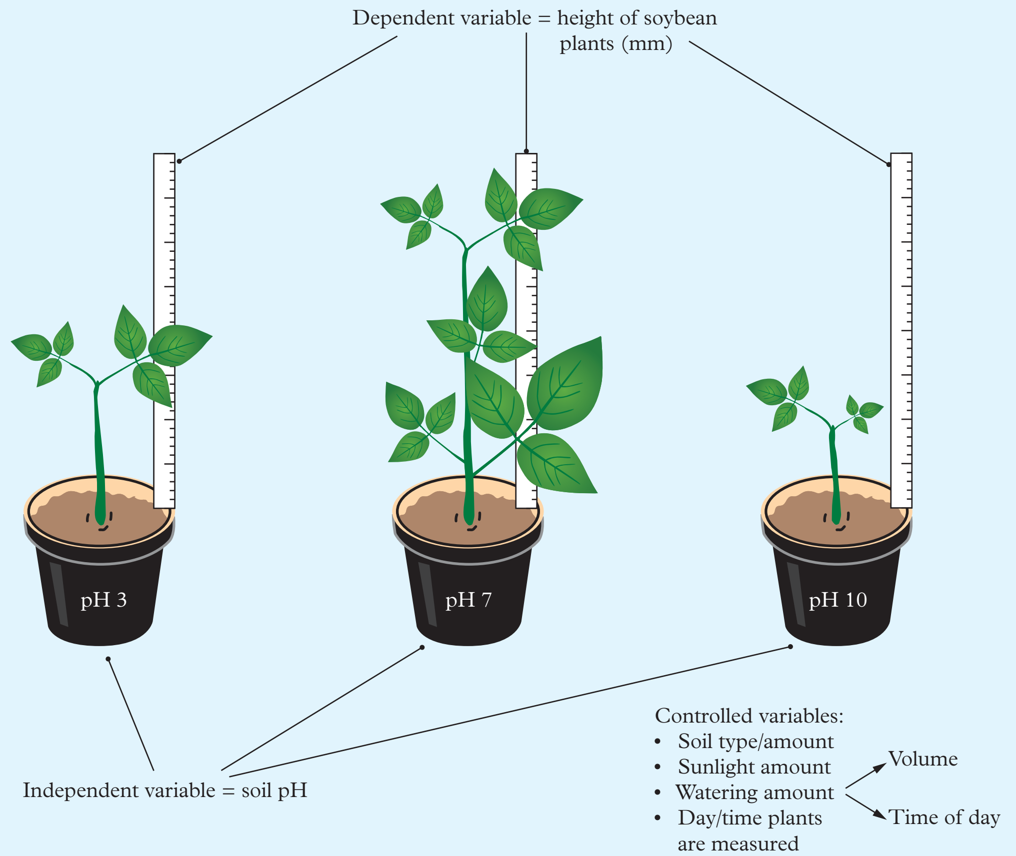


FIGURE 4 Controlled experiment demonstrating variables for the question 'How does the pH of soil affect soybean growth?'

Choosing a methodology

You can approach your scientific question in a number of ways. How you choose to approach your investigation is called your methodology. Different scientific methodologies include:

- controlled experiment
- fieldwork
- modelling
- simulation
- case study
- classification and identification
- product, process, or system development.

Information on each of these methodologies is provided in Chapter 1 Chemistry toolkit. The flow chart in Figure 5 can help you to choose the methodology best suited to your scientific question.

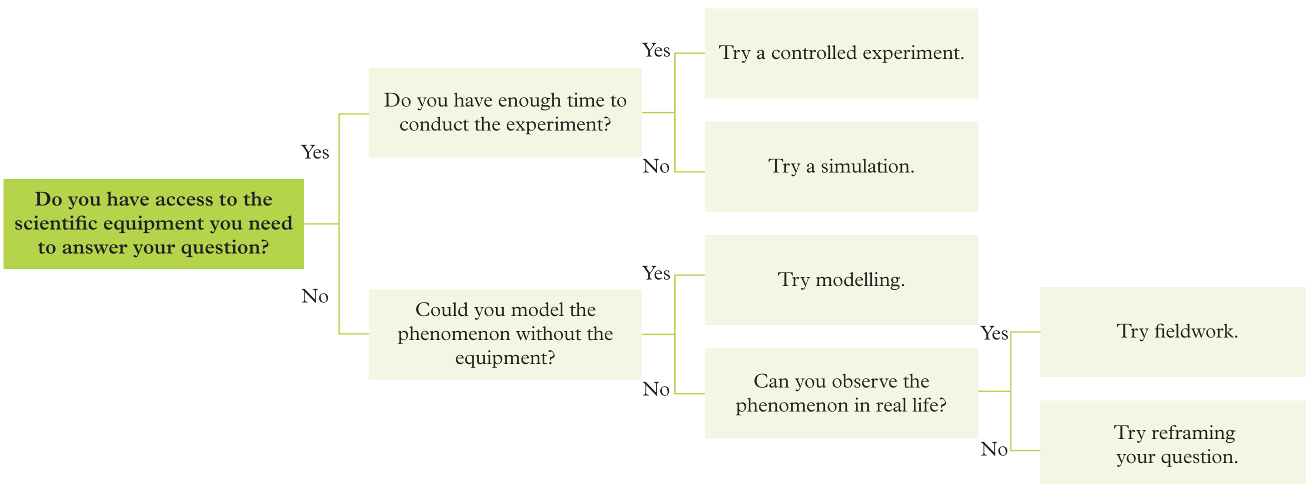


FIGURE 5 A flow chart to assist with selecting your scientific methodology

Generating primary data

For this investigation, you will need to generate your own primary data. There are two types of primary data that you can collect:

- **quantitative data –** numbers, things that are counted or measured and given a numerical value
- **qualitative data –** descriptive rather than numerical, using observations, it is more categorical.

Tests you can run to generate primary data are summarised in Table 1.

FIGURE 6 Numeric values, such as weight, is quantitative data

TABLE 1 Approaches to collecting primary data

What kind of data do you collect?	Test	How do you do it?
Quantitative data	pH	pH test strips Universal indicator pH meter
	Mass	Gravimetric analysis Weighing mass change with an analytical balance
	Temperature	Using a thermometer
	Electrical conductivity	Using an electrical conductivity kit
	Measuring concentration	Volumetric analysis
Qualitative data	Corrosion of metals	Observing the amount or colour of rust
	Production of gases	Looking for bubbling Using the pop test for hydrogen gas
	Presence of CO_2	Adding limewater and observing results

Writing your method

For detailed information on how to write up your method, refer to Chapter 1 Chemistry toolkit. Key things to remember when writing a method are to:

- include detailed sequential steps:
  - ensure the steps are in correct order
  - number each of your steps
- ensure that your controlled and independent variables are appropriately manipulated:
  - the independent variable is the thing you change, so that change should be in the method correctly
  - the controlled variables are the things you need to control during the experiment; providing detail on this is important for replication (see below)
- include appropriate measurements and concentrations with accurate use of units:
  - include the quantity of each item
  - include correct units
  - include the correct equipment
- write the method in past tense.

TO-DO LIST

- Choose a methodology for your investigation and justify why you have chosen this as the best way in which to answer your scientific question.
- Identify the independent and dependent variables that you will be investigating.
- Identify which variables you will control for your investigation.
- Write a testable hypothesis for your investigation, using the 'If …, then … because …' style.
- Write a succinct and detailed method for your investigation. (Remember to include enough detail so someone else could conduct your investigation for you – aim for repeatability and reproducibility.)

18.3 Preparing to conduct your investigation

KEY IDEAS

In this topic, you will:

+ complete a risk assessment
+ consider the ethics around your experiment.

In this topic, we will look at how to prepare to conduct your scientific investigation. Preparing to conduct your investigation includes**:**

- conducting a risk assessment
- considering the ethics of your investigation.

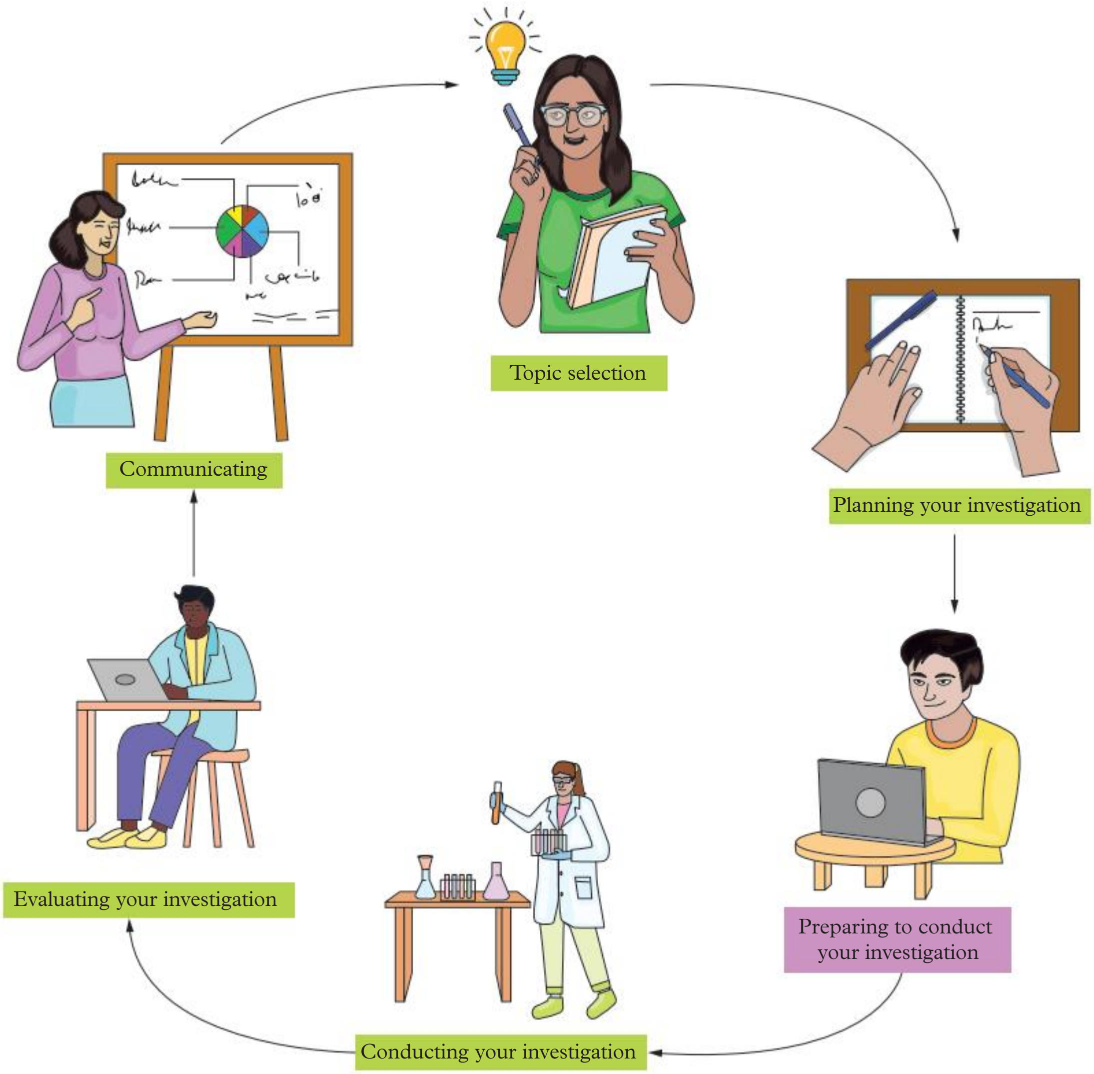


FIGURE 1 The scientific investigation process – this topic is about preparing to conduct your investigation.

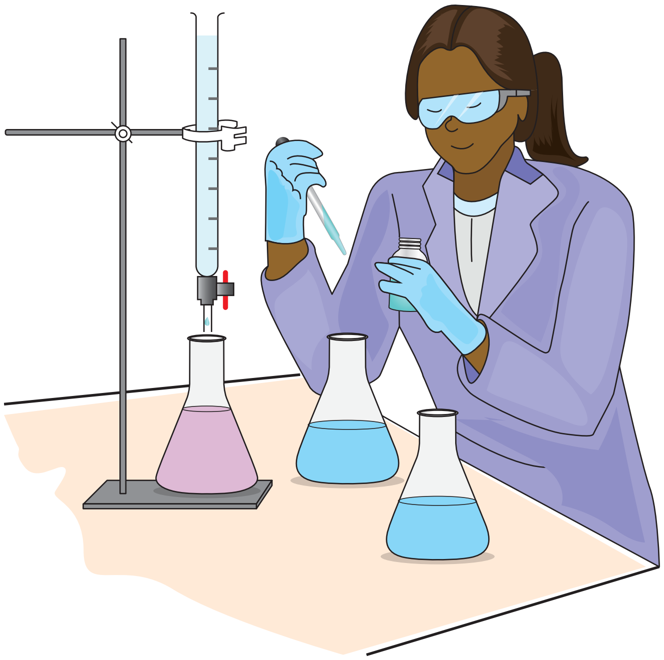


FIGURE 2 Personal protective equipment (PPE), including safety glasses and lab coats, should be worn to protect you during an experiment.

Conducting a risk assessment

When conducting an experiment, there are many health and safety considerations you need to take to protect yourself and those around you in the laboratory. Before you conduct an experiment, you must conduct a risk assessment. Risk assessments help you identify all the potential risks associated with your materials, chemicals and health and safety.

The instructions for conducting a risk assessment for an experiment are outlined in Chapter 1 Chemistry toolkit (page 17). A risk assessment template is also provided in your Student obook pro for you to download and print.

Ethical considerations

For your scientific investigation, you will need to generate primary data. This means when writing and planning your investigation, you will need to consider any potential ethical considerations. This includes considering the impact of your research on:

- the greater community
- the environment
- living organisms and non-living things.

Other ethical factors to be aware of when preparing for your investigation include:

- the sourcing of materials
- disposal of chemicals and equipment
- the green chemistry principles
- the United Nations Sustainable Development Goals.

You are expected to apply integrity when recording and reporting results and data. This means you must not falsify results and you must acknowledge prior research. It is also important to consider the impact of your results and how they may be used in the future by others.

TO-DO LIST

- Conduct a risk assessment for your investigation.
- Evaluate the ethics of your investigation.

18.4 Conducting your investigation

KEY IDEAS

In this topic, you will:

+ create a data table in which to collect and record your data
+ prepare to conduct your experiment.

In this topic, you we will look at conducting your scientific investigation. This will include:

- generating and recording data
- conducting the experiment.

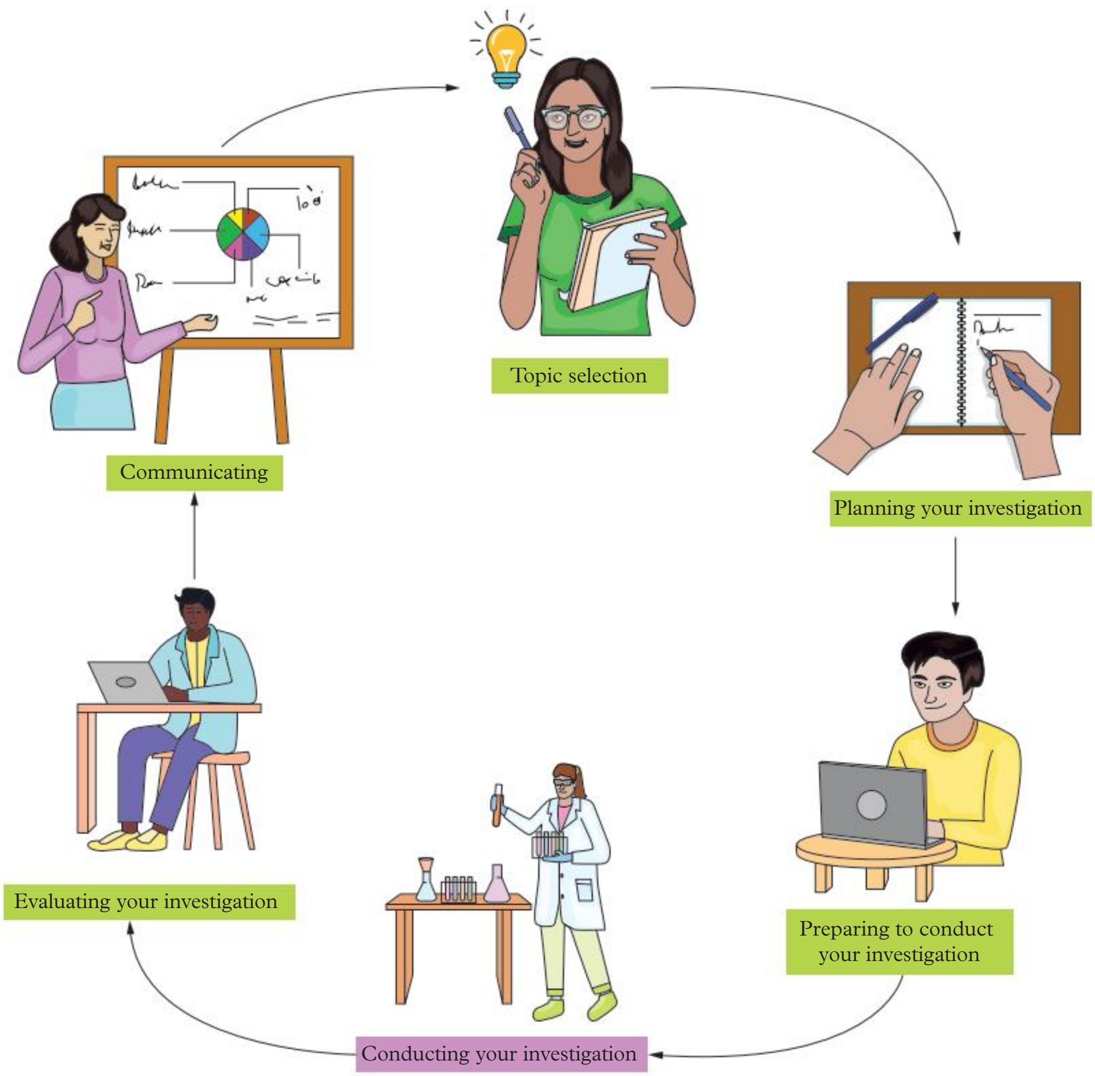
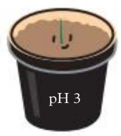
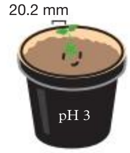


FIGURE 1 The scientific investigation process – this topic is about conducting your scientific investigation.

Recording results, data and observations

Establishing good data-recording habits will help you greatly when it comes to writing your discussion, and evaluating your errors. Data collection does not just include recording numbers and data, but also recording your observations. Figure 2 shows an example of an annotated and complete table for recording results, data and observations for the practical investigation 'How does the pH of soil affect the height of soybean growth?'

TEST	*IV: pH of soil*	*DV: Height of soybean plant (mm)*	*Observations/sketches*	
Day 5 Time: 10.30–10.40 am	3	20	• One tiny little green sprout has come up • No leaves • Has a little kind of bulging protrusion at the very top of the sprout – looks like it might be a leaf soon	
	7	35.5	• Largest of the three plants so far (as hypothesised) • Two tiny little leaves on sprout along with a little bulging protrusion at the top	
	10	N/A	• No growth – can't see any green • No change to the top of the soil since the day experiment was started	N/A
Day 10 Time: 10.30 –10.40 am	3	83.2	• The plant is larger than on day 5 • Two original leaves are now bigger (approximately 20.2 mm each), they are not as big as those on the pH 7 plant • Four new smaller leaves at the base of the stem • The stem is not as thick as that of the pH 7 plant • Stem and leaves are lighter in colour than those of pH 7 plant	
	7	110.7	• Largest of the three plants • Two original leaves are now bigger (approximately 32.0 mm each) and are right at the top of plant • Four new smaller leaves near bottom of the stem • Stem and leaves are much darker green in colour than those of pH 3 plant	
	10	24.1	• Small yellowish green sprout with two tiny little leaves that are just starting to unfurl • Did not measure the length of the leaves because I did not want to damage them to do this • Plant does not look healthy but there is growth	

Includes day and time that results were recorded, and observations made

Uses the same terminology to describe things

Links back to hypothesis

Even though there were no new differences, observations were still made

Compared results from different days and between different test conditions

Measured leaves (something that might not have been considered at beginning of experiment but worth recording)

Explains why no measurements were taken

Includes general thoughts on health of plant

FIGURE 2 An annotated and completed table for recording results, data and observations when testing the effect of soil pH on soybean growth

A considerable amount of time can elapse between when you first conduct the experiment and when you write it up. Good records of data, results and observation will be extremely helpful when you need to write because they will refresh the details of the experiment. Record your results in your logbook because these will need to be verified for authenticity by your teacher.

Research tip

When conducting an experiment, scientists constantly take notes and annotate changes to the method. This ensures the method can be reproduced by others in the future.

Conducting your experiment

When conducting your experiment, you should remember the following:

- Be safe at all times.
- Consider the safety of those around you.
- Follow your method or make annotations in your logbook when you change anything.
- Record all results at the time of taking them.
- Record any errors or things that seem out of place.
- Record observations and comparisons between tests.
- Check before disposing of experiment equipment, chemicals and materials.
- If you have any questions, always ask your teacher.

FIGURE 3 When setting up and conducting your research make sure you are following the correct safety procedures.

TO-DO LIST

- Set up a logbook so that you can record all results and observations as you conduct the investigation.
- Check with your teacher to make sure you are ready to begin generating data.

18.5 Evaluating your investigation

KEY IDEAS

In this topic, you will:

+ understand different ways to display your results
+ evaluate your results
+ identify and discuss any errors
+ write a discussion about your investigation.

In this topic, you will look at evaluating your scientific investigation. This includes:

- displaying your data
- writing a discussion
  - analysing your results
  - evaluating your errors
  - modifications to the method
- writing a conclusion.

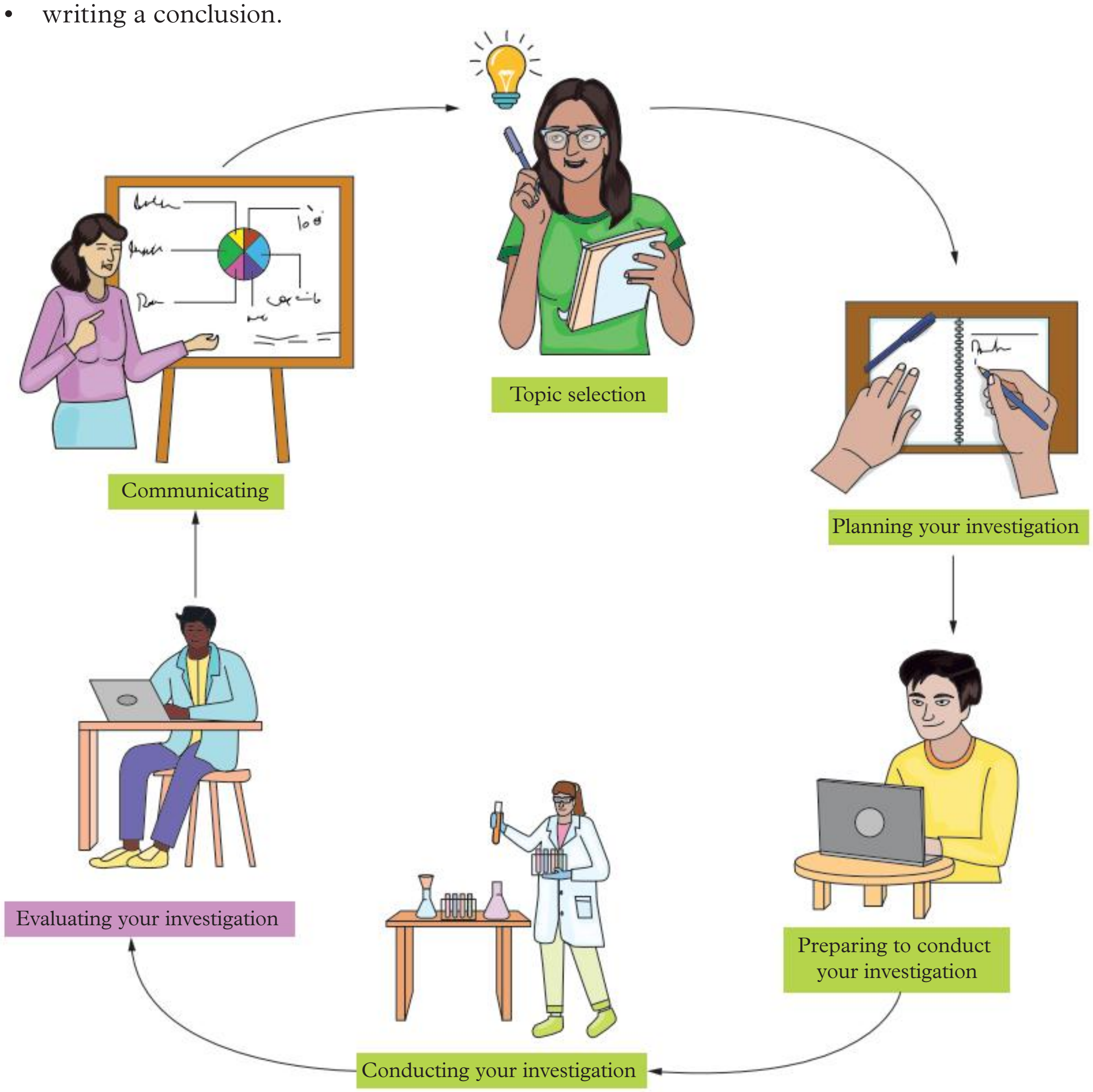


FIGURE 1 The scientific investigation process – this topic is about evaluating your investigation.

Displaying your data

After conducting your experiment and recording your data and observations, you will need to select the right way to display your data so that others can get a quick snapshot of your results. Your data presentation should also outline any trends that may exist between your variables.

The way in which you display your data is important; some graphs cannot be used for a particular purpose. For example, you cannot use pie charts to compare data. The flow chart in Figure 2 summarises different methods of displaying data depending on how you wish to use the information.

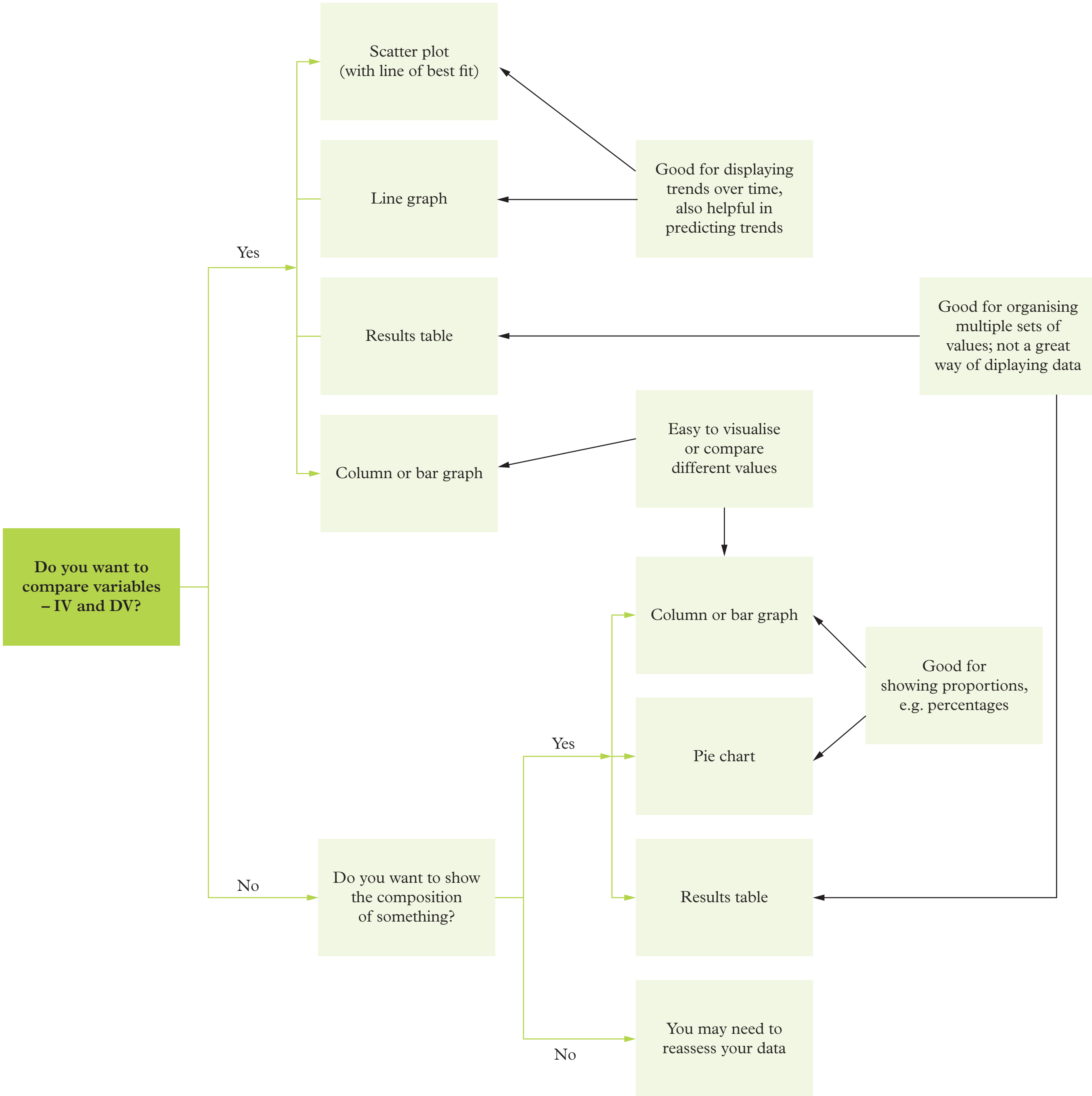


FIGURE 2 A flow chart to assist your decision on how you will display your data

Remember that what you are trying to show the people reading or assessing your investigation is the relationship between variables, or the trend or pattern from your results.

18.5A WORKED EXAMPLE

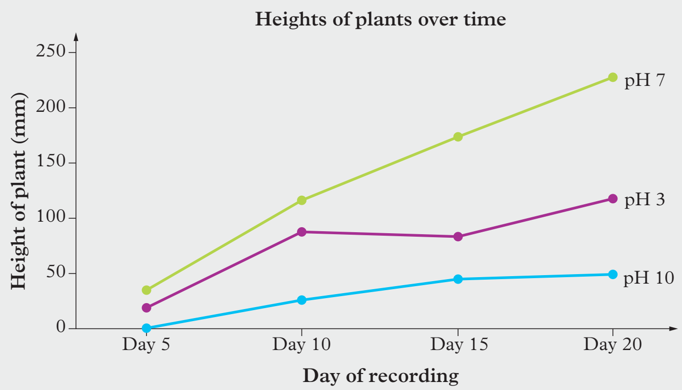
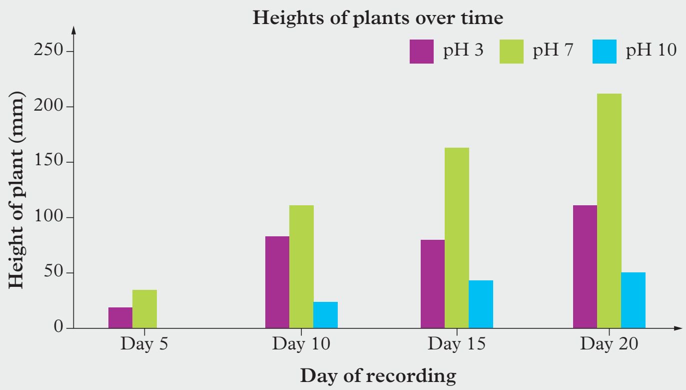
DISPLAYING DATA

The data in Table 1 was collected in an investigation exploring the research question: 'How does the pH of soil affect soybean growth?'.

TABLE 1 A results table of pH and height of plant collected over 20 days

Day	Height (mm) at pH:		
	3	7	10
5	20.0	35.5	0
10	83.2	110.7	24.1
15	80.4	163.0	44.4
20	111.2	212.7	50.0

Select and display the results from the table on the most appropriate graph.

Think	Do
Step 1: Use the flow chart in Figure 2 to determine which graph type best suits the data.	The data has variables to compare, so we could use a line graph or a bar chart. Both are presented here.
Step 2: Graph the data in the table.	Line graph of the heights of plants grown in different soil pH over 20 days:  Bar graph of the heights of plants grown in different soil pH over 20 days 

Research tip

Play around with different graph styles if you are having trouble deciding. There are graphing programs that will give you a lot of different types of graphs; if in doubt, ask someone to tell you what your graph says to them.

Think	Do
Step 3: Determine which graph will be most effective at displaying the information you want to get across.	Comparing the two graphs displaying the investigation data shows that the line graph is the superior display. It shows the trend of each pH condition over time. You can see the clear trends in each pH growth rate, and you can also clearly see that pH 7 is the best soil pH. In the line graph, you can see the odd measurement of pH 3 at days 10 and 15, which you cannot see as well in the bar graph. This graph is the clear choice for displaying the results data effectively.

Research tip

If you are drawing graphs that have an *x*- and a *y*-axis, then the IV is on the *x*- axis and the DV is on the *y*-axis.

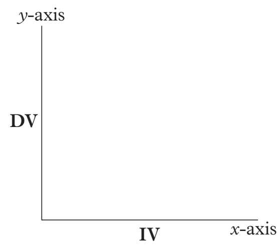


FIGURE 3 How to graph your IV and DV

Writing a discussion

The information you need to provide in your discussion depends on what the assessment criteria for your investigation says to include. For the most part, it will include:

- analysis of your results
- evaluation of your errors
- modifications of the method.

Analysis of your results

After displaying your data to show your results and trends, it is time to analyse and evaluate your results. To do this, there are a few questions you need to ask and answer:

- What trend can you see from the data?
- How does the trend relate to the theory?

One approach to addressing these questions is to simply go through each of the questions and answer them. If you need more guidance, another approach is to use the discussion flow chart in Figure 4. Figure 4 can help you understand how to answer each question by unpacking it, and format your responses so they can be included in your discussion.

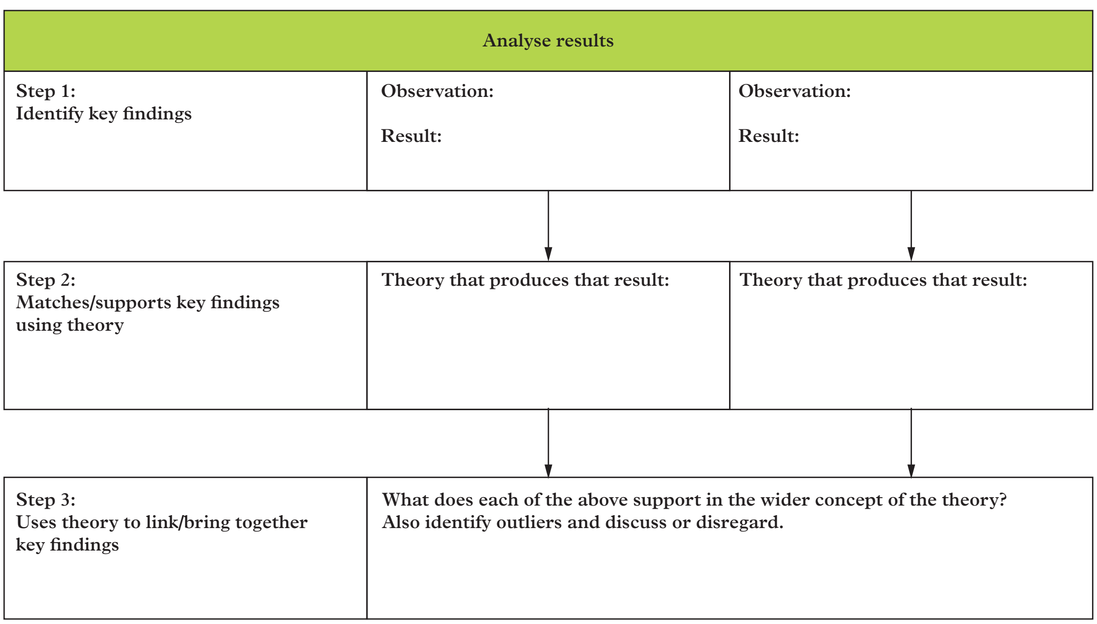


FIGURE 4 A flow chart for analysing results for discussion

For example, Figure 5 shows annotated results from an investigation on the effect of soil pH on soybean growth. You will need to use research and your understanding to link and support the findings from your experiment. You may use information you have already gathered from the topic selection phase. or conduct more research to inform and support your data and trends.

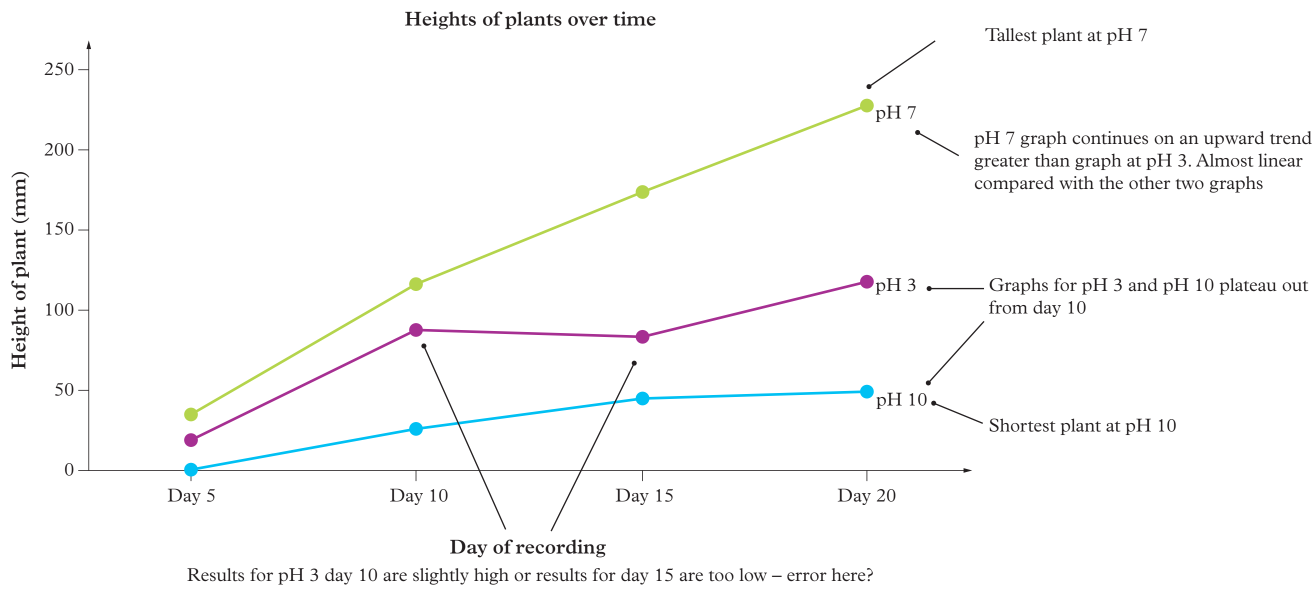


FIGURE 5 Annotation of the displayed results from the investigation of soil pH and plant growth

You may also need to gather more information if you find that your data or trend isn't supported by the theory, or if your hypothesis isn't supported. Figure 6 shows a worked example of how you can use the flow chart for analysis of data to inform your discussion.

Analyse results		
Step 1: **Identify key findings**	**Observation:** Tallest plants grew in soil of pH 7 **Result:** pH 7 is the best for growing plants	**Observation:** pH 10 plant growth plateaued after day 10 **Result:** pH 10 soils are not ideal for growing plants
Step 2: **Matches/supports key findings using theory**	**Theory that produces that result:** Soils are normally more neutral, with pH values of about 7, rather than higher or lower pH values, so plants are better adapted to growing in neutral pH soils.	**Theory that produces that result:** A soil with pH 10 does not occur naturally unless something is added to it that isn't meant to be there, so the plants have not adapted to this environment and will not grow.
Step 3: **Uses theory to link/bring together key findings**	**What do steps 1 and 2 support in the wider concept of the theory?** **Also identify outliers and discuss or disregard.** No outliers – but only one set of results for each: • Naturally occurring pH values are more neutral. • Plants don't grow well outside their pH range. • Plants become stunted and yellow outside their pH range, and their growth plateaus.	

FIGURE 6 An example of how the analysis of data flow chart can be used

Evaluating errors

Chapter 1 Chemistry toolkit explains types of errors, accuracy, precision and reproducibility, repeatability and reliability. You can refresh your memory of each of these by reading pages 24–25 in Chapter 1. Figure 7 shows how errors impact each of these factors.

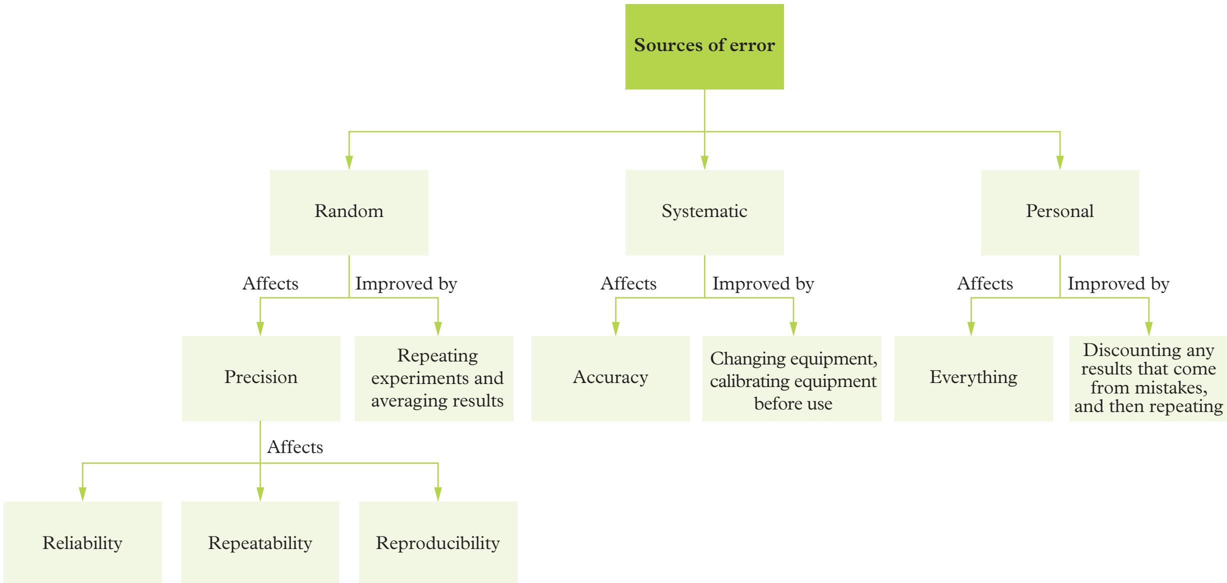


FIGURE 7 Errors and their effects and how they can be improved during an experiment

It is important that you evaluate your errors in your discussion. Figure 8 is a flow chart that shows how you can evaluate your errors and format your responses to be included in your discussion.

Evaluate errors		
Step 1: **Identify errors** **Types of error:** Personal Random Systematic Outlier	**Error:** **Type of error:**	**Error:** **Type of error:**
Step 2: **Explain reasons for errors**	**How did it come about/get introduced?**	**How did it come about/get introduced?**
Discuss effect of error on quality of experiment/data e.g. Higher or lower result and why	**How did it affect the result?**	**How did it affect the result?**
Step 3: **Propose ways to reduce or remove errors**	**How could you prevent this from happening again?**	**How could you prevent this from happening again?**

FIGURE 8 A flow chart for evaluating errors for discussion

An example how to use the error evaluation flow chart is shown in Figure 9.

Evaluate errors		
Step 1: **Identify errors** **Types of error:** Personal Random Systematic Outlier	**Error:** For pH 3, day 10 result is higher than day 15 result (plants can't grow backwards) **Type of error:** Random – subjective reading of height on ruler	**Error:** When water (pH 7) was added, the pH of the soil may have fluctuated over the 20 days. **Type of error:** Random
Step 2: **Explain reasons for errors**	**How did it come about/get introduced?** Could have measured from a different side of the plant, where the soil was higher, so when the ruler was touching the soil it gave a different reading.	**How did it come about/get introduced?** Adding neutral water (pH 7) could cause the soil in the pH 3 and 10 tests to return to a more neutral pH.
Discuss effect of error on quality of experiment/data e.g. Higher or lower result and why	**How did it affect the result?** Higher result at day 10 than at day 15 – and plants cannot grow smaller.	**How did it affect the result?** Unsure because pH of the soil was not tested or controlled
Step 3: **Propose ways to reduce or remove errors**	**How could you prevent this from happening again?** Have a marked spot that the ruler sits at on the soil for every measurement to control where the height is measured from.	**How could you prevent this from happening again?** Test pH every time before and after watering to ensure no fluctuations. Add water of controlled pH to each of the test pots instead of neutral water (pH 7).

FIGURE 9 Evaluating errors using the discussion flow chart for the investigation of soil and plant growth

Modifying the method

Any errors that occurred during your investigation will affect the repeatability, reproducibility and reliability of your experiment. You need to consider your controlled variables during your experiment and when you are evaluating your errors. Figure 11 shows a flow chart that can help you determine modifications you can make to your method.

FIGURE 10 Be sure to keep track of any errors that occur during your experiment.

Modifications to method	
Step 1: **Propose modification** e.g. Something that will improve repeatability, reproducibility and/or reliability of the experiment	**What needs to be controlled?** ↓ **Propose modification/change to a specific step.**
Step 2: **Explain how modification will improve design** **Note:** Be specific (increase/decrease etc.) about the effect on repeatability, reproducibility and/or reliability of the experiment	**How will this make the practical results better?**
Step 3: **Predict effects of modification on quality of data** e.g. Increase accuracy and/or reliability, with example of how	**What will you get for these 'improved' data/results?**

FIGURE 11 Modifications to method flow chart for discussion

FIGURE 12 Think about how you could modify your method to reduce potential errors. For example, repeating experiments and averaging results reduces random error.

An example of how to use the modifications to method flow chart to add to your discussion is given in Figure 13.

Modifications to method		
Step 1: **Propose modification** e.g. Something that will improve repeatability, reproducibility and/or reliability of the experiment	**What needs to be controlled?** More accuracy and precision of measurements in experiment	**What needs to be controlled?** pH of the soil
	Propose modification to a specific step Increase the number of plants tested at each pH	**Propose modification to a specific step** Add a step to the method so that every time watering occurs, pH-adjusted water (3, 7 and 10) is added to each pot. Add step to test pH using pH meter every time that the height is measured.
Step 2: **Explain how modification will improve design** **Note:** Be specific (increase/decrease etc.) about the effect on repeatability, reproducibility and/or reliability of the experiment	**How will this make the practical results better?** This will allow us to understand if the experiment is accurate. With only one plant at each pH, there is no understanding of whether the values were true values or not. Also will allow identification of outliers.	**How will this make the practical results better?** Controlling and testing the pH of the soil will allow the experiment to be completed under the same conditions (repeatability) and under different conditions (reproducibility).
Step 3: **Predict effects of modification on quality of data** e.g. Increased accuracy and/or reliability, with example of how	**What will you get for these 'improved' data/results?** Increased accuracy because measurements will be able to be averaged	**What will you get for these 'improved' data/results?** Increase reliability of results because experimenters will know controlled variables are accounted for

FIGURE 13 An example of the modifications to method flow chart for the investigation regarding soil and plant growth

Conclusion

The conclusion is the end to your poster, report or presentation, a wrap-up of everything you did. You should check the assessment criteria for your investigation to write your conclusion, but conclusions often involve answering the following questions.

- What were your key findings?
- Was your hypothesis supported or disproved?
- Were there any limitations in your experiment?
- How can you overcome those limitations?

Was your hypothesis supported or disproved?

It is important to link all aspects of your poster, report or presentation to the key idea – the question, or hypothesis. In your conclusion, you should first summarise the key findings, then link them straight to the hypothesis and explain if they supported your hypothesis or disproved it.

Limitations

In your conclusion, you need to consider the limitations in your experiment. Check if these are assessed in your conclusion or in your discussion. Limitations are things that affect the overall research and conclusion that you are making. For example, if you were testing the corrosion of metals, and only used zinc to test it, then you cannot conclude that all metals will act this way. The final step in a conclusion is to propose a way that you could overcome these limitations.

18.5B WORKED EXAMPLE

WRITING A CONCLUSION

Write a conclusion for the investigation on the effect of soil pH on soybean growth.

Think	Do
Step 1: Summarise your key findings.	The plant at pH 7 grew the best; the plants at pH 3 and 10 did not grow as tall. pH does affect plant growth.
Step 2: Link your key findings to your hypothesis (has it been supported or disproved).	The initial hypothesis was 'If the pH of the soil is increased or decreased from neutral, then the soybean plant will be shorter'. The hypothesis was supported because findings showed that soybean plants were shorter when soil pH was above and below a neutral pH.
Step 3: Identify any limitations.	One limitation was that only three pH values were tested, which is not a large range. Another limitation was that only soybeans were tested, so data cannot be extrapolated to other plants.
Step 4: Provide recommendations to overcome limitations.	Limitations can be overcome by testing more plants within a smaller range closer to neutral pH. The investigation can also be expanded to include other plants.

TO-DO LIST

- Find the best way to display your data to best represent your results.
- Analyse your results:
  - identify key findings
  - match/support key findings using theory
  - link key findings and theory.
- Evaluate your errors:
  - identify errors
  - explain reasons for errors
  - discuss effects of errors on quality of data
  - propose way to remove or reduce errors.
- Modify the method:
  - propose modifications to the method
  - explain how modifications will improve experiment
  - predict effect of modifications on quality of data.
- Identify whether your hypothesis was supported.
- Identify the limitations of your investigation design.
- Propose recommendations to overcome limitations.

18.6 Communicating your findings

KEY IDEAS

In this topic, you will:

+ select a format for your investigation
+ understand where everything needs to go on your poster
+ create a reference list and acknowledgements.

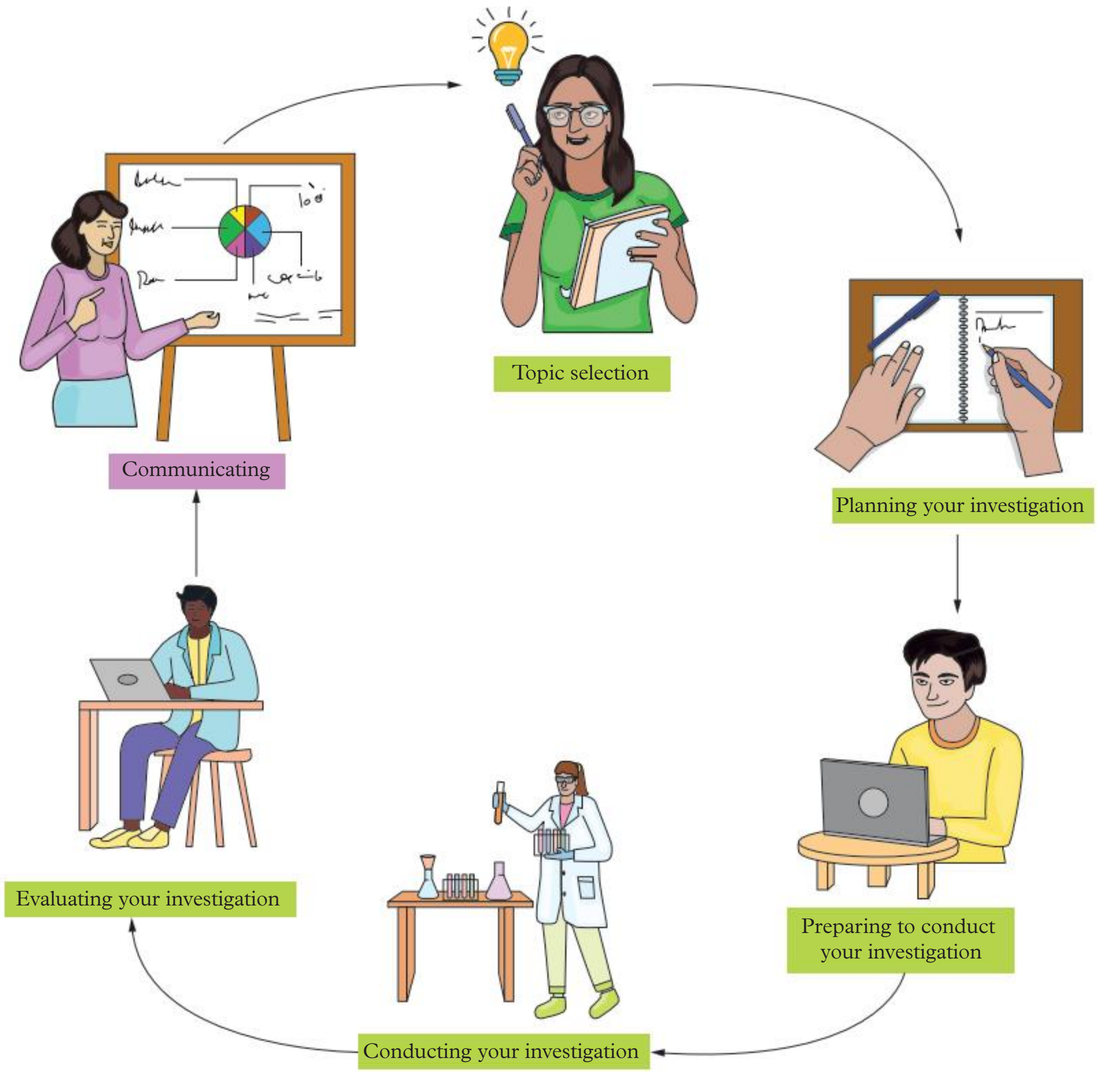


FIGURE 1 The scientific investigation process – this topic is about communicating your findings.

In this topic, we will look at the reporting phase of your scientific investigation. This will include:

- communicating your results as a poster
- using scientific terminology
- referencing
- acknowledgements.

Communicating your results as a poster

There are all different formats you can use to present your investigation. However, when you get to Unit 4 Area of Study 3, you are only allowed to report your findings as a scientific poster so you may wish to choose this format to get some practice in presenting this way. The poster format outlined in the VCAA study design is shown in Figure 3, which includes VCAA's suggestions for what should be under each heading. The poster may be produced electronically or in hard-copy format and must not exceed 600 words.

The 600-word limit does not include supporting text, such as:

- tables
- graphs
- image captions
- references
- acknowledgements.

FIGURE 2 Scientific posters on display at a science conference

Title **Student name**		
Introduction • Brief explanation or reason for undertaking the investigation, including a clear aim, a hypothesis and/or a prediction and relevant background chemical concepts **Methodology and methods** • Brief outline of the selected methodology used to address the investigation question • Summary of data generation method(s) and data analysis method(s) **Results** • Presentation of generated data/evidence in appropriate format to illustrate trends, patterns and/or relationships	**Communication statement reporting the key finding of the investigation as a one-sentence summary**	**Discussion** • Interpretation and evaluation of analysed primary data • Identification of limitations in data and methods, and suggested improvements • Cross-referencing of results to relevant chemical concepts • Linking of results to investigation question and to the aim to explain whether the investigation data and findings support the hypothesis **Conclusion** • Conclusion that provides a response to the investigation question • Identification of the extent to which the analysis has answered the investigation question, with no new information being introduced
References and acknowledgements • Referencing and acknowledgement of all quotations and sourced content relevant to the investigation		

FIGURE 3 The VCAA scientific poster format.

Source: *VCE Chemistry Study Design (2023–2027)* reproduced by permission © VCAA

Within the 600 words, you need to summarise everything that you have done throughout your investigation. This means it is important be careful with how many words you designate to each section.

The following tips can help you generate your scientific poster or report:

- **Choose your content wisely**. Don't include fluff and things that don't relate to your question; find one or two *key* points and stick with those throughout the poster.
- **Don't waste words when there are only a small number of marks allocated**. Check the assessment criteria – if there is only one mark for something, then don't waste 200 words trying to explain it.
- **Proofread your work**. Edit your poster to make sure there is no irrelevant information. After proofreading once, go back and proofread/edit again. Having a friend proofread it for you is also valuable.
- **Make your poster visually appealing**. Figure 4 shows the dos and don'ts of formatting your poster.

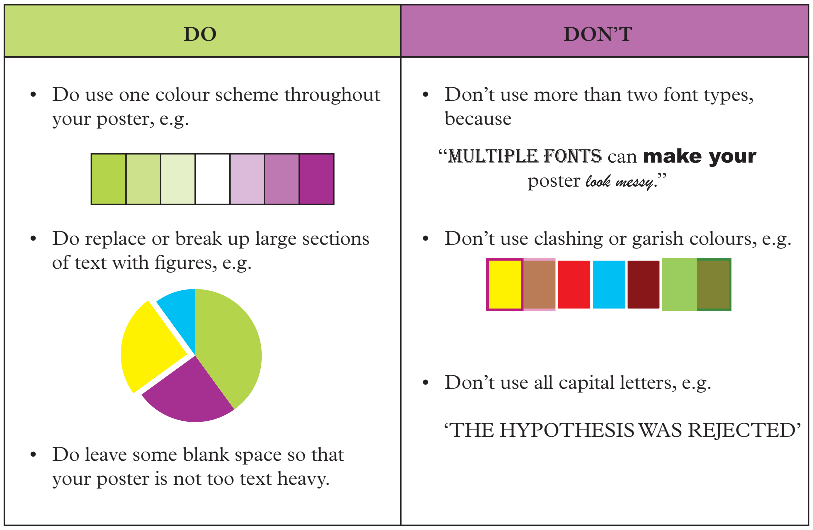


FIGURE 4 A summary of what to do and what not to do to make your poster visually appealing

The communication statement

The centre of the poster will occupy 20–25% of the poster space and will be a one-sentence summary of the major finding of the investigation that answers the investigation question.

Because the communication statement is such a large area of your poster, when writing your statement, you should:

- keep it succinct and use just one sentence
- avoid using long, confusing words
- avoid phrasing it as a question
- make sure it relates to the question.

For example, a communication statement based on the sample investigation relating to soybean growth covered in this chapter might be 'The optimum soil pH for growing soybean plants is pH 7' or 'The tallest soybean plants come from soil with a pH of 7'.

Using scientific terminology

Every scientific discipline uses different terminology and presents ideas in slightly different ways. In chemistry, you need to convey your ideas by using the correct scientific terminology and aligning them with the way chemical ideas are represented. Make sure you use correct symbols, equations and units where possible. Personal pronouns are not used when presenting scientific research, so don't include 'I', 'we' or 'us'. Aim to write your poster in third-person past tense.

Referencing

There are many styles of scientific referencing that can be used. Refer to your assessment criteria to see if a particular type is stated. A common referencing style is APA (American Psychological Association), which you can see in Figure 5. You need to include your reference list at the bottom section of your poster.

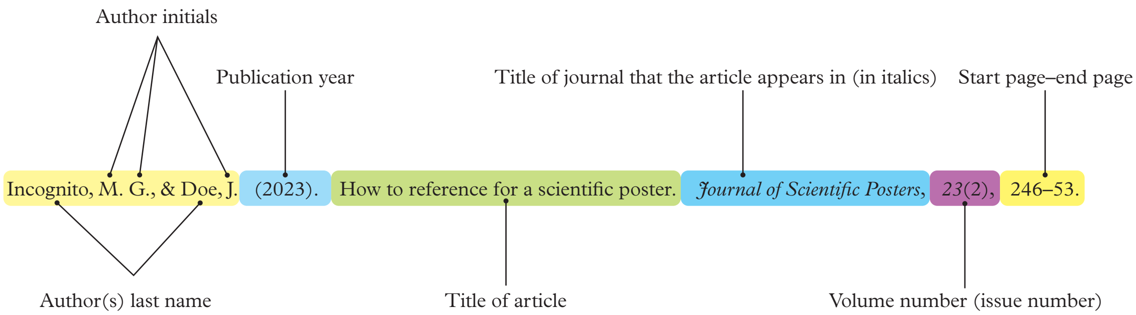


FIGURE 5 An example of how to include a journal article in your reference list using APA referencing

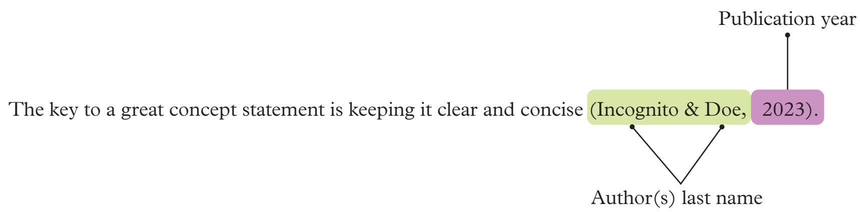


FIGURE 6 An example of how to reference a journal article in the body text of your report

Acknowledgements

The acknowledgements section is where you, the author of the poster, acknowledge and list the people who contributed to or supported your research investigation. People who directly contributed, such as your group members, should appear at the top of the poster as group members and not in the acknowledgements.

People you may wish to acknowledge are:

- the lab assistant who prepared all your equipment and materials
- your teacher for contributing ideas
- class members who helped with ideas
- people who helped proofread your poster.

TO-DO LIST

- Write a communication statement to summarise the results of your investigation.
- Use the format of your choice to communicate your findings in 600 words or fewer.
- Format your list of references.
- Include an acknowledgement for everyone who helped you – don't forget your teacher!

CHAPTER 18 Review

Chapter summary

- As part of Unit 2 Area of Study 3, you will conduct a research investigation related to chemical principles explored in Unit 2 Area of Study 1 and Area of Study 2.
- For your research investigation you will need to collect and analyse primary data; that is, data that has been generated from your own experimentation.
- To answer your investigation question, you can break down the question into smaller parts and generate expanding questions to guide your research.
- Planning your investigation using a research outline is helpful to keeping you on track.
- Annotating your results with detailed notes when collecting data for your investigation can make the research investigation process easier.
- Information can be organised in different ways.
- You must acknowledge your sources by referencing them in the text and in a bibliography.

Research investigation checklist

Use the following checklist to make sure you have completed the research investigation.

I have ...	
written or decided on a scientific question to investigate	☐
defined the key scientific theories and terms that are relevant to my investigation topic	☐
analysed the criteria I will be marked on for this assessment	☐
chosen a methodology for my investigation that will best answer my scientific question	☐
identified the independent and dependent variables that I will be investigating and identified which variables I will control for my investigation	☐
written a testable hypothesis	☐
written a succinct and detailed method for my investigation	☐
conducted a risk assessment	☐
evaluated the ethics of my investigation	☐
set up a logbook to record all results and observations as I conduct the investigation	☐
determined the best way to display my data to best represent my results	☐
analysed my results by identifying key findings and linking key findings to theory	☐
evaluated my errors	☐
modified the method where needed	☐
identified if my hypothesis was supported	☐
identified the limitations of my investigation design and proposed recommendations to overcome limitations	☐
used a poster format to communicate my findings in 600 words or fewer	☐
written a communication statement	☐
formatted my list of references	☐
included acknowledgements	☐

UNIT 2 Review

PART A – Revisit and revise

Part A of the Unit Review will help you revisit and revise all of the key concepts and terms from Unit 2 and test your understanding to identify the strengths and weaknesses in your knowledge.

Unit 2 overview

The chart below shows all of the Areas of Study for Unit 2 and the relevant chapters in your student book. Go to the pages shown to review the key concepts for each chapter.

- **Unit 2**
  - Area of Study 1
    - **Chapter 11** Water as a unique chemical (page 320)
    - **Chapter 12** Acid–base reactions (page 352)
    - **Chapter 13** Redox reactions (page 374)
  - Area of Study 2
    - **Chapter 14** Measuring solubility and concentration (page 400)
    - **Chapter 15** Analysis for acids and bases (page 422)
    - **Chapter 16** Measuring gases (page 446)
    - **Chapter 17** Analysis for salts (page 470)
  - Area of Study 3
    - **Chapter 18** Student-designed investigation (go to your Student obook pro)

Test your understanding

Use the following table to guide your revision:

Step 1 – Read the Key knowledge for this Unit.

Step 2 – Test your understanding of the Key knowledge by answering the question(s). (You can check your answers online).

Step 3 – Rate your understanding of each Key knowledge from low to high.

Step 4 – Use the topic and page numbers to revise the concepts you've identified as needing practice.

Only the first few Key knowledge dot points are shown. Access the rest of the Test your understanding questions in your **obook pro**.

Key knowledge	Test yourself	Rate yourself	Focus your revision
• the existence of water in all three states at Earth's surface, including the distribution and proportion of available drinking water	**1** Identify the feature of water that allows it to exist in all three states at once on Earth's surface.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 11.1 Pages 310–311
• explanation of the anomalous properties of H_2O (ice and water), with reference to hydrogen bonding – trends in the boiling points of group 16 hydrides – the density of solid ice compared to liquid water at low temperature – specific heat capacity of water, including units and symbols	**2** Analyse the following statements and determine if they are true or false. Justify why the statements you have selected are true, and if they are false, correct the statements to make them true. **a** Ice floats on water because it has a higher density than water. **b** H_2O has a higher boiling point than H_2S because H_2O has covalent bonding within the molecules and H_2S doesn't. **c** H_2O has a higher boiling point than the other group 16 hydrides because H_2O can form hydrogen bonds. **d** The specific heat capacity of water is high because of the hydrogen bonding between the water molecules.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 11.2 Pages 312–315
• the relatively high latent heat of vaporisation of water and its impact on the regulation of the temperature of the oceans and aquatic life	**3** Draw a heating curve for water. **a** Label: **i** the temperatures for melting ice and boiling water **ii** L_{fus} **iii** L_{vap} **b** Describe the behaviour of the water molecules at each step on the heating curve.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 11.3 Pages 316–319
	4 Explain how the high L_{vap} of water is a benefit for humans, animals and aquatic life.	☐ **High** – I've got this! ☐ **Medium** – I could use a bit more practice. ☐ **Low** – I have some work to do!	Topic 11.3 Pages 316–319

PART B – Exam essentials

Now that you've completed your revision for Unit 2, it's time to learn and practise some of the skills you'll need to answer exam questions like a pro! To help you, our expert authors have created the following advice and tips so you can maximise your results on the end-of-year examination.

Exam tip 1: Check your significant figures

- Sometimes marks are awarded for the correct number of significant figures (even if it's not specified!), so you will need to be aware of them.
- Don't round off values too early or by too much – this can cause an error in your calculations.
- Double-check that you have given your final answer in the correct number of significant figures. Go back to Chapter 1 Science toolkit to refresh your understanding.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 4
A mixture contains several different organic liquids all of which boil at temperatures greater than 50.0°C. The compounds present in the mixture are separated. Three of the compounds, compounds X, Y and Z, are analysed as follows.

Compound X is vaporised. At a temperature of 120°C and a pressure of 115 kPa, a 0.376 g sample of the vapour occupies 124 mL.

a. Calculate the molar mass, in g mol^{-1}, of compound X.

2 marks

Source: *VCE 2008 Chemistry Exam 1* reproduced by permission © VCAA

A high-scoring response

Student looked and considered all the significant figures in the question and answered with the same number

A mixture contains several different organic liquids all of which boil at temperatures greater than 50.0°C. The compounds present in the mixture are separated. Three of the compounds, compounds X, Y and Z, are analysed as follows. 3 sig figs
Compound X is vaporised. At a temperature of 120°C and a pressure of 115 kPa, a 0.376 g sample of the vapour occupies 124 mL.
a. Calculate the molar mass, in g mol^{-1}, of compound X.

m
n | M

$PV = nRT$

$P = 115$ kPa

$V = \frac{124}{1000} = 0.124$ L

$n = ?$

$T = 120 + 273 = 393$ K

$n = \frac{115 \times 0.124}{393 \times 8.31}$

$n = 0.00437$ mol

$M = \frac{m}{n}$

$M = \frac{0.376}{0.00437}$

$M = 86.0$ g mol^{-1}

2 marks

Student defined all the variables

Correct answer to 3 sig figs

A low-scoring response

A mixture contains several different organic liquids all of which boil at temperatures greater than 50.0°C. The compounds present in the mixture are separated. Three of the compounds, compounds X, Y and Z, are analysed as follows.

Compound X is vaporised. At a temperature of 120°C and a pressure of 115 kPa, a 0.376 g sample of the vapour occupies 124 mL.

a. Calculate the molar mass, in g mol^{-1}, of compound X.

$PV = nRT$ $\qquad n = 0.004$ mol

Do not round too soon. If you do, you will get the wrong answer, so it is best to keep the numbers in your scientific calculator and work through with them.

$n = \frac{PV}{RT} = \frac{115 \times 0.124}{8.31 \times (120 + 273)} \qquad \frac{0.376}{0.004}$

$= 94$ g mol^{-1}

2 marks

Because the student rounded too soon, they got a higher answer and wrote it exactly – they forgot about the 3 sig figs required.

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam (below and on the next page). Imagine you are an examiner and use the marking guide on the next page to mark the response.

Question 8 (12 marks)

The conversion of sulfur dioxide to sulfuric acid is used in a number of analytical techniques to determine the amount of analyte present in a substance. The half-equation for this reaction is

$$SO_2(aq) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$$

a. What type of reaction is this? 1 mark

Redox

b. Sulfur dioxide is often used as a preservative in food and drink. The sulfur dioxide content in dried apricots was determined by gravimetric analysis as follows:

- The dried apricots were powderised in a blender.
- A sample of the apricot powder weighing 50.00 g was put into a conical flask containing 100 mL of de-ionised water.
- A 3% solution of hydrogen peroxide was added to convert the dissolved sulfur dioxide to sulfate ions.
- An excess of barium chloride solution was then added. The barium sulfate precipitate was filtered off, dried and weighed to constant mass.

The equation for the precipitation of barium sulfate is

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The following results were recorded.

mass of dry filter paper	0.864 g
mass of dry filter paper and $BaSO_4$ sample	1.338 g

$M(BaSO_4) = 233.4\ g\ mol^{-1}$ $M(SO_2) = 64.1\ g\ mol^{-1}$

i. Determine the percentage, by mass, of SO_2 in the apricot sample. 4 marks

$n(BaSO_4) = n(Ba^+)$

$n(BaSO_4) = 1.338 - 0.864$

$= \frac{0.474}{233.4}$

$n(BaSO_4) = 0.002031$ mol

$n(Ba^+) = 0.002031$ mol $\times 64.1$

$m(Ba^+) = \frac{0.13}{50} \times 100$

$= 0.26\%$

Source: *VCE 2014 Chemistry Exam 1* reproduced by permission © VCAA

Marking guide

Question 8bi	1 mark for correctly identifying the equation type
	1 mark for accurately calculating $m(BaSO_4)$
	1 mark for accurately calculating $m(SO_2)$
	1 mark for accurately calculating $\%(SO_2)$
	1 mark for expressing answer to correct number of significant figures

Exam tip 2: Be careful of units

- In exam questions, you may be asked to give your answer in specific units, e.g. 'Find the volume in mL' or 'Find the mass in mg'.
- Make sure you highlight or make note of these in the question before you start.
- Sometimes a question will give you one set of units, such as mg, but to calculate the number of moles, it needs to be in grams. Make sure you remember which units are needed to calculate different values. For example:
  - For $PV = nRT$, P is in kPa, V is in L, n is in mol and T is in K
  - For $n = CV$, n is in mol, C is in M (or mol L^{-1}) and V is in L
  - For $n = \frac{m}{M}$, n is in mol, m is in g and M is in g mol^{-1}

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses.

Question 7

Sulfur dioxide gas is commonly used as a preservative in wine. An important source of SO_2 is solid sodium metabisulfite ($Na_2S_2O_5$; molar mass 190 g mol^{-1}). $Na_2S_2O_5$ reacts readily with acid as follows.

$$Na_2S_2O_5(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + 2SO_2(g)$$

3 marks

Source: *VCE 2007 Chemistry Exam 1* reproduced by permission © VCAA

A high-scoring response

Highlighted the variables + noted units

a. Calculate the volume, in litres, of SO_2 produced at 1.00 atm pressure and 15.0°C when 250 g of $Na_2S_2O_5$ reacts with excess acid.

kPa K

$\frac{1}{0.987} \times 100 = 101.32$ kPa

$15 + 273 = 288$ K

$1\,n(Na_2S_2O_5) = 2\,n(SO_2)$

$n(Na_2S_2O_5) = \frac{250}{190} = 1.316$ mol $\times 2$

$n(SO_2) = 2.632$ mol

$PV = nRT$

$101.32 \times V = 2.632 \times 8.31 \times 288$

$V = 62.2$ L

Identified the mol ratio

Completed the conversions first, knew what they were looking for

Correct sig figs + answer

3 marks

A low-scoring response

a. Calculate the volume, in litres, of SO_2 produced at 1.00 atm pressure and 15.0°C when 250 g of $Na_2S_2O_5$ reacts with excess acid.

$n = \frac{250}{190}$

$PV = nRT$

$1 \times V = 1.316 \times 8.31 \times 15$

$= 164$ L

$n = 1.316$ mol

Forgot the mol ratio $Na_2S_2O_5$:$2SO_2$

3 marks

Didn't convert the units

Think like an examiner

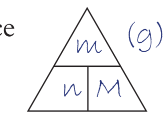
To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guide below to mark the response.

Question 4

a. In order to help prevent tooth decay, fluoride ions at a level of 0.9 mg L^{-1} of F^- are added to Melbourne's public water supplies. The fluoride ions are obtained by adding sodium fluoride (NaF) to the water.

i. Calculate the mass of sodium fluoride in mg that must be present in one litre of water to produce a concentration of fluoride ions of 0.90 mg L^{-1}.



$F^- = \frac{0.9 \text{ mg}}{1 \text{ L}} = n(F^-) = n(NaF) = 1:1$

Mass in g $= \frac{0.9}{1000}$

$= 0.0009 \text{ g} \times 42$

$= 0.00199 \text{ g}$

2 marks

Source: *VCE 2004 Chemistry Exam 1* reproduced by permission © VCAA

Marking guide

Question 1a	1 mark for correct conversion of mg into g
	1 mark for multiplying by the molar mass (42.0) and the correct answer in mg

Exam tip 3: Show your working

- Showing your working is important. If you get something wrong early on in a question, examiners are able to follow your working and can award you consequential marks. If you don't show your working, you won't get the marks at all.
- If you don't finish the question, the examiners may award marks for what you have shown you can do.

See it in action

Read the real exam question below and see how the tip has made a difference in the high-scoring and low-scoring responses on the next page.

Question 3

A student is to accurately determine the concentration of a solution of sodium hydrogencarbonate in a titration against a standard solution of hydrochloric acid, HCl.

The first step in this experiment is to accurately dilute 100.0 mL of a 1.00 M HCl stock solution to a 0.100 M solution using a 1.00 L volumetric flask.

However, instead of using distilled water in the dilution, the student mistakenly adds 900.0 mL of 0.0222 M sodium hydroxide, NaOH, solution.

Source: *VCE 2009 Chemistry Exam 1* reproduced by permission © VCAA

A high-scoring response

b. Calculate the concentration of the hydrochloric acid in the 1.00 L volumetric flask after the student added the sodium hydroxide solution.
Give your answer to correct significant figures. =3

$n(NaOH) = \frac{900}{1000} \times 0.0222$ Original $n(HCl) = \frac{1\ M}{100\ ml} = 0.1$ mol

$n(NaOH) = 0.01998$ mol $n(HCl)$ remaining $= 0.1 - 0.01998$

$n(NaOH) = n(HCl) = 1:1$ $= 0.0800$ mol

$\therefore n(HCl) = 0.01998$ mol $c(HCl) = \frac{0.0800}{1\ L}$

Student showed clear and ALL working for the question, noted that the sig figs were important and showed them correctly

$= \underline{0.0800\ M}$

Correct answer

2 marks

A low-scoring response

b. Calculate the concentration of the hydrochloric acid in the 1.00 L volumetric flask after the student added the sodium hydroxide solution. Give your answer to correct significant figures.

$n = cV$

$n(NaOH) = \frac{900}{1000} \times 0.0222$ mol

$= 0.01998$ mol

Mol ratio $= 1:1$

Student showed working and got the first part correct.

0.1998 mol of HCl $\times$ 1 L

$= \underline{0.020\ M}$ of HCl

If student had written this with no working, they would not have received any marks. Because they showed how they calculated n(NaOH) correctly, with working, and they knew the mol ratio, they will get 1 mark for this question.

2 marks

Think like an examiner

To maximise your marks on an exam, it can help to think like an examiner. Consider how many marks each question is worth and what information the examiner is looking for.

A student has given the following response in a practice exam. Imagine you are an examiner and use the marking guide below and on the next page to mark the response.

Question 6

Some students were set the task of determining the concentration of acetic acid in a particular brand of vinegar. An outline of the method they used is given below.

1. A burette is filled with a standard solution of sodium hydroxide.
2. The vinegar is diluted by a factor of 10 in a volumetric flask. A pipette is used to transfer 20.00 mL of diluted vinegar to a conical flask and a few drops of phenolphthalein indicator is added.
3. The diluted vinegar is titrated with the base. Titrations are repeated until three concordant results are obtained.

The equation for the reaction is

$$NaOH(aq) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

d. One student's results are given below. The data shown in the student's laboratory book was

concentration of NaOH(aq)	=	0.11 M
volume of undiluted vinegar	=	10.00 mL
total volume of diluted vinegar	=	100.00 mL
volume of diluted vinegar used in each titration	=	20.00 mL
average titre of NaOH	=	15.35 mL

Based on these results, calculate the concentration, in mol L^{-1}, of acetic acid in the undiluted vinegar solution. Be careful to use the correct number of significant figures in your answer.

$n(NaOH) : n(HCl) = 1:1$ $\quad n(HCl) = \frac{0.00717 \text{ mol}}{20.00}$

$n(NaOH) = \frac{0.11}{15.35}$

$c = 0.00036 \text{ M (diluted)} \times 10 \text{ (DF)}$

$n(NaOH) = 0.00717 \text{ mol}$ $\quad c = 0.0036 \text{ mol } L^{-1}$

4 marks

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Marking guide

Question 6d	1 mark for correct concentration of NaOH
	1 mark for correctly calculating the concentration of acetic acid in the diluted solution
	1 mark for correctly calculating the concentration of acetic acid in the undiluted solution
	1 mark for correct number of significant figures

Practice makes perfect

Now that you know all these tips, it's time for you to move on to Part C – Exam practice to put them into practice.

PART C – Exam practice

Multiple choice

Question 1

Which of the following cannot be explained by the existence of hydrogen bonding in water?

A The higher melting point of water relative to other molecules of about the same size

B The ability of water to dissolve other substances

C The expansion of water upon freezing

D The bent shape of the water molecule

Question 2

Find the pH of 0.0050 M calcium hydroxide solution that has completely dissociated in water.

A 12.0

B 11.7

C 2.3

D 2.0

Question 3

Identify the reaction in which water is acting as an acid.

A $H_3O^+(aq) + HPO_4^{2-}(aq) \rightarrow H_2O(l) + H_2PO_4^-(aq)$

B $H_2O(l) + HCO_3^-(aq) \rightarrow H_3O^+(aq) + CO_3^{2-}(aq)$

C $H_2O(l) + NH_3(g) \rightarrow NH_4^+(aq) + OH^-(aq)$

D $H_3O^+(aq) + HS^-(aq) \rightarrow H_2S(aq) + H_2O(l)$

Question 4

For the following equation, identify the reducing agent and the oxidising agent:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

A Reducing: O_2. Oxidising: NH_3

B Reducing: H_2O. Oxidising: NO

C Reducing: NH_3. Oxidising: O_2

D Reducing: NO. Oxidising: H_2O

Question 5

A clean iron nail is dipped into a silver nitrate solution. Predict the reaction that would be observed.

A No reaction will occur.

B The iron will be reduced.

C The iron will become silver plated.

D Oxygen will form around the iron nail.

Question 6

A 150 mL solution of sodium chloride was made by using 2.00 g of salt. Which of the following is not a concentration of the solution?

A 0.228 M

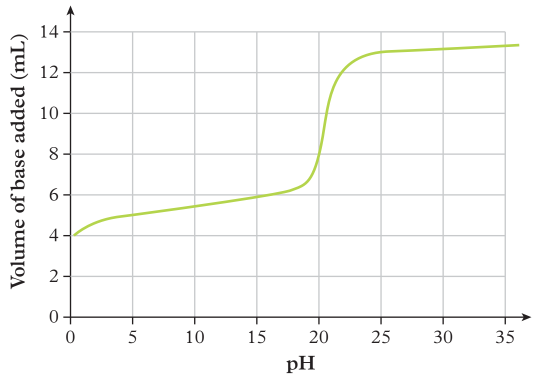
B 13.3 g L^{-1}

C 1.33×10^4 ppm

D 13.3% (m/v)

Question 7

The pH curve for a titration between a 0.10 M solution of a strong base and 20 mL of a solution of 0.10 M weak acid is shown below.



Identify the indicator that would be the *least* suitable to detect the endpoint of the titration.

A Bromophenol blue

B Bromothymol blue

C Phenolphthalein

D Phenol red

Question 8

A titration between 100 mL of hydrobromic acid (HBr) and a 1.5 M NaOH solution reached an end point after the addition of 24.0 mL of the hydroxide titrant. The concentration of the acid solution was:

A 0.36 M

B 0.72 M

C 1.4 M

D 3.1 M

Question 9

What is the maximum volume of carbon dioxide gas, measured at SLC, that could be produced by the complete combustion of 65.0 g of octane?

A 113.1 L

B 14.13 L

C 56.50 L

D 176.7 L

Question 10

A hydrated salt of mass 26.30 g was heated and dried until a final mass of 14.67 g was maintained. Identify the hydrated salt.

A $CaCO_3 \cdot 6H_2O$

B $Na_2CO_3 \cdot 6H_2O$

C $Mg(NO_3)_2 \cdot 6H_2O$

D $CaSO_4 \cdot 6H_2O$

Short answer

Question 1 (8 marks)

Consider the chemical equation:

$HF(aq) + NO_3^-(aq) \rightarrow HNO_3(aq) + F^-(aq)$

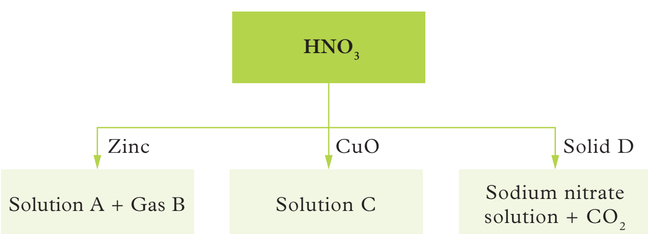
a Give the definition of a strong acid. (1 mark)

b Identify one base from the equation. (1 mark)

c State the conjugate acid for the base in part **b**. (1 mark)

d 0.1 M solutions of both HF and HNO_3 solutions are prepared. Determine if the pH of the solutions will be the same or different from each other. Justify your answer. (2 marks)

e A number of chemical reactions with dilute HNO_3 are shown below.



Identify the following.

i Solution A

ii Solution C

iii Solid D (3 marks)

Question 2 (8 marks)

The Statue of Liberty, in New York Harbor, was originally made from thin sheets of copper bolted onto a steel (iron) frame.

The exposed copper sheets eroded to a green patina with little deterioration of the copper sheets. However, the steel frame that was in contact with the copper sheets underwent extensive deterioration and has now been replaced with stainless steel (an alloy that includes nickel and chromium).

For the reaction between the original steel (iron) frame and the copper sheets:

a construct a balanced oxidation half-equation for the reaction. (1 mark)

b construct a balanced reduction half-equation for the reaction. (1 mark)

c construct a balanced overall equation for the reaction. (1 mark)

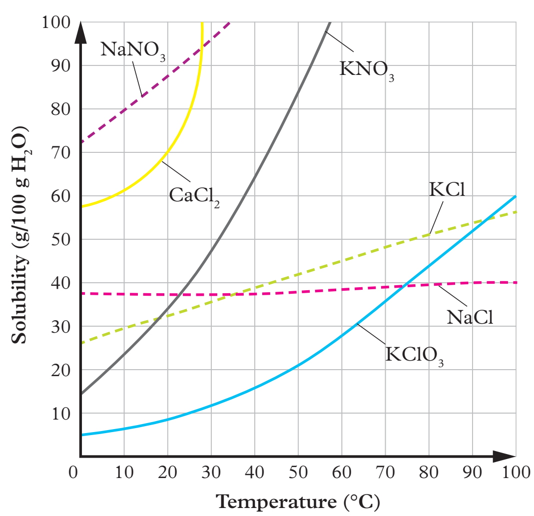
d analyse both metals to explain why the steel corroded more rapidly when in contact with the copper sheets. (1 mark)

e Draw a cell diagram that could be used to measure the potential difference for the reaction between copper and steel (iron). Label the:

- anode and cathode
- positive and negative electrodes (and suitable materials for each)
- salt bridge and the direction of movement of each of the ions in the salt bridge
- direction of flow of electrons in the external circuit. (4 marks)

Question 3 (5 marks)

Use the solubility curve to answer the questions.



a What is the maximum amount of the following salts that can be dissolved in 25 g of water at 20°C?

  i Calcium chloride

  ii Potassium nitrate (2 marks)

b A solution containing 40 g of potassium chlorate is dissolved in 100 g of water at 90°C.

  i Is this solution saturated or unsaturated? Justify your answer. (2 marks)

  ii If the potassium chlorate solution was cooled to 20°C, calculate the mass of potassium chlorate that will crystallise out. (1 mark)

Question 4 (7 marks)

A sample of vinegar was titrated with sodium hydroxide:

$$CH_3COOH(aq) + NaOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l)$$

A 10.0 mL sample of the vinegar was diluted to 100.0 mL with distilled water and 10.0 mL aliquots of the diluted solution were titrated with 0.250 M NaOH, using a suitable indicator. The following results were obtained.

Titration	Titre value (mL)
Trial	31.64
1	31.44
2	31.68
3	31.70

a Select a suitable indicator for the titration and justify your selection. (2 marks)

b Calculate the concentration of ethanoic acid in the original sample. (3 marks)

c Before the titration, all the glassware was rinsed with deionised water. Discuss the effect that rinsing all the glassware with deionised water will have on the final result. (2 marks)

Question 5 (6 marks)

A 6.50 L sealed container has 7.80 g of an unknown small alkane in it, at SLC.

a Determine the identity of the unknown alkane (2 marks)

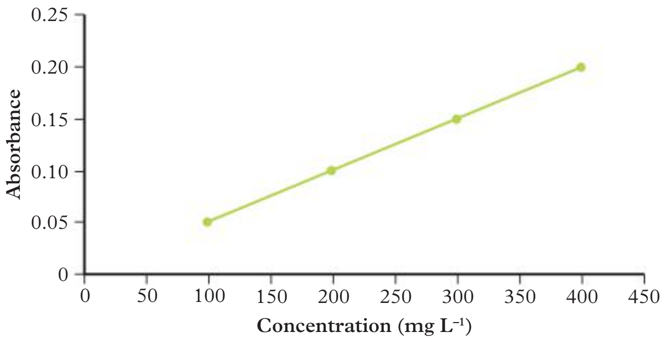
b The container is heated to 45.0°C. Calculate the pressure of the container after it is heated. (1 mark)

c Write an equation for the combustion of the alkane that you determined is in the container. (1 mark)

d If the full amount of the alkane was fully combusted at SLC, calculate the amount of CO_2 that would be produced by this reaction. (2 marks)

Question 6 (6 marks)

The concentration of sodium ions was analysed in two different water samples by UV–visible spectroscopy. A set of four known concentrations was analysed and used to create a calibration curve, shown below.



The absorbance of the first water sample was 0.135.

a Use the calibration curve to determine the concentration of sodium ions in the first water sample. (1 mark)

b The solutions were diluted before they were analysed. A 20 mL sample of the original solution was diluted into a 100 mL conical flask.

  i Calculate the dilution factor for the dilutions. (1 mark)

  ii Calculate the concentrations of the first undiluted solution in g L^{-1}. (2 marks)

c The second absorbance was 0.21. Explain what you would need to do in order to find the concentration for this solution. (2 marks)

TOTAL MARKS

______ /50 marks

You can find the following resources for this section in your obook pro:

Unit 2 Review Part A
Test your understanding questions

Unit 2 Review Part C
Exam practice questions

Weblink
Past examinations and examiners' reports

Unit 1 & 2 Exam practice
Now that you've completed Units 1 & 2, put your knowledge and skills to the test with these extra practice exam questions.

CHAPTER 19

Practical work

To complete VCE Chemistry, you will need to complete at least 10 hours of practical work for each of Units 1 and 2. Practical work can cover a range of scientific investigation methodologies, such as controlled experiments, modelling, case studies, classification and identification, literature reviews, fieldwork, simulations, and product, process or system development. All investigations that are undertaken as part of your course, as well as internal assessment tasks, should be written in a logbook that will be monitored and submitted to your teacher. Before undertaking an investigation for the first time, you should consider any ethical concerns, including the importance of sociocultural, economic, political and legal factors that may arise from science-related decision making.

SAFETY IN THE LABORATORY

This chapter will highlight key safety concerns for each practical, although there are some general safety concerns to be considered before completing all practical work.

- → Tie long hair back.
- → Do not eat or drink in the lab.
- → Always be aware of your peers and act in a way that will not cause harm.
- → Wear a lab coat, safety glasses, closed-toe shoes and gloves.
- → Review the school's safety procedures and the locations of the eyewash, shower, spill kits and first aid kits.
- → Handle all chemicals with care and consult your teacher and risk assessments for all hazards involved with each chemical.
- → Keep open flames away from flammable materials.
- → Handle hot material with the appropriate equipment (i.e. heat-resistant gloves or tongs).
- → Before use, always check that electrical equipment is not damaged and that there are no exposed wires.
- → Conduct fieldwork in groups, and complete a full risk assessment before any excursion.

It is the responsibility of the teacher and school to conduct a risk assessment before any practical covered in this book.

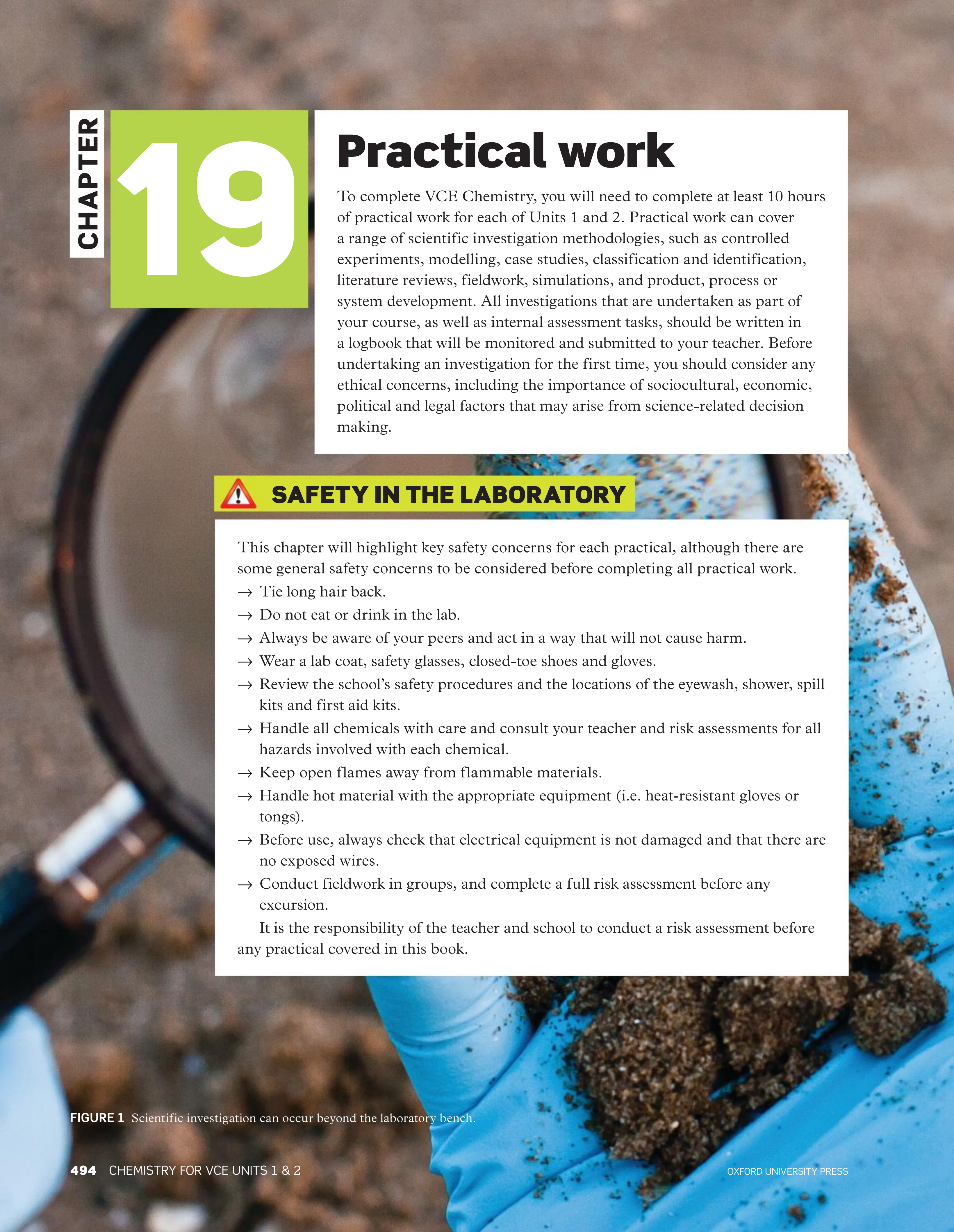


FIGURE 1 Scientific investigation can occur beyond the laboratory bench.

UNIT 1 PRACTICALS

UNIT 2 PRACTICALS

2.3

PRACTICAL:
LITERATURE REVIEW

Are we running out of helium?

Practical worksheet
2.3 Are we running out of helium?

Context

Helium is required worldwide for medical equipment, scientific research and party balloons. Supplies of helium are depleting fast. Some estimates predict that helium could become virtually unobtainable by the middle of the 22nd century. Once helium supplies have been exhausted, there is no way of making more.

In this practical, you will research and review how helium is currently used and how it might be used in the future.

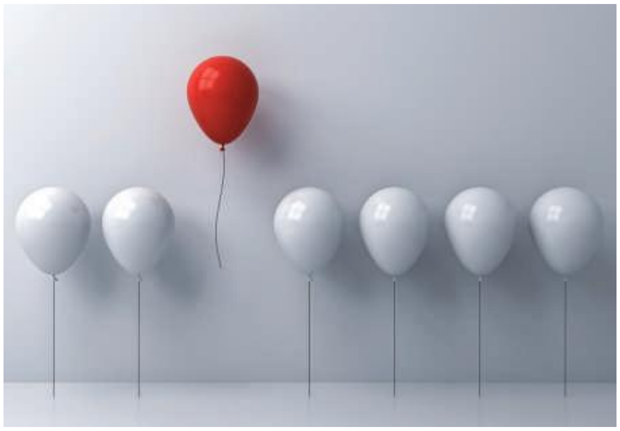


FIGURE 1 Helium is a critical element.

Aim

To investigate the current uses of helium and identify strategies for using helium more sustainably

Instructions

1. Look for secondary sources of information. These can include the internet, books, scientific magazines, videos, podcasts, or interviewing an expert.
2. Use the CRAAP method to evaluate the reliability of your sources. This method is outlined in Topic 10.4, Step 5.
3. Make notes about what you have learnt. You can organise your information in different ways. Examples can be found in Topic 10.5, Step 6.
4. Prepare an oral presentation as a 2–3-minute video that answers the following questions:
   - **a** What are the origins, structure and abundance of helium on Earth?
   - **b** Where is helium found today on Earth, and how much does it cost?
   - **c** What are the current uses of helium? Give statistics where possible.
   - **d** Can we replace helium in each of the uses outlined in part **c**?
   - **e** How can we use helium more sustainably?

   Tips for preparing a presentation can be found in Topic 10.6, Step 8.
5. Make sure you record the details of all of the sources you use, including the title of the source, who it was written by, when it was written, page numbers or URLs and the date you accessed it. You will need this information to put together a bibliography (Topic 10.7, Step 9).

3.2A PRACTICAL: MODELLING

How can modelling be used to understand molecular shapes?

Video demonstration
3.2A How can modelling be used to understand molecular shapes?

Practical worksheet
3.2A How can modelling be used to understand molecular shapes?

Risk assessment
3.2A How can modelling be used to understand molecular shapes?

Context

A scientific model is a way of representing an idea or concept that helps you to understand it.

In this practical, you will use modelling to understand the 3D shape of molecules. You will also discuss the limitations of modelling.

Aim

To model covalently bonded molecules and determine, understand and visualise their 3D shape, by creating your own molecular modelling kit

Materials

- Plasticine
- Cotton buds
- Aluminium foil
- Playdough (different colours)
- Beads (different colours)
- Toothpicks
- Straws
- Cardboard strips
- Pipecleaners

Method

1 Identify the materials that you will use in your modelling kit to represent the atoms and bonds.

2 Assign colour codes to your atoms so you remember which is which.

3 Determine the molecular shape of the following molecules by making 3D models of them.

a CH_4	**b** CO_2	**c** H_2O
d NH_3	**e** HF	**f** CCl_4
g C_2H_4	**h** CH_3Cl	**i** CH_3NH_2
j O_2	**k** HCN	

Questions

1 In your logbook, represent the 3D models you have made as drawings. Keep them as 3D as you can.

2 For each molecule:

  a show the molecular formula, structural formula and Lewis structure

  b identify the molecular shape.

3 Design a set of instructions for another Year 11 student who wants to use your modelling kit.

4 Swap modelling kits and instructions with a partner. Follow their instructions and use their modelling kit to create the same molecules as in step 3 of the Method. Once you are finished, check with your partner to see if you have used their kit correctly.

5 Explain why scientific models are useful and identify when you would use them.

6 Identify the possible limitations in your 3D modelling. Discuss how you could overcome these limitations.

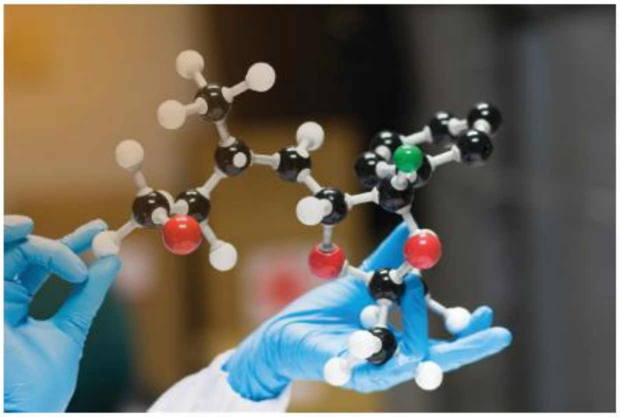


FIGURE 1 A molecular model used to visualise molecular shape

3.2B
PRACTICAL: MODELLING

How can balloons help us understand VSEPR theory?

Video demonstration
3.2B How can balloons help us understand VSEPR theory?

Practical worksheet
3.2B How can balloons help us understand VSEPR theory?

Risk assessment
3.2B How can balloons help us understand VSEPR theory?

Context

Valence shell electron pair repulsion (VSEPR) theory can be challenging to understand with 2D models alone. In Practical 3.2A, you built your own molecular modelling kit which helped you see molecules in 3D. However, it can be hard to predict the exact molecular shape and position of electrons by using a kit.

In this practical, you will use balloons (representing electron pairs) to understand VSEPR theory.

Aim

To model different molecular shapes using balloons

Materials

- Balloons
- Paperclips, bulldog clips or pegs
- Markers

Method

Part A – Linear molecules

1 Inflate a balloon and tie it.
2 Inflate a second balloon to the same size as the first one, and tie it.
3 Tie the balloons together or attach them with a clip or peg.

Part B – Bent molecules

1 Inflate and tie three balloons.
2 Tie the balloons together or attach them with a clip or peg.

Part C – Pyramidal molecules

1 Inflate and tie four baloons.
2 Tie the balloons together or attach them with a clip or peg.
3 On one balloon, draw a pair of dots.

Part D – Tetrahedral molecules

1 Inflate and tie four balloons.
2 Tie the balloons together or attach them with a clip or peg.

Questions

1 In your logbook, represent the 3D models you have made as drawings. Keep them as 3D as you can.
2 Describe the bond angles in each balloon model. Refer to VSEPR theory in your answer.
3 Compare the models from parts C and D. What do the dots on the balloon represent in part C?
4 For each balloon model, identify one example of a covalent molecule that you would expect to exist in that arrangement.
5 Explain why the balloons need to be the same size.
6 Identify one possible limitation of your balloon models.

FIGURE 1 In this practical, balloons represent electron pairs.

3.5 PRACTICAL: CONTROLLED EXPERIMENT

How do intermolecular forces influence the physical properties of covalent molecules?

Video demonstration
3.5 How do intermolecular forces influence the physical properties of covalent molecules?

Practical worksheet
3.5 How do intermolecular forces influence the physical properties of covalent molecules?

Risk assessment
3.5 How do intermolecular forces influence the physical properties of covalent molecules?

Context

The properties of covalently bonded substances can differ because of the intermolecular forces present between molecules.

In this practical, you will conduct a controlled experiment to collect primary data and observations. This will help you understand how different intermolecular forces affect the following properties:

- polarity – the separation of electric charge in a molecule
- rate of evaporation – how quickly a substance changes from liquid to gas
- capillary action – the ability of a liquid to move through a narrow space without any help, usually against gravity – related to how adhesive ('stickiness' to unlike molecules) or cohesive ('stickiness' to like molecules) a molecule is
- surface tension – the attraction that molecules have for each other that helps to minimise surface area (the area of the exposed surface).

Aim

To compare the strengths of intermolecular forces by investigating the physical properties of polarity

Materials

- Water
- Hexane
- Ethanol
- Glycerol
- Acetone
- 4 burettes (per group)
- Small coins (5 cent or 10 cent)
- Pasteur pipette
- Capillary tubes
- Cotton buds
- Beakers
- Balloons
- Dark coloured tile (or plastic sheet)
- Stopwatch
- Ruler
- Measuring cylinders
- Retort stand and clamp
- PPE (lab coat, safety glasses)

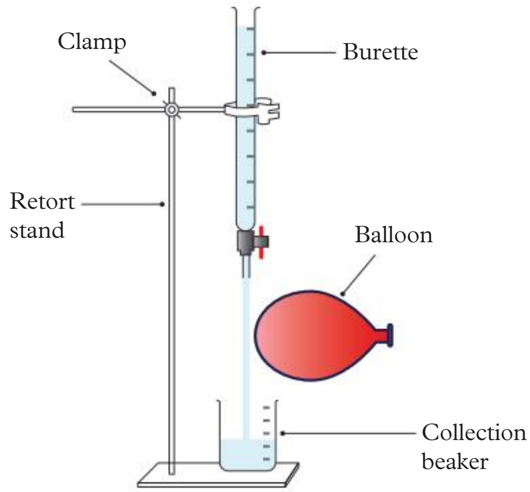


FIGURE 1 How to set up the burette to test polarity

Method

Part A – Polarity

1. Set up four burettes with an empty beaker underneath each (Figure 1).
2. Fill each burette with a different solvent: water, hexane, ethanol and acetone.
3. Rub a balloon against your hair or shirt to build up static electricity.
4. Open the tap of the burette and let a stream of solvent flow. Bring the balloon close to the stream.

5 Record your observations.
6 Repeat steps 3–5 for the three other solvents in the burettes.

Part B – Rate of evaporation

1 Dip a cotton bud into a beaker of water.
2 Use the cotton bud to create a 5 cm streak across the dark tile or plastic sheet.
3 Start the timer.
4 Watch until all of the water has evaporated and stop the timer when it has all gone. Record the time.
5 Repeat steps 1–4 for hexane, ethanol, glycerol and acetone.

Part C – Capillary action

1 Measure 3 mL each of water, hexane, ethanol into separate small beakers.
2 Put a capillary tube into each beaker and record the heights of the liquids.

Part D – Surface tension

1 Use a Pasteur pipette to place one drop of water at a time onto the head side of a coin, until the water spills over the edge.
2 Count the number of drops and record your results.
3 Repeat steps 1 and 2 for hexane, ethanol, glycerol and acetone.

Results

Download the practical worksheet in your obook pro and fill in the Results table.

Discussion

1 Complete the molecular formula, structural formula, polar or non-polar, and intermolecular forces rows for each solvent in the Results table.

Part A – Polarity

2 Explain any differences in how each solvent reacted to the charged balloon. Refer to the properties of each substance in your response.

Part B – Rate of evaporation

3 Arrange the solvents in order from fastest to slowest rate of evaporation.
4 Contrast the rates of evaporation. Refer to the structure of each molecule and intermolecular forces present in your response.

Part C – Capillary action

5 Arrange the solvents in order from smallest to greatest height in the tubes.
6 Given that glass is highly polar, explain the differences in height.

Part D – Surface tension

7 A larger number of solvent drops that sit on a coin before spilling indicates higher surface tension. Arrange the solvents in order from lowest to highest surface tension.
8 Contrast the surface tension of each solvent. Refer to the intermolecular forces between the molecules in your response.
9 You will have identified hydrogen bonding as an intermolecular force of some of the solvents. Compare the strengths of hydrogen bonding between these molecules.
10 Compare the strength of the intermolecular forces in hexane, ethanol and acetone. How do the different intermolecular forces influence the properties of each?

Inquiry: What if chain length of the non-polar molecules was increased?

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify the *dependent* variable that you will measure and/or observe.
4 Identify two variables that you will need to control to ensure a fair test. Describe how you will control these variables.
5 Write down the method you will use to complete your investigation.
6 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.

Conclusion

Describe how the strengths of different intermolecular forces between covalent molecules can affect their physical properties.

4.2

PRACTICAL: CONTROLLED EXPERIMENT

How can we determine the reactivity series for metals?

Video demonstration
4.2 How can we determine the reactivity series for metals?

Practical worksheet
4.2 How can we determine the reactivity series for metals?

Risk assessment
4.2 How can we determine the reactivity series for metals?

Context

The chemical reactivity of a metal is determined by its position in the periodic table.

In this practical, you will observe and prepare a series of metal reactions, using water, acid and oxygen, to create your own reactivity series.

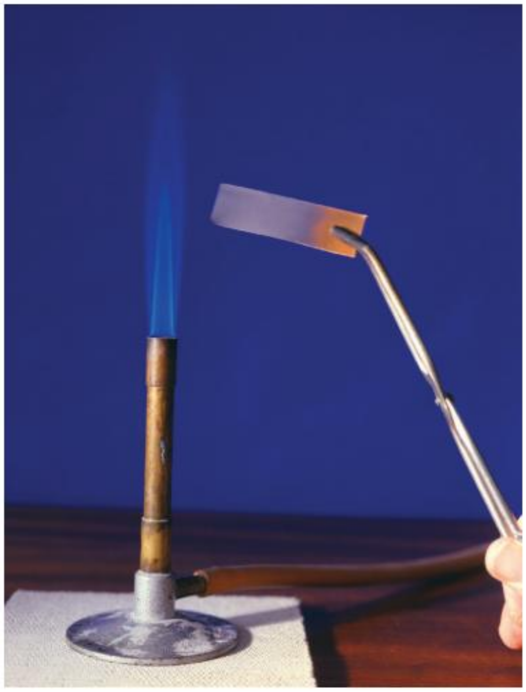


FIGURE 1 Burning copper with a Bunsen burner

Aim

To construct a reactivity series of selected metals by observing and comparing their reactions with hydrochloric acid, oxygen and water

Materials

For Part A:

- Sodium metal pieces
- Potassium metal pieces
- Calcium metal pieces
- Trough (three-quarters filled with water)
- Safety screen with wooden stabilisers
- Paper towel

For the class to share:

- Kettle (for hot water)
- Striker or a box of matches

For Parts B–D:

- Pieces of metals (copper, iron, magnesium ribbon, tin, aluminium, zinc)
- Dropper bottle of 2 M hydrochloric acid
- Dropper bottle of phenolphthalein solution
- Dropper bottle of water
- Wooden test-tube holder
- Wooden test-tube rack
- Plastic tweezers
- Marker pen
- Small glass test tubes
- Piece of emery or sandpaper
- Bunsen burner
- Metal tongs
- Heatproof mat
- Heatproof gloves
- Black tile
- PPE (safety screen, lab coat, safety glasses)

Method

Part A – Teacher demonstrations

1 Observe the teacher demonstrations of the reactions of sodium, calcium and potassium with cold water.

2 Record your observations in the Results table.

Part B – Reactions with water

3 Using tweezers, place a 1–2 cm sample of each metal, pre-cleaned with emery or sandpaper, into labelled test tubes.

4 Add water to each of the test tubes to a depth of 2 cm.

5 Add 2–3 drops of phenolphthalein solution to each test tube.

6 Record your observations in the Results table.

7 Gently heat the test tubes until the water is hot (but not boiling).

8 Record your observations in the Results table.

9 Carefully pour the water out of the test tubes (while retaining the metal samples).

Part C – Reactions with acid

10 Add 2 M hydrochloric acid to a depth of 2 cm *only* to each of the test tubes containing the metals that did not react readily with water.

11 If there is a reaction, hold a lit match in the neck of the test tube.

12 Record your observations in the Results table.

Part D – Reactions with oxygen

13 Using the Bunsen burner, heat the metal samples (iron, copper and zinc), one metal at a time, over the black tile.

14 Remove the metal sample from the heat and observe any changes.

15 Record your observations in the Results table.

Results

Download the practical worksheet in your obook pro and fill in the Results table.

Discussion

1 Construct balanced chemical equations for the metals that reacted with water.

2 Construct balanced chemical equations for the metals that reacted with hydrochloric acid.

3 Construct balanced chemical equations for the metals that reacted with oxygen.

4 Explain why a lit match was held in the neck of the test tube for reactions between the metal and hydrochloric acid.

5 Explain what a change in the appearance of the metals after rubbing with emery or sandpaper suggests.

Conclusion

Write a conclusion for your experiment that summarises your results, includes an analysis of possible errors, suggestions for improvement and a statement of which metals were the most reactive to the least reactive.

4.3
PRACTICAL: FIELDWORK

How is metal recovered from scrap metal?

Practical worksheet
4.3 How is metal recovered from scrap metal?

Risk assessment
4.3 How is metal recovered from scrap metal?

Context

Scrap metals are metals in discarded products (e.g. whitegoods, junk vehicles, alloy wheels, bicycles, used batteries). They often contain a mixture of different substances.

Metal recycling facilities accept (or often buy) scrap metal, which they can melt down, purify and sell to manufacturers.

In this practical, you will investigate current processes for recycling metals by visiting a metal recycling facility.

FIGURE 1 Scrap metals are recycled.

Aim

To visit a scrap metal recycling facility and gain a better understanding of the processes involved in recovering metals from scrap metal

Instructions

Visit a scrap metal recycling facility (e.g. Norstar Steel Recyclers) to learn about the processes they use to recover metals from scrap metal. Answer the following questions.

Questions

1 Create a flow chart of the steps involved in metal recycling.

2 Identify the green chemistry principles that relate to the recycling of scrap metal.

3 Distinguish between ferrous and non-ferrous scrap metal. Provide an example of each.

4 You will notice that many different types of scrap metals can be recycled.
  a List the types of scrap metal recycled at the facility you are visiting.
  b Select one of the scrap metals you listed in part **a** and identify the main metallic element that can be recovered.
  c Comment on how much of the pure metal you identified in part **b** is recovered from the scrap metal.
  d Analyse and evaluate the process(es) used to recover the metal. Consider the cost (financial and environmental) of the process to determine whether it is sustainable.

5 There are some substances and products that metal recycling facilities will not accept, even if they contain metal.
  a List three examples of substances/products that metal recycling facilities will not accept.
  b Suggest why one of the examples you listed in part **a** may not be recycled.
  c Describe the environmental impact of not recycling these types of substances. Identify where these products end up.

5.1 PRACTICAL: CONTROLLED EXPERIMENT

How are the properties of ionic and covalent compounds different?

Video demonstration
5.1 How are the properties of ionic and covalent compounds different?

Practical worksheet
5.1 How are the properties of ionic and covalent compounds different?

Risk assessment
5.1 How are the properties of ionic and covalent compounds different?

Additional resource
Crystal shapes

Context

The physical properties of a substance such as flame colour, crystal structure, solubility and conductivity can give information about the type of bonding in a compound.

In this practical, you will compare the general properties of selected ionic compounds and a covalent compound.

FIGURE 1 A flame test involves observing the flame colour.

Aim

To determine the general properties of ionic compounds and compare them with the properties of a covalent compound

Materials

For the class to share:

- Box of matches
- Paper towel

For each group of 2 or 3 students:

- Stereomicroscope
- 25 mL measuring cylinder
- Cut wooden skewers (pre-soaked in water and with points removed)
- Dropper bottle of distilled water
- Petri dishes
- Bunsen burner
- Heat mat
- Solid sodium chloride
- Solid copper(II) sulfate
- Solid magnesium sulfate
- Solid potassium chloride
- Sugar (sucrose)
- Graphite electrodes
- Power pack
- Connecting wires with alligator clips
- Globe
- Spatula
- 50 mL beakers
- Electronic balance
- Stopwatch
- Glass stirrer
- PPE (lab coat, safety glasses)

Method

1 Carefully place a sample of each of the solid compounds into five separate Petri dishes.
2 View each compound under the stereomicroscope and record your observations in Results table 1.
3 Set up the power pack, connecting wires, globe and graphite electrodes as a closed circuit.
4 Dip both of the electrodes into the Petri dish of one solid compound, without letting them touch each other, to see if the globe lights up. Clean the electrodes after testing the compound.
5 Repeat with the other solids and record your findings in Results table 2.

6 Turn the gas on and carefully light the Bunsen burner. Place a pre-soaked cut skewer into one of the solid compounds, then hold it in the blue flame of the Bunsen burner until a colour appears. Repeat with the other compounds, using a new skewer each time, and record your findings in Results table 2. Remember to turn the gas off when you are finished.
7 Measure 1.0 g of one solid compound into a small beaker. Add 25 mL of distilled water and stir with the glass rod. Record the time it takes for the solid to dissolve in the water in Results table 3. Repeat with the other compounds. Remember to clean the spatula and glass rod each time.
8 Measure the conductivity of the dissolved compound in solution with the same electrodes and circuit set up as in step 4. Clean the electrodes after testing the solution. Repeat with the other solutions and record your findings in Results table 3.

Results

1 Download the practical worksheet in your obook pro and fill in the Results tables.
2 In the **Description** column of Results table 1, describe the colour and crystalline shape of the compounds that you observed through the stereomicroscope.
3 In the **Crystal structure and name** column of Results table 1, sketch a representative crystal and identify the name of the crystalline shape as closely as possible. You can access the index of crystal shapes on your obook pro to help you.
4 In the **Solid conductivity** column of Results table 2, record the conductivity level as none, low or high.
5 In the **Flame test colour** column of Results table 2, describe your observations when the solid was held in the blue flame of the Bunsen burner.
6 In the **Time to dissolve** column of Results table 3, record the time taken for the solid to dissolve. In the **Solubility column**, record the solubility level as completely, partially or insoluble.
7 In the **Solution conductivity** column of Results table 3, record the conductivity level as none, low or high.

Discussion

1 Identify the covalent compound investigated.
2 From Results table 1, list any similarities and differences in the appearance and crystalline structure of the solid compounds.
3 From Results table 2:
   a describe the electrical conductivity of the compounds in the solid state
   b identify whether all compounds produced a flame colour.
4 From Results table 3:
   a identify the compound that was the most soluble
   b identify the compound that was the least soluble
   c identify whether all of the dissolved compounds conducted electricity.
5 Use the internet to find the melting points for NaCl, KCl, $CuSO_4$, $MgSO_4$ and sugar. Record and compare the values.

Inquiry: What if the covalent compound investigated was graphite?

1 Write a hypothesis for this investigation.
2 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.
3 Identify one assumption you made that could affect your results.
4 Compare the properties of graphite with those of sugar used in the first investigation. Explain any differences and refer to the structure of graphite in your answer.

Conclusion

Summarise your findings from this practical by comparing the properties of ionic and covalent compounds.

6.1
PRACTICAL: SIMULATION

How do the sample components behave in chromatography?

Practical worksheet
6.1 How do the sample components behave in chromatography?

Context

Chromatography involves manipulating the properties of the mobile and stationary phases in order to separate the components of a mixture or sample. Components of a mixture that have a higher affinity for the stationary phase spend longer interacting with it and do not move as far from the origin line. Meanwhile, components with a higher affinity for the mobile phase spend longer interacting with it and move further from the origin line.

In this practical, you will simulate different properties of the mobile phase to determine their effect on the separation of the components of a mixture.

Aim

To determine the effect of changing the properties of the mobile phase on the distance that a sample component moves and, therefore, its R_f value

Equipment

- 21 pieces of paper, laminated paper or cardboard (to mark points on the floor), one labelled 'Origin' and the remaining 20 labelled from '1' to '20'. (This can also be done by using chalk on a pavement if you are working outside)
- Sample component cards – 4 red cards, 4 yellow cards, 4 blue cards and 4 green cards that students can hold up to demonstrate which component of the sample they are
- One stopwatch per student
- Pen/pencil and a piece of paper to record results

Instructions

1. Lay the 21 pieces of paper in a straight line where they can be easily seen. The first will be the Origin (as shown in the Results table).
2. Four students will be chosen. You will each receive a stopwatch and a different coloured card. Once you have everything you need, line up with the Origin (as shown in the Results table).
3. Your teacher will explain these instructions to you:

   When your teacher says go, you will move from the Origin to the final piece of paper. However, each coloured card has a different length of time on it:

   Red: 10 seconds

   Yellow: 8 seconds

   Green: 6 seconds

   Blue: 4 seconds

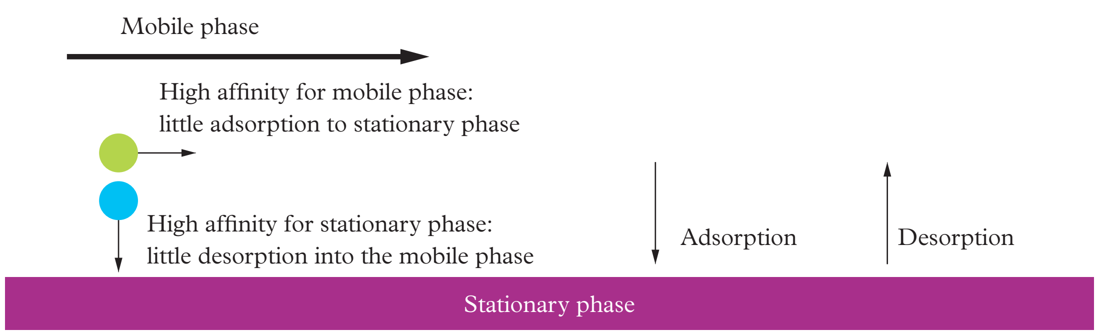


FIGURE 1 The blue and green components in a sample can be separated depending on their affinity for the mobile and stationary phases.

You must spend this length of time at each piece of paper before you move on to the next piece of paper. When your teacher says stop, stay at the piece of paper you are currently at.

4 In the first simulation, you are moving with a very polar mobile phase. Your teacher will say 'go', run the stimulation for 60 seconds, then say 'stop'.

5 Each student must record how many pieces of paper they are from the origin. This number will be on the piece of paper at their location.

6 Repeat the simulation by altering the properties of the mobile or stationary phases. As a comparison, the first four students can remain where they are and four new students can be selected and given new coloured cards and stopwatches. Options include:
   - **a** decreasing the polarity of the mobile phase – increasing the time of the blue, green and yellow cards and decreasing the time of the red card
   - **b** making the mobile phase non-polar – giving the red card the shortest time on each piece of paper and giving blue the longest.

Results

1 Download the practical worksheet in your obook pro and fill in the Results table. Add a new column for every simulation conducted. For each experiment, you must record the:
   - **a** properties of the mobile phase
   - **b** distance travelled by each component of the mixture (red, yellow, green and blue cards).

2 For each simulation, calculate the R_f values for each component of the mixture. Note: in all simulations, the solvent travels from the origin to piece of paper '20'.

Questions

1 Rank the components of the sample (the colour of each card) from most to least polar, based on the results of the first simulation.

2 Based on the results, explain which component has a higher affinity to the mobile phase and which component has a higher affinity to the stationary phase.

3 Describe what would happen if the mobile phase was half as polar as it was in the first simulation.

4 Describe what might have occurred if the R_f value was 0 or 1.

5 Explain what happens to the order of the components when a non-polar mobile phase is used.

6 Of the separations that were simulated by your class in this activity, identify the one that resulted in the best separation. Justify your response.

7 A student in your class wants to simulate the effects of changing the properties of the stationary phase. Suggest what variables you could change in this simulation to investigate this.

6.2 PRACTICAL: CONTROLLED EXPERIMENT

How can chromatography be used to separate food dyes in Skittles?

Video demonstration
6.2 How can chromatography be used to separate food dyes in Skittles?

Practical worksheet
6.2 How can chromatography be used to separate food dyes in Skittles?

Risk assessment
6.2 How can chromatography be used to separate food dyes in Skittles?

Context

The components of a mixture can be separated by paper chromatography and/or thin-layer chromatography (TLC).

In this practical, you will compare the dyes in food products with the claims made on the packets. Skittles are used in this experiment, because the shells contain food dyes.

The food dyes in Skittles are shown in Table 1.

TABLE 1 Food dyes in Skittles

Number	Name	Colour
171	Titanium dioxide	White
129	Allura red AC	Red
133	Brilliant blue FCF	Blue
110	Sunset yellow FCF	Yellow
102	Tartrazine	Yellow
132	Indigo carmine	Blue

Aim

To determine the identity of food dyes in various colours of Skittles and compare them with the manufacturer's claims

Materials

- Large packet of Skittles (approximately 10 of each colour are needed)
- Standard dyes (10% – diluted 1:10 in water)
- 200 mL of saturated salt solution (NaCl)
- 50 mL beakers
- Wooden skewers or toothpicks
- Filter paper or chromatography paper cut into strips (the size depends on your chosen methodology – if you analyse the colours separately, use a smaller piece and a smaller beaker. If you analyse them all together, use a larger piece and a larger beaker.)
- Beaker (size dependent on size of the filter paper chosen)
- Sticky tape
- Scissors
- Pencil
- Ruler
- PPE (lab coat, safety glasses)

Method

1. Sort similar-coloured Skittles into separate piles.
2. Add 10 of each coloured Skittle to separate 50 mL beakers. Add 10 mL of water to dissolve the dye in the water. Leave them for 5 minutes.
3. Set up the filter paper as shown in Figure 1. Draw the origin in pencil. You can add the solution front after the experiment.
4. Use the wooden skewers or toothpicks (a different one for each sample) to dot the sample onto the filter paper. The dots must be spaced approximately 1 cm apart. One or two drops is enough on each spot.

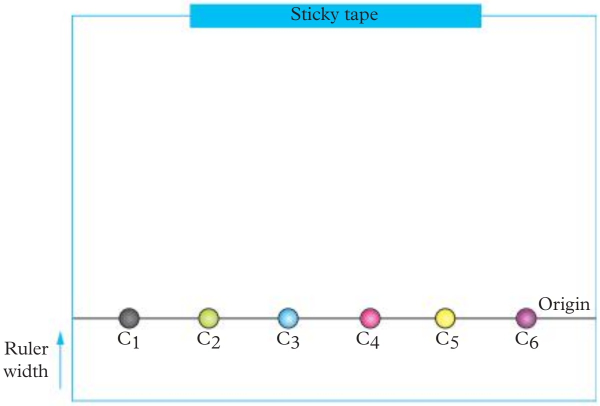


FIGURE 1 The set-up of filter paper for separation of coloured samples, where C represents a different colour. Note: C may be changed to the colour name, e.g. blue, green.

5 Add the filter paper to the beaker, making sure that the mobile phase is not above the sample spots. If it is, the dots will dissolve into the mobile phase at the bottom of the beaker and you will have to start again. Use the sticky tape to tape the top of the filter paper to the beaker.
6 Leave the beaker until the mobile phase moves towards the top of the filter paper. Do not let the mobile phase reach the bottom of the sticky tape.
7 Remove the filter paper from the beaker and allow it to dry.
8 Once dry, draw a line to show the solvent front.
9 Repeat the method by spotting the standard dyes onto a separate sheet of filter paper as shown in Figure 2.

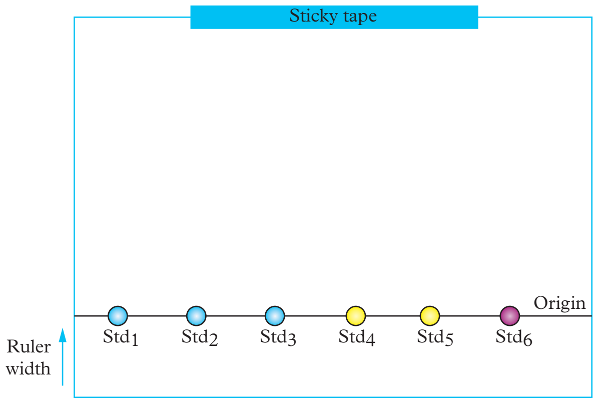


FIGURE 2 The set-up of filter paper for separation of standard dyes, where Std represents a different standard. Note: Std may be changed to the dye name, e.g. titanium dioxide, sunset yellow FCF.

Results

1 Download the practical worksheet in your obook pro and fill in the Results tables.
2 Using a ruler, determine the distance between the origin line and the:
   a solvent front
   b components of the dyes; and
   c standards.
   Glue your chromatogram into your logbook.
3 Calculate the R_f of the sample dyes and the standards. Record your results in the Results tables. Add as many rows as necessary depending on the number of components in each coloured sample. For example, if brown separates into three colours, you can have three rows: Colour 1, Colour 2 and Colour 3.

Discussion

1 Compare the coloured samples with the standard dyes to identify the dyes. Are the identities of the dyes on the back of the food packaging correct? Justify your response.
2 Use the R_f values that you calculated and the properties of the mobile phase to describe the polarity of the dyes.
3 Identify any issues with the separation or distance moved by the dyes.
4 Identify one way the resolution of the separation (how well separated the components are) could be improved.

Inquiry: What if the mobile phase was more or less polar?

1 Write an aim for your investigation.
2 Write a hypothesis for this investigation.
3 Predict what will happen to the R_f values in a separation with a more polar mobile phase.
4 Predict what will happen to the R_f values in a separation with a less polar mobile phase.

Conclusion

Summarise your findings from this practical.

7.4
PRACTICAL:
CONTROLLED EXPERIMENT

How can we experimentally determine the empirical formula of magnesium oxide?

Video demonstration
7.4 How can we experimentally determine the empirical formula of magnesium oxide?

Practical worksheet
7.4 How can we experimentally determine the empirical formula of magnesium oxide?

Risk assessment
7.4 How can we experimentally determine the empirical formula of magnesium oxide?

Context

Magnesium ribbon, a shiny solid metal, reacts with oxygen in the heat of a Bunsen burner flame to form magnesium oxide, a light grey, crumbly ash.

In this practical, you will compare the mass of magnesium at the start of the reaction with the mass of magnesium oxide at the end of the reaction to determine the empirical formula of magnesium oxide.

FIGURE 1 Magnesium oxide is a light grey, crumbly ash.

Aim

To determine the empirical formula of magnesium oxide by measuring the mass of magnesium oxide formed after burning a known mass of magnesium ribbon in a Bunsen flame

Materials

- Magnesium ribbon
- Electronic balance
- Metal tongs
- Bunsen burner
- Tripod
- Clay triangle
- Matches
- Crucible and lid
- Heat-resistant mat
- Scissors
- Sandpaper
- PPE (lab coat, safety glasses)

Method

1 Cut a piece of magnesium 10–15 cm long.
2 Rub the surface of the magnesium ribbon with sandpaper to remove any black tarnished areas.
3 Measure the mass of the crucible and the lid together.
4 Measure the mass of the crucible and lid with magnesium ribbon inside.
5 Set up a Bunsen burner on a heat-resistant mat with a tripod and clay triangle over the top. Place the crucible containing the magnesium ribbon onto the clay triangle.
6 Light the Bunsen burner. Use an open air hole (blue flame). Heat the magnesium in the crucible with the lid on for 5–10 minutes or until a bright light is emitted. Use the metal tongs to open the lid periodically or leave the lid on but slightly ajar.
7 When the reaction is complete, the shiny magnesium ribbon will have all reacted and a light grey, brittle magnesium oxide solid will have formed. At this point, turn off the Bunsen burner and allow the crucible to cool.
8 Measure the mass of the crucible with the magnesium oxide inside and the lid together. Do not place hot objects directly onto the electronic balance.

Results

1 Record the masses of the:
  a crucible and lid together
  b crucible and lid with magnesium ribbon inside
  c crucible and lid with magnesium oxide inside.
2 Calculate the mass of magnesium ribbon that was used in the experiment.
3 Calculate the mass of oxygen that was added to the magnesium.
4 Use an $m:n$ ratio table with Mg and O as headings to calculate the empirical formula.

Discussion

1 Write a word equation for the reaction taking place in this experiment.
2 Compare the empirical formula of magnesium oxide that you determined with the actual empirical formula. Explain any differences between the values.
3 Evaluate whether the results are accurate. Justify your answer.
4 Evaluate the validity of the experiment. Justify your answer.

Inquiry: What if oxygen was in limited supply?

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify two materials or pieces of equipment that require a risk assessment.
4 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.
5 Identify one assumption you made that could affect your results.

Conclusion

Write a conclusion for your experiment that summarises your results, includes an analysis of possible errors, suggestions for improvement and the empirical formula of magnesium oxide.

8.1

PRACTICAL: CLASSIFICATION & IDENTIFICATION

Can you identify organic compounds from their physical properties?

Practical worksheet
8.1 Can you identify organic compounds from their physical properties?

Context

The presence of different functional groups, and the size and structure of organic compounds, affect their physical properties, including boiling and melting points.

In this practical, you will apply this knowledge to identify organic compounds on the basis of their physical and chemical properties.

FIGURE 1 The boiling point of an organic compound depends on its structure.

Aim

To identify organic chemicals from their physical properties

Questions

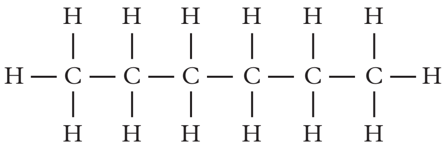
Part A – Size and structure

The boiling points of four organic compounds A–D are shown in Table 1.

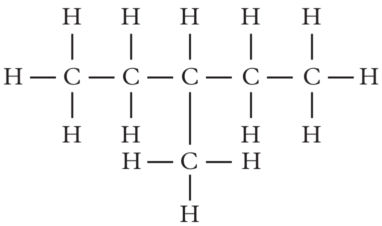
TABLE 1 The boiling points of organic compounds A–D

Compound	Boiling point (°C)
A	50
B	58
C	64
D	69

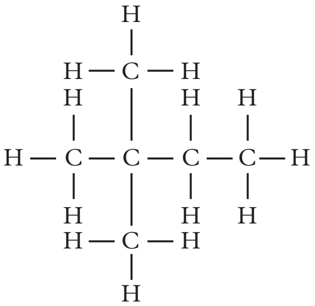
We know that these are the boiling points of hexane and three of its isomers (shown below), but we don't know which is which.



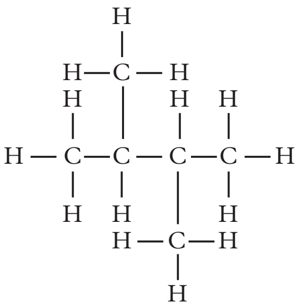
Hexane



3-Methylpentane



2,2-Dimethylbutane



2,3-Dimethylbutane

1 Copy the structural formulas (and names) of these molecules onto small cards.
2 Copy Table 1 into your logbook, leaving space to paste in the matching structural formulas you have drawn.
3 Match the compounds on your cards with their correct boiling point. Justify your choices.
4 Discuss if you had doubts about matching the isomers with their boiling points. Explain why or why not.

Part B – Functional groups

The melting and boiling points of organic compounds E–K are shown in Table 2.

TABLE 2 The melting and boiling points of organic compounds E–K

Compound	Melting point (°C)	Boiling point (°C)
E	–114.0	78.0
F	–143.2	–37.7
G	–182.79	–88.6
H	16.2	117.5
I	119.0	38.4
J	–169.0	–104.0
K	–138.7	12.3

We know that these are the properties of the molecules listed below, which have the same number of carbons but different functional groups. Again, we don't know which is which.

- Ethane
- Ethene
- Fluoroethane
- Chloroethane
- Bromoethane
- Ethanol
- Ethanoic acid

1 Write the structural formulas (and names) of the molecules above on small cards.
2 Copy Table 2 into your logbook, leaving space to paste in the matching structural formulas you have drawn.
3 Match the compounds on your cards with their correct boiling point. Justify your choices.
4 Outline your methodology for matching the structures to their physical properties.
5 Discuss the usefulness of having both melting point and boiling point data for the compounds. Did you use both? Why or why not?

8.4 PRACTICAL: CONTROLLED EXPERIMENT

Which milk makes the strongest glue?

Video demonstration
8.4 Which milk makes the strongest glue?

Practical worksheet
8.4 Which milk makes the strongest glue?

Risk assessment
8.4 Which milk makes the strongest glue?

Context

Milk contains a protein called casein, which can be used to make a strong glue. This process involves two steps.

1 Using vinegar to separate milk into its curds (which contain the casein) and whey
2 Reacting (neutralising) the curds with a base to produce the glue

However, different types of milk sold in supermarkets may have different amounts of casein.

In this practical, you will use different types of milk to make glue and determine which milk makes the best glue.

You can investigate:

- different brands of milk
- whole milk versus reduced-fat milk, low-fat milk or skim milk
- cow's milk versus goat's milk.

FIGURE 1 Which milk makes the strongest glue?

Aim

To investigate the strength of glue made from different types of milk and identify the type of milk that makes the strongest glue

Materials

- A variety of milks (100 mL of each)
- Vinegar
- At least one solid base (e.g. sodium hydrogen carbonate, magnesium carbonate, calcium carbonate, sodium hydroxide)
- 100 mL and 25 mL measuring cylinders
- 250 mL beakers
- 250 mL conical flasks
- Filter funnel
- Filter paper or Chux wipe
- pH indicator paper (suitable for neutral substances)
- Spatula
- Stirring rod
- Bunsen burner
- Heatproof mat
- Tripod
- Gauze mat
- Icy-pole sticks
- 100 g masses
- PPE (lab coat, safety glasses)

Method

Part A – Making the glue

1 Measure 100 mL of the first milk you will investigate and pour it into a large beaker.
2 Measure 20 mL of vinegar and add it to the beaker containing the milk.

3 Place the beaker on the tripod set up over the Bunsen burner flame and warm the milk. Stir the milk constantly.
4 When small lumps start to form in the mixture, turn off the heat. Continue stirring until no more lumps form.
5 Let the lumps settle, then decant the liquid to separate it from the solids.
6 Using the filter funnel , filter paper (or Chux® wipe) and conical flask, separate the rest of the liquid from the solids. The solids are the curds, which you will need to keep. You can dispose of the liquids.
7 Gently squeeze off all the remaining liquid from the curds and add them to a clean beaker. Add 15 mL of water to the curds and stir the mixture until it is smooth.
8 Add approximately half a spatula full of the solid base to the smooth curd mixture and stir it well.
9 Check if the mixture is neutral by using the pH indicator paper. If it is still acidic, add more base, a small amount at a time. Stir and check the pH between each addition.
10 Stop once you have a neutral pH.
11 Repeat steps 1–10 with each milk. Make sure you clearly label the beakers for each type of milk.

Part B – Testing the glue strength

12 Use the glue to stick two icy-pole sticks together. Make sure about 2 cm of the sticks are overlapping. Label the icy-pole sticks with the type of milk used to make the glue. Prepare three sets of sticks for each glue.
13 Leave the glue to dry (overnight, if possible).
14 Arrange two stacks of books, two tripods or two tables together so that you can lay the icy-pole sticks on them like a bridge.
15 Stack masses on the icy-pole stick bridges as close to the centre as possible. Add them 100 g at a time.
16 Record the mass required to break the glue. The more mass the icy-pole sticks and glue can hold, the stronger the glue is.

Results

1 Create a results table to record your results in your logbook.
2 Create a graph of your results.

Discussion

1 Identify the milk that made the strongest glue.
2 Compare your results with those of the rest of the class. Identify whether the results were consistent across all groups.
3 Identify and discuss any errors you made during this experiment.
4 Discuss how the concept of making glue from milk is an example of the green chemistry principle ‘Renewable feedstocks’, but not an example of a linear to circular economy.

Inquiry: Does the type of base used affect the strength of the glue?

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify two materials or pieces of equipment that require a risk assessment.
4 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.
5 Identify one assumption you made that could affect your results.
6 Discuss the results of your investigation.
7 Discuss why the type of base is important for neutralising the acid in the glue. Use the internet to help you.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors, suggestions for improvement and a statement of the type of milk that makes the strongest glue.

9.3 PRACTICAL: CONTROLLED EXPERIMENT

What effects do cross-links have on the properties of polymers?

Video demonstration
9.3 What effects do cross-links have on the properties of slime?

Practical worksheet
9.3 What effects do cross-links have on the properties of slime?

Risk assessment
9.3 What effects do cross-links have on the properties of slime?

Context

Slime can be made by cross-linking polyvinyl alcohol (PVA). PVA is a polymer made of repeating monomers of ethenol (vinyl alcohol) (Figure 1).

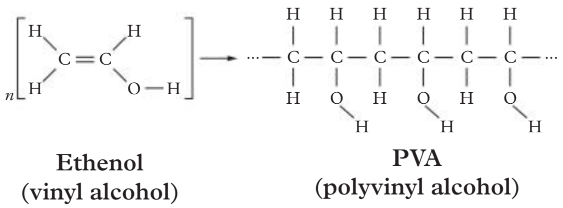


FIGURE 1 Ethenol (vinyl alcohol) monomers make the polymer PVA (polyvinyl alcohol).

PVA polymers can be cross-linked using borax. In solution, borax forms borate ions (Figure 2).

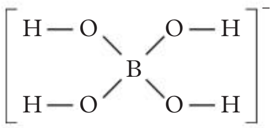


FIGURE 2 Borate ions form when borax is in solution

These form weak bonds with the hydroxyl (–OH) functional groups in the PVA polymer chains (Figure 3).

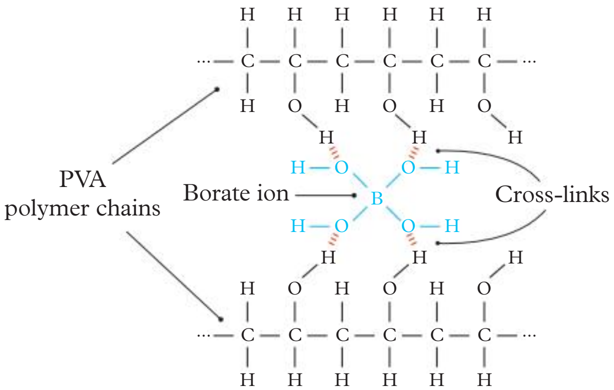


FIGURE 3 Cross-linking occurs between the –OH groups in the PVA polymer chains and the borate ions.

In this practical, you will create a slime polymer with the most cross-linking and compare it with slime with less cross-linking to determine the effect that the cross-links have on slime properties.

Aim

To compare the properties of slime with different amounts of cross-linking between the polymer chains

Materials

- Polyvinyl alcohol (PVA) solution (40 mL of 6–12% solution)
- Borax solution
- Food colouring
- 100 mL beaker or disposable plastic cup
- 10 mL and 50 mL measuring cylinders
- Glass stirring rod
- Petri dish or watch glass
- Wax paper or similar
- Disposable gloves
- PPE (lab coat, safety glasses)

Method

Part A – Slime 1: 40 mL PVA and 10 mL borax

1. Measure 40 mL of the PVA solution and pour it gently into a beaker or plastic cup.
2. Add one drop of food colouring of your choice.
3. Measure 10 mL of the borax solution and add it to the beaker or cup with the PVA solution.
4. Stir vigorously until a 'slime' is formed.
5. Wearing disposable gloves, remove the slime from the cup and knead it thoroughly until it is completely mixed. Roll the slime around in your hands and squeeze it gently to remove all the air bubbles.

Part B – Slimes 2–5: 40 mL PVA with different amounts of borax

6 Repeat steps 1–5, but at step 3, change the volume of borax solution to 2, 4, 6 or 8 mL. You can make as many new slimes as you like, but to compare results, two or three different mixtures of slime are sufficient (e.g. 2, 6 and 10 mL).
7 Use different food colourings for the different volumes of borax and record the colours in your logbook. Keep slime away from clothes because it can leave permanent stains.

Results

1 Create a results table to record the results of each of the following tests:

Test 1: Try to pour each slime.

Test 2: Stretch and pull the slime apart slowly.

Test 3: Pull the slime apart quickly.

Test 4: Roll the slime into balls then try to bounce them (on wax paper or similar so you don't stain the bench or floor).

Test 5: Roll the slime into balls and sit the slime in Petri dishes or watch glasses. Draw a circle around the base of each, then let the slime sit for 10 minutes. Record how far they have gone beyond the circle you drew.

Test 6: Roll out the slime into the longest ropes you can. Measure the lengths and record them.

Discussion

1 Discuss the effect that the different amounts of cross-linking had on the properties of the slime.
2 Explain how the different amounts of cross-linking between the polymer chains can change the properties of the slime.
3 Identify whether slime is a liquid or a solid. Refer to the properties of liquids and solids in your answer. Explain if it is something completely different.
4 Describe the properties that a great slime should possess.
5 Identify which slime you thought was the best. Justify your answer with reference to cross-linking.

Inquiry: How do the properties of slime polymers change with temperature?

1 Write a hypothesis for this investigation.
2 Put your slime into plastic bags and store them in the fridge or freezer until cooled below room temperature. Remove and roll into ropes again (Test 6). Measure and record the lengths.
3 Explain how the properties of a polymer change with temperature.
4 Discuss whether you can conclude this for all polymers.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors, suggestions for improvement and a statement of which slime was the best and why.

FIGURE 4 Slime is made from polyvinyl alcohol.

9.5 PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can the design of copolymers be used to meet product requirements?

Practical worksheet
9.5 How can the design of copolymers be used to meet product requirements?

Context

Every polymer you encounter daily is designed for a specific purpose, with specific properties in mind. The polymers in your phone case are different from the polymers in the insulation surrounding your phone charging cord, which are also different from the polymers found in the tiny computer chips that make your phone work.

In this practical, you will analyse the properties of monomers to design copolymers with specific properties.

Aim

To analyse the properties of monomers and design copolymers with specific properties for a designated purpose

Instructions

Table 1 shows six monomers and their specific properties. You will use these monomers to create copolymers to be used in products 1–5.

Product 1 Design a copolymer that can be used for food storage bags. They should be able to store and protect foods for long periods of time under different conditions (e.g. fridge, freezer, lunch box, cupboards, kitchen bench). They need to be flexible, and most importantly compostable once used.

FIGURE 1 Product 1 – food storage bags

TABLE 1 Six monomers and their specific properties

Monomer A	Monomer B	Monomer C	Monomer D	Monomer E	Monomer F
• soft, tacky, and rubbery • high insulation resistance and low electrical losses • soluble in all common organic solvents	• rigid and very strong but not easily moulded into shapes • lightweight • transparent • not UV resistant	• flexible and easy to mould • biodegradable in a short amount of time and non-toxic • poor oxygen resistance • stable in the temperature range 15–30°C	• very strong but brittle • low moisture absorption, excellent weathering properties	• can be formed into fibres • insoluble in all organic solvents • poor temperature resistance • UV resistant	• very strong, flexible and tough • transparent and colourless but can be coloured • poor heat and light resistance

Product 2 Design a copolymer that would be suitable for making toy blocks. They should be brightly coloured and appealing to children. They must be extremely durable, resistant to light and heat, and not wear or discolour over time. They should also be non-toxic.

FIGURE 2 Product 2 – toy blocks

Product 3 Design a copolymer that can be used as the windscreen on a small speed boat. It should be transparent, extremely durable and weather resistant. It must be able to withstand high force winds and salt water.

FIGURE 3 Product 3 – boat windscreen

Product 4 Design a copolymer that would be suitable for the soles of workboots. They need to be light weight and flexible. They must be also be durable, resilient and hard-wearing as well as chemically resistant for use in different situations.

FIGURE 4 Product 4 – workboot soles

Product 5 Design a copolymer that would be suitable for single-use containers to store various consumer products, such as shampoo, conditioner, detergents, sprays and body lotions. They need to be UV and temperature resistant because the storage conditions for the products will vary. They should also be chemically resistant, so that they can store liquid products with different chemical compositions.

FIGURE 5 Product 5 – plastic containers

1 Design the copolymers for products 1–5.
2 For each copolymer, identify the monomers you will add. Justify your selection.
3 Describe the desired properties that each monomer will bring to the finished polymer.
4 Describe the undesired properties that a monomer may bring and how you will overcome them.

Questions

1 Describe the obstacles that you had to overcome for each design. What kinds of challenges did you face when selecting monomers to make your copolymers?

2 Imagine you had access to one more monomer, monomer G.

  a For each copolymer, identify any specific monomer properties that were missing. Determine whether access to monomer G would have helped your design.

  b Describe the properties that monomer G would have.

  c Identify in which products you would have used monomer G.

3 Design a sixth product using monomers A–F. Outline the properties that it should have and the monomers that you would use to create it. Justify your choices.

4 Analyse and evaluate the life cycle of each of the copolymers you created.

  a Identify whether they are to be used forever or need to be replaced.

  b If they need to be replaced, discuss whether you could recycle them after they have been used.

  c Discuss whether your copolymer products follow the linear or circular economy model.

  d Explain what you could do to each product to make them more sustainable.

FIGURE 6 Different polymers have specific properties to make them usable for their own tasks.

11.1A PRACTICAL: FIELDWORK

What is the quality of water sourced from different locations?

Practical worksheet
11.1A What is the quality of water sourced from different locations?

Risk assessment
11.1A What is the quality of water sourced from different locations?

Context

Water quality varies greatly depending on where it is found. Surface water is often potable (drinkable) and easily accessible.

In this practical, you will compare the quality of water from various sources, including tap water from different areas of the school and in your home, rainwater tanks, distilled water, mineral water and water from the local environment where aquatic life is found.

FIGURE 1 Water samples can be collected from a local pond.

Aim

To investigate the quality of water from different sources by measuring the amounts of various minerals and pollutants by using rapid diagnostic water-testing strips

Materials

- 10 small bottles or test tubes with lids for collecting water samples
- 10 water-testing strips
- Thermometer

Method

1 Collect 10 samples of water from various locations:
   - taps around the school and home
   - the local environment where aquatic life is found
   - a rainwater tank
   - distilled water from the lab
   - mineral water from a bottle.

2 Record any observations about the appearance of the water.

3 Test the temperature of the water immediately at the source.

4 Test the water samples for minerals and various pollutants by using rapid diagnostic water-testing strips.

Results

1 Record your results in a results table.

2 Use the instructions provided with the water-testing strips to determine whether the water is above or below ‘safe’ levels for human consumption.

Discussion

1 Based on your findings, rank the water samples in order of how safe they are for humans to drink.

2 Determine whether any of the water samples are considered dangerous for consumption. Explain why this might be the case.

3 Discuss any limitations of this investigation. List any important factors that could affect the quality of water but were not tested in this investigation.

4 Discuss whether the appearance of the water source was a good indicator of water quality.

Conclusion

Write a conclusion for your experiment that summarises your results, includes an analysis of possible errors, suggestions for improvement and a statement of the quality of water sourced from different locations.

11.1B
PRACTICAL:
PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can contaminated water be purified for drinking?

Practical worksheet
11.1B How can contaminated water be purified for drinking?

Context
Most water on Earth is not safe for human consumption. Some water sources are too salty or too contaminated with dirt and chemical pollutants to drink.

In this practical, you will design a process to purify water that has been contaminated with salt and dirt (i.e. seawater).

FIGURE 1 Can you purify seawater for drinking?

Aim
To design a process to make seawater drinkable by using equipment available in your school laboratory

Instructions
1 Conduct some research into how salt and dirt can be removed from water to make it drinkable.
2 List all the materials and equipment that you will need. Remember that you can only use what is available in your school laboratory.
3 Design a method and list all of the steps you will need to follow to purify the seawater. You could also include instructions on how you will collect the seawater.

Questions
1 Identify at least two tests you will need to perform on your purified water to check that it is safe for drinking.
2 Determine whether a risk assessment is required for the process you designed. Justify your response.
3 Share your method with another student in your class. Discuss any differences in your methods. Identify any changes you might like to make to your method and explain why.
4 Discuss whether your method could be used by any Chemistry student in Victoria. Explain why or why not.
5 Discuss whether your method could be used on a larger scale to purify enough water to supply the whole of Victoria. Explain why or why not.
6 Discuss whether purification of seawater for drinking is sustainable. Refer to the relevant sustainability perspectives (sustainable development, green chemistry principles, the move from a linear economy towards a circular economy) in your answer.

11.2
PRACTICAL: CONTROLLED EXPERIMENT

How can you create a heating curve for liquid water?

Video demonstration
11.2 How can you create a heating curve for liquid water?

Practical worksheet
11.2 How can you create a heating curve for liquid water?

Risk assessment
11.2 How can you create a heating curve for liquid water?

Context

As water is heated at a constant rate using a hot plate, heat energy is absorbed by the hydrogen bonds. The temperature increases as hydrogen bonds are gradually broken. The change in state and temperature of H_2O as heat energy is absorbed can be shown on a heating curve.

In this practical, you will construct your own heating curve for water.

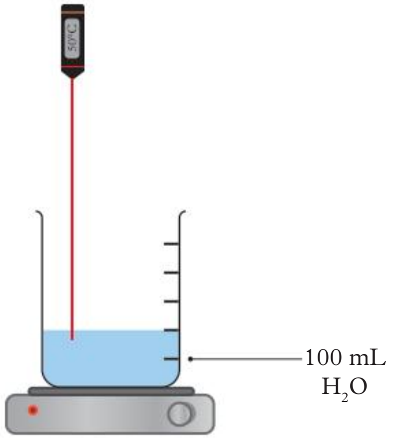


FIGURE 1 The experimental set-up for constructing a heating curve for water

Aim

To experimentally determine the change in temperature of H_2O during heating and construct a heating curve

Materials

- Hot plate with known power output (you can calculate the power output in joules per second by multiplying the voltage by the current, which should be written on the device)
- 250 mL beaker
- 100 g of cold distilled water (5–20°C)
- Balance
- Digital thermometer or temperature sensor
- Stopwatch
- PPE (lab coat, safety glasses)

Method

1. Turn on the hot plate to maximum power.
2. Place the 250 mL beaker on the balance and tare it. Measure 100 g of water into the beaker.
3. Add 100 g of cold water to the 250 mL beaker and insert the thermometer or temperature sensor. Wait for the temperature to stabilise and record the initial temperature of the water.
4. When the hot plate has reached maximum temperature, place the beaker of water onto the hot plate and start the stopwatch.
5. Record the temperature every 30 seconds as it increases. Stir slowly and continuously.
6. Take note of the temperature when the water starts boiling. Continue recording the temperature.
7. Continue boiling the water until the total volume has evaporated, then turn the hot plate off.
8. Repeat steps 2–6 at least two more times with fresh water.

Results

1. Record your results in a results table.
2. Convert the 30 s intervals to amounts of energy supplied to the water using the formula:

   energy supplied (joules) = power of hot plate (watts) × time (seconds)

   For example, if the hot plate is 1400 W, then each 30 s interval has supplied 42 000 J or 42 kJ to the water:

   energy supplied (joules) = 1400 × 30 = 42 000 J
3. Plot your results on a graph with energy supplied on the *x*-axis and temperature on the *y*-axis. This is your heating curve.

4 Identify the boiling point of H_2O from the results you obtained.

5 Calculate the specific heat capacity of water (c) from the equation:

$$q = mc\Delta T$$

where q is energy (J), m is mass of the substance being heated (g), c is the specific heat capacity of the substance being heated (J g^{-1} K^{-1}) and ΔT is the temperature change of the substance being heated (K).

Discussion

1 Describe what is happening to any intermolecular forces between the water molecules at the boiling point temperature.

2 Compare your boiling point with the theoretical boiling point of water (100°C). Discuss the factors that could account for any differences between your observed result and the theoretical boiling point of water.

3 Compare your calculated specific heat capacity with the theoretical specific heat capacity of water (4.18 J g^{-1} $°C^{-1}$). Discuss the factors that could account for any differences.

Inquiry: What if seawater was used instead of distilled water?

1 Write a hypothesis for this investigation.

2 Identify the *independent* variable that you will change from the first method.

3 Identify two materials or pieces of equipment that require a risk assessment.

4 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.

5 Identify one assumption you made that could affect your results.

6 Discuss the results of your investigation.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors and suggestions for improvement.

12.1
PRACTICAL: CLASSIFICATION & IDENTIFICATION

Can you identify acids, bases and neutral compounds?

Practical worksheet
12.1 Can you identify acids, bases and neutral compounds?

Context

Acids are compounds that can donate hydrogen ions (H^+) or protons. Bases are compounds that can accept protons. Neutral compounds cannot donate or accept protons.

In this practical, you will construct models of compounds and predict whether they will be acidic, basic or neutral based on their structure.

FIGURE 1 A molecular modelling kit

Aim

To create models of compounds and predict whether they will be acidic, basic or neutral based on their structure

Materials

- Molecular modelling kit
- Plastic tub (to keep parts together)

If a molecular modelling kit is not available:

- Icy-pole sticks (to represent chemical bonds)
- Black paper (for carbon atoms)
- Red paper (for oxygen atoms)
- Blue paper (for nitrogen atoms)
- Green paper (for halogen atoms)
- White paper (for hydrogen atoms)
- Purple paper (for phosphorus atoms)
- Brown paper (for metal atoms)
- Scissors (to cut the paper or cardboard into circles)

Method

1 Using the modelling kit (or alternative materials), construct models of the compounds whose structural formulas are listed in Results table 1 in your obook pro.

Results

1 Download the practical worksheet in your obook pro and fill in Results table 1.
2 Redraw the molecular structures of the compounds listed in Results table 1.
3 Predict whether each compound will ionise fully or partially in solution and write your prediction in Results table 1.
4 Predict whether each compound will form an acid, form a base, or remain neutral in solution and write your predictions in Results table 1.
5 If the compound is predicted to be an acid or a base, draw the structural formulas of the conjugate pair in Results table 1.

Discussion

1 List the acids. Justify your response by identifying the hydrogen that the compound can donate.
2 List the bases. Justify your response by identifying the atom in the compound that can accept a proton.
3 Identify the neutral compounds. Justify your response.
4 Use the internet to determine whether your classifications and the struture of your conjugate pairs were correct.

12.2 PRACTICAL: CONTROLLED EXPERIMENT

What happens when commercial indicators are added to different solutions?

Video demonstration
12.2 What happens when commercial indicators are added to different solutions?

Practical worksheet
12.2 What happens when commercial indicators are added to different solutions?

Risk assessment
12.2 What happens when commercial indicators are added to different solutions?

Context

Commercial indicators produce a visible colour change in acidic, basic or neutral solutions. Therefore, they can help us determine the pH of a solution.

In this practical, you will observe the colour changes of various commercial indicators in acidic, basic and neutral solutions.

FIGURE 1 The colour of the indicator can be compared to a scale.

Aim

To observe and record visible colour changes of commercial indicators in acidic, basic and neutral solutions

Materials

For the class to share:

- Paper towel
- Spare (clean) test tubes for mistakes
- Plastic containers or tubs (for used test tubes and chemical waste)

For each group of three students:

- Small, glass test tubes
- Test-tube rack
- Litmus paper (2 cm in length)
- 3 watch glasses or Petri dishes
- Dropper bottles of:
  - universal indicator solution
  - bromothymol blue solution
  - phenolphthalein solution
  - methyl orange solution
  - 0.1 M hydrochloric acid solution
  - 0.1 M sodium hydroxide solution
  - 0.1 M sodium chloride solution
- Heat mat (to prevent mess when using litmus paper)

Method

1 On a watch glass, add 3–5 drops of hydrochloric acid to one strip of litmus paper. Record the colour change in Results table 1.

2 Repeat step 1 with sodium chloride solution and then sodium hydroxide solution on fresh pieces of litmus paper. Record the colour changes in Results table 1.

3 Add 10 drops of hydrochloric acid to a test tube, and then add 3–5 drops of phenolphthalein solution. Record the colour change in Results table 1.

4 Repeat step 3 with sodium chloride solution and then sodium hydroxide solution. Record the colour changes in Results table 1.

5 Add 10 drops of hydrochloric acid to a test tube, and then add 3–5 drops of bromothymol blue solution. Record the colour change in Results table 1.

6 Repeat step 5 with sodium chloride solution and then sodium hydroxide solution. Record the colour changes in Results table 1.

7 Add 10 drops of hydrochloric acid to a test tube, and then add 3–5 drops of methyl orange solution. Record the colour change in Results table 1.
8 Repeat step 7 with sodium chloride solution and then sodium hydroxide solution. Record the colour changes in Results table 1.
9 Add 10 drops of hydrochloric acid to a test tube, and then add 3–5 drops of universal indicator solution. Record the colour change in Results table 1.
10 Repeat step 9 with sodium chloride solution and then sodium hydroxide solution. Record the colour changes in Results table 1.
11 For the clean-up, tip the contents of the test tubes in a plastic container (for off-site chemical waste disposal) and return the used test tubes to a plastic tub for washing. Wipe down the laboratory bench with a damp piece of paper towel and dispose of the paper towel in the rubbish bin.

Results

1 Download the practical worksheet in your obook pro and fill in the Results table.

Discussion

1 Identify the indicators that showed a visible colour change in acidic solutions.
2 Identify the indicators that showed a visible colour change in basic solutions.
3 Identify the indicators that were better for identifying a neutral solution.
4 Use the commercial indicator pH chart to predict the pH range of the acid, base and neutral solutions tested from your observations recorded in Results table 1. Compare this with the pH range of each indicator.

Inquiry: What if a pH meter was used?

1 Write a hypothesis for this investigation.
2 Compare the pH values obtained using a pH meter with the results using the commercial indicators. Discuss which method you think would give a more accurate result.
3 Discuss the advantages and disadvantages of using a pH meter. Compare this with the advantages and disadvantages of using commercial indicators.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors and suggestions for improvement.

13.3
PRACTICAL: LITERATURE REVIEW

Are redox reactions beneficial or harmful to society and the environment?

Practical worksheet
13.3 Are redox reactions beneficial or harmful to society and the environment?

Context

Redox reactions are common in nature. By understanding how these reactions work, we have been able to design useful redox reactions. Some redox reactions are also undesirable.

In this practical, you will research and review how redox reactions are used and evaluate their benefits and harms to society and the environment. This could include combustion reactions in fireworks, batteries, film photography, use of bleach, corrosion, breathalysers, use of antiseptics and composting.

FIGURE 1 Film photography uses redox reactions.

Aim

To investigate redox reactions and evaluate their benefit and harm to society and the environment

Instructions

1 Look for secondary sources of information. These can include the internet, books, scientific magazines, videos, podcasts, or interviews with an expert.

2 Use the CRAAP method to evaluate the reliability of your sources. This method is outlined in Topic 10.4, Step 5.

3 Make notes about what you have learnt. You can organise your information in different ways. Examples can be found in Topic 10.5, Step 6.

4 Prepare a written report that answers the following questions.
  - **a** What are the current uses of redox reactions? Which are natural reactions and which are designed by humans?
  - **b** How can chemistry knowledge affect how we use redox reactions?
  - **c** Select one redox reaction to explore in more depth. Write a balanced redox equation for the process involved.
  - **d** How does the redox reaction you chose in part **c** affect society and the environment? Is it harmful or helpful, neither or both?
  - **e** Is the redox reaction you chose in part **c** sustainable? What alternatives exist? What are their advantages and disadvantages?

  Tips for preparing a written report can be found in Topic 10.6, Step 8.

5 Make sure you record the details of all of the sources you use, including the title of the source, who it was written by, when it was written, page numbers or URLs, and the date you accessed it. You will need this information to put together a reference list (Topic 10.7, Step 9).

14.2

PRACTICAL:
CLASSIFICATION & IDENTIFICATION

Can you determine the solubility of molecules in water from their structure?

Practical worksheet
14.2 Can you determine the solubility of molecules in water from their structure?

Context

The solubility of covalent molecules in water can be predicted from their polarity. Depending on the shape of a molecule, the dipoles can cancel each other out. In other cases, they do not cancel; instead, they have a permanent dipole that makes them polar.

You can see this in the water molecule (Figure 1). A dipole forms between the oxygen and hydrogen because the oxygen is more electronegative and becomes partially negative. The hydrogen becomes partially positive. Because water is a bent molecule and the dipoles point in the same direction, they do not cancel out. Therefore, water has a permanent dipole and is polar.

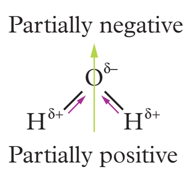


FIGURE 1 A water molecule is polar because its dipoles do not cancel out.

This means that water readily interacts with molecules that also have permanent dipoles or are polar (Figure 2). This allows these molecules to dissolve in water.

In this practical, you will analyse the structure of different molecules to predict their water solubility.

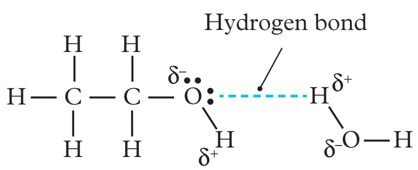


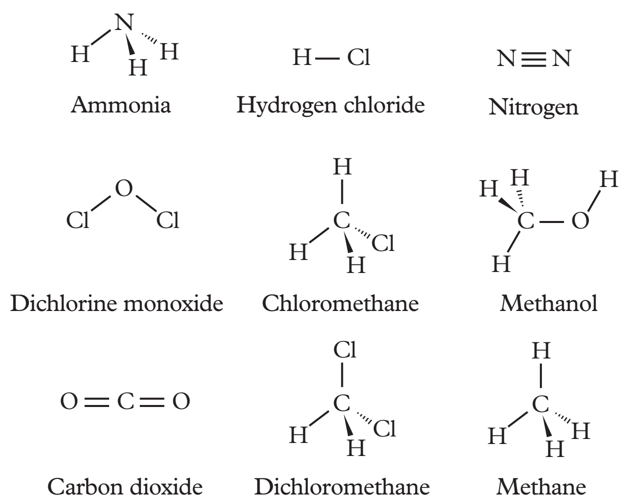
FIGURE 2 Ethanol is a polar molecule that can form hydrogen bonds with water. Therefore, ethanol is soluble in water.

Aim

To determine whether substances are polar and soluble in water from their structure

Questions

1 Copy the following molecules into your logbook.



2 Analyse the molecules and determine whether they are polar or non-polar. Assign dipoles where relevant.

3 Identify which of these molecules will be soluble in water. Justify your response.

4 For each of the molecules you have identified as soluble in water, draw diagrams to show the interaction between the molecule and water.

5 Describe the effect of changing temperature on the solubility of molecules in water.

14.3 PRACTICAL: PRODUCT, PROCESS OR SYSTEM DEVELOPMENT

How can you remove cations from hard water?

Practical worksheet
14.4 How can you remove cations from hard water?

Context

All water in Australia is sourced from rainwater. Deposits of rainwater form as rivers, lakes and groundwater etc. As rain falls, it absorbs CO_2 from the atmosphere. The water then passes over soil and rocks, where minerals are dissolved into the water. These minerals are predominantly carbonates of calcium and magnesium. They can also be chlorides, sulfates and other metal cations such as copper, lead or zinc.

FIGURE 1 Hard water contains metal cations.

Hardness is a measure of the concentration of calcium and magnesium cations in water. Therefore, the more calcium and magnesium cations that are dissolved in water, the harder the water is. In Australia, this concentration is expressed in mg L^{-1} or ppm. Table 1 shows a grading system for hardness based on calcium carbonate ($CaCO_3$) concentration.

TABLE 1 Grading system for degrees of hardness of water

Concentration of $CaCO_3$ (mg L^{-1})	Hardness
<60	Soft but possibly corrosive
60–200	Good quality
200–500	Increasing scaling problems
>500	Severe scaling

Water obtained from surface water catchments in Australia can have $CaCO_3$ levels as low as 14 mg L^{-1}. In groundwater, they can be as high as 440 mg L^{-1}.

The easiest test to determine whether water is hard is to observe whether it forms a lather with soap. If the water is hard, a lather will not form. When people shower in hard water and shampoo their hair, they often report that their hair feels unclean or oily. This is because the shampoo has not formed a lather to remove oil from the hair and thoroughly clean it.

When heated, minerals deposit as lime scale ($CaCO_3$). You might have noticed this as white residue on the inside of pipes, kettles and hot water cylinders or around taps. This build-up can be removed by vinegar (ethanoic acid). The weak acid reacts with the basic calcium carbonate molecule according to the equation:

$$CaCO_3(s) + 2CH_3COOH\ (aq) \rightarrow (CH_3COO)_2Ca(aq) + CO_2(g) + H_2O(l)$$

This reduces the hardness of the water.

In this practical, you will design a procedure to remove mineral cations from Victorian hard water.

Aim

To develop a procedure for the effective removal of mineral cations from hard water

Instructions

1 Analyse the water sample that you will use. This may have come from a water catchment in a Victorian region. Look up the water quality for this region and determine the theoretical hardness of the sample. This may be found on the water website for your region.

2 Brainstorm the chemicals that can be added to the hard water to remove the metal cations.

3 Create a flow chart outlining the procedure that you will use to remove cations from the hard water. Note: The precipitate must be removed after each experiment, dried and weighed. This must be included in your procedure.

4 Explain how you would determine the hardness of your water sample to compare it to the theoretical value in order to determine whether it is within Australian standards.

Questions

1 Explain why it is important to understand which metals cations are present in the water sample.

2 Identify the metal cations in your water sample. Explain how you know.

3 Write balanced chemical equations, including states, for the reactions that you have used in your procedure.

4 Compare your procedure with those of other students in your class. Identify any differences. Evaluate their methods and deduce which one would be the most effective to achieve your aim.

5 Research the chemical EDTA and explain how water-quality chemists are able to use it to determine the hardness of water.

15.2A

PRACTICAL:
CONTROLLED EXPERIMENT

Do more expensive vinegars have a higher concentration of ethanoic acid?

Video demonstration
15.2A Do more expensive vinegars have a higher concentration of ethanoic acid?

Practical worksheet
15.2A Do more expensive vinegars have a higher concentration of ethanoic acid?

Risk assessment
15.2A Do more expensive vinegars have a higher concentration of ethanoic acid?

Context

Ethanoic acid is found in vinegar; it is produced by the oxidation of ethanol into a carboxylic acid, and it gives vinegar some of its distinct flavour. You might have noticed at the supermarket that some vinegars are more expensive than others. Could this be because they have a higher concentration of ethanoic acid in them?

In this practical, you will determine and compare the concentrations of ethanoic acid in different vinegars by using an acid–base titration and stoichiometry calculations.

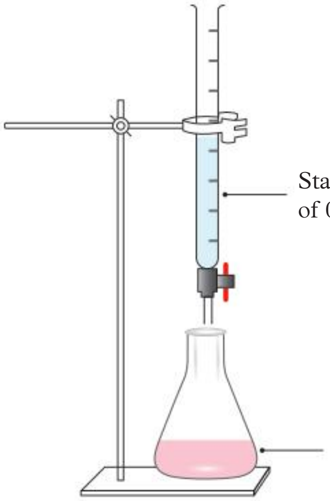


FIGURE 1 The set-up for the titration

Aim

To determine and compare the concentration of ethanoic acid in different vinegars

Materials

- Various vinegar samples (different prices)
- 0.1 M NaOH solution
- Deionised water in a wash bottle
- 25 mL pipette and pipette filler
- Burette
- Retort stand
- Boss head clamp
- White tile
- Conical flasks
- Phenolphthalein indicator
- Analytical balance
- 200 mL volumetric flask
- Stopper
- 3 × 100 mL transfer beakers
- Marker
- 250 mL waste beaker
- Funnel
- 20 mL pipette

Method

1. Weigh approximately 20 g of vinegar on the analytical balance. Record the actual mass.
2. Add the vinegar to a clean 200 mL volumetric flask and top up to the line with deionised water. Add a stopper to the volumetric flask and invert the flask to mix thoroughly.
3. Wash the pipette with water and then with diluted vinegar solution, emptying each into the waste beaker. Pipette a 25.00 mL aliquot of the vinegar into a 100 mL conical flask. Add three drops of phenolphthalein indicator to the conical flask.
4. Place a funnel in the top of the burette. Wash the burette with water and then with the 0.1 M NaOH solution. Fill the burette with the standard 0.1 M NaOH solution. Record the initial burette reading to two decimal places.

5 Place the conical flask on a white tile beneath the burette and titrate until the solution just becomes permanently pale pink.
6 Record the final burette reading in a results table.
7 Repeat the titration (steps 2–4) until three concordant results have been obtained. This means that the highest and lowest titre volumes should be within 0.10 mL of each other.
8 Repeat steps 1–7 with different samples of vinegar or use results from other groups.

Results

1 Record your results in a results table.
2 Calculate the average titre volume of NaOH that you used for the titration.
3 Calculate the amount, in mol, of NaOH used in the average titre and then use the mole ratio to determine the amount of ethanoic acid in the sample.
4 Calculate the mass of ethanoic acid in the vinegar.
5 Using the original mass of vinegar you used to make up your vinegar solution and the mass of ethanoic acid in the sample, calculate the concentration as % (m/m) of the ethanoic acid in the vinegar.

Discussion

1 Write a balanced chemical equation for the reaction between ethanoic acid and sodium hydroxide.
2 Compare the concentrations of each of the vinegar samples with the cost of the sample of vinegar. Does a higher price indicate more ethanoic acid? What is the trend?
3 Identify and discuss any errors you made during this experiment.

Inquiry: Does the source of the vinegar have an effect on the concentration of ethanoic acid?

There are many types of vinegars, including red wine, white wine, apple cider, malt, brown, and white vinegar.

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify two materials or pieces of equipment that require a risk assessment.
4 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.
5 Identify one assumption you made that could affect your results.
6 Discuss the results of your investigation.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors, suggestions for improvement and a statement of whether there is a relationship between the price of vinegar and the concentration of ethanoic acid.

15.2B

PRACTICAL: LITERATURE REVIEW

How have methods for measuring pH changed over time?

Practical worksheet
15.2B How have methods for measuring pH changed over time?

Context

pH is important for the safety, stability and quality of many chemical products and environments. This includes food products, personal care products and cosmetics, drinking water, agriculture, medicines, in pools and fish tanks. There are several methods or tools that can be used to measure and monitor pH.

In this practical, you will research and review the history of tools and methods for measuring the pH of different substances.

FIGURE 1 pH meters are used to measure the pH of marine aquariums.

Aim

To investigate the tools and methods for measuring pH and identify reasons why they have or have not changed over time

Instructions

1 Look for secondary sources of information. These can be the internet, books, scientific magazines, videos, podcasts, or interviewing an expert.

2 Use the CRAAP method to evaluate the reliability of your sources. This method is outlined in Topic 10.4, Step 5.

3 Make notes about what you have learnt. You can organise your information in different ways. Examples can be found in Topic 10.5, Step 6.

4 Prepare a written report that answers the following questions.

   a What are the different tools and methods used to measure pH? How do they work?

   b What types of samples are each of these tools and methods used for?

   c How have these tools and methods changed over time?

   d What are the limitations of the existing tools and methods?

   e Do we need better methods for pH measurement? Why or why not?

   Tips for preparing a written report can be found in Topic 10.6, Step 8.

5 Make sure you record the details of all of the sources you use, including: the title of the source, who it was written by, when it was written, page numbers or URLs and the date you accessed it. You will need this information to put together a bibliography (Topic 10.7, Step 9).

16.2A PRACTICAL: CASE STUDY

How do hot-air balloons use our understanding of gas behaviour?

Practical worksheet
16.2A How do hot air balloons use our understanding of gas behaviour?

Context

The idea of using hot gases to make objects lift into the air dates back to ancient China (200–300 CE) when military personnel would release paper lanterns as a form of communication. In this practical, you will investigate the history of hot-air balloons and draw connections to chemistry concepts.

The first hot-air balloon was developed by French brothers Joseph-Michel and Jacques-Etienne Montgolfier in 1783 (Figure 1). To inflate the balloon and lift it off the ground, they used a basket of hot coals to heat the air. It flew for 10 minutes, unmanned, and laid the foundation for the design of future manned hot-air balloons.

FIGURE 1 A model of the first hot-air balloon developed by the Montgolfier brothers

The first manned hot-air balloon flight followed shortly after.

At about the same time, hydrogen (H_2) gas was discovered, and French inventor Jacques Charles formed the hypothesis that this light gas could also lift a balloon. Charles enlisted the help of French engineers and brothers Anne-Jean and Nicolas-Louis Robert to build the balloon (Figure 2). The design included a release valve to expel gas from the balloon into the atmosphere.

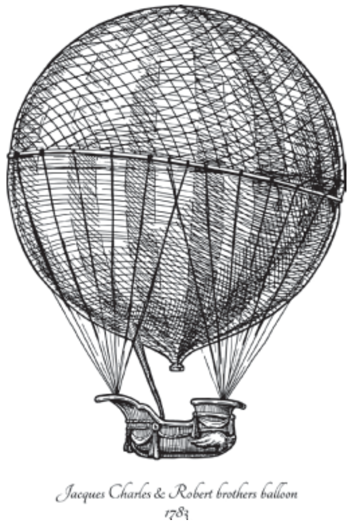


FIGURE 2 The hydrogen balloon designed by Jacques Charles and the Robert brothers

To generate the hydrogen required to fill the balloon, Charles used the reaction between an acid and a metal to produce salt and hydrogen. He used 250 kg of dilute sulfuric acid (H_2SO_4) and 500 kg of scrap iron:

$$H_2SO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + H_2(g)$$

This reaction released hot hydrogen gas, making it challenging to fill the balloon. As the gas cooled, the volume of the balloon decreased. Charles needed to produce more hydrogen gas to fill the space. This led to the development of Charles' law,

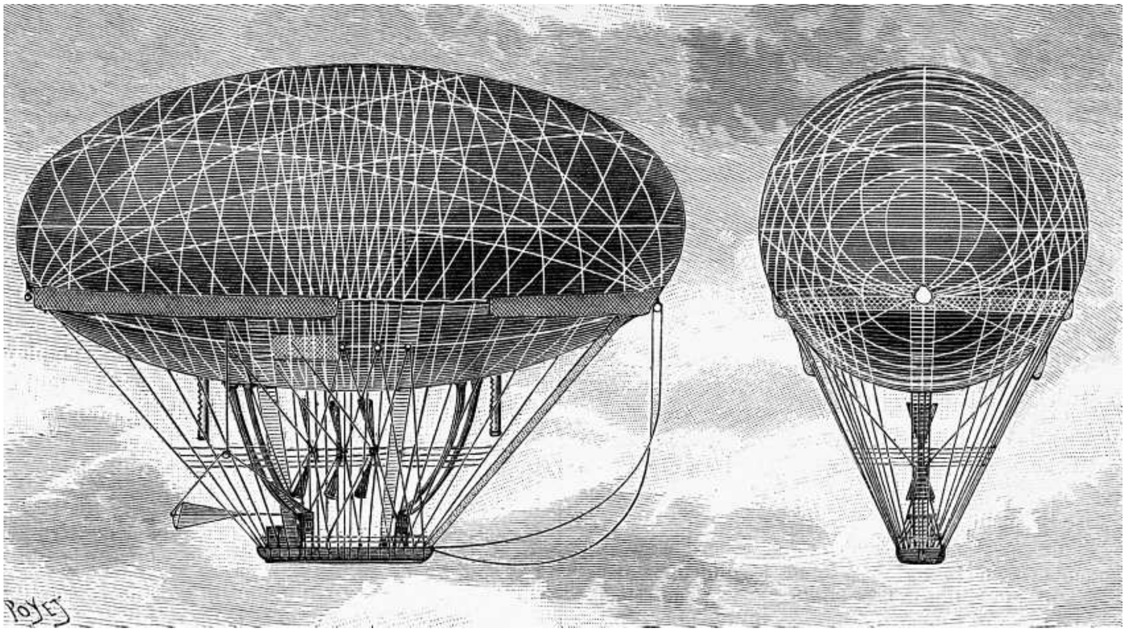


FIGURE 3 One of the early designs of Meusnier's dirigible

which describes the relationship between the temperature of a gas and the volume of space that it occupies.

The flight of Charles' hydrogen gas balloon was highly successful and led the team to build an elongated balloon. This was modelled from French mathematician Jean Baptiste Meusnier's design of the first dirigible (Figure 3).

Aim

To explain how hot-air balloons work by using your understanding of gas behaviour

Questions

1 Explain how hot-air balloons can rise through the atmosphere. Refer to density and the kinetic molecular theory in your answer.

2 Describe what would need to happen for a hot-air balloon to descend.

3 Explain the advantages and disadvantages of using hydrogen gas to fill the balloon.

4 Propane gas is used in modern hot-air balloons. Conduct some research to answer the following questions.

   a Identify the type of reaction involved in inflating a modern hot-air balloon.

   b Write a balanced chemical equation for the reaction of propane in a hot-air balloon.

   c Describe how modern hot-air balloons work.

   d Suggest a reason for the replacement of hydrogen gas with propane in hot air balloons.

   e Compare the applications of modern hot-air balloons and hydrogen hot-air balloons, by describing their purpose.

   f Discuss the impact of modern hot-air balloons on the environment.

   g Suggest strategies that could be used to make the use of hot-air balloons 'greener'. Identify the green chemistry principles you could address.

16.2B

PRACTICAL:
CONTROLLED EXPERIMENT

How are pressure and volume related in a variable volume system?

Video demonstration
16.2B How are pressure and volume related in a variable volume system?

Practical worksheet
16.2B How are pressure and volume related in a variable volume system?

Risk assessment
16.2B How are pressure and volume related in a variable volume system?

Context

Boyle's law states that when pressure is increased in a variable-volume system, the volume of gas within the system decreases.

In this practical, you will use masses to apply pressure to a gas system and investigate the changes in volume.

FIGURE 1 A syringe can be used to measure volume changes.

Aim

To determine the relationship between pressure and volume in a variable volume system

Materials

- Wooden block Boyle's law apparatus (if one is unavailable, a cork can be placed in the end of a plastic syringe. The corked syringe must be held firmly against the bench using a retort stand and clamp to ensure that the cork does not come off the nose of the syringe.)
- 60 mL plastic syringe
- 0.5 kg and 1 kg masses

Method

1. Pull the plunger of the syringe until it measures the maximum volume (60 mL).
2. Place the syringe in the apparatus or the alternative (see Materials).
3. Record the volume of the syringe at a weight of zero.
4. Add the 0.5 kg mass to the end of the apparatus with the plunger. Record the volume of the syringe.
5. Continue to add masses, 0.5 kg at a time, recording the volume of the syringe each time.

Results

1. Record your results in a results table. The table must include pressure (as mass applied to the syringe), volume and inverse volume $\left(\frac{1}{V}\right)$.
2. Create a graph to demonstrate the relationship between pressure and volume.
3. Create a second graph to demonstrate the relationship between pressure and inverse volume.

Discussion

1. Explain why the plunger cannot be pushed to the end of the barrel of the syringe.
2. Draw a diagram to demonstrate the behaviour of gas molecules:
   - **a** before masses are added
   - **b** after a 0.5 kg mass is added
   - **b** after a 3 kg mass is added.
3. Use the graphs to explain the relationship between pressure and volume in a variable-volume system. Refer to the kinetic molecular theory in your answer.
4. Explain why a second graph was created. Was it useful to help you answer Question 3? Why or why not?

Inquiry: What if the size of the syringe is increased/decreased?

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify two materials or pieces of equipment that require a risk assessment.
4 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.
5 Identify one assumption you made that could affect your results.
6 Discuss the results of your investigation.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors, suggestions for improvement and a statement of the relationship between pressure and volume in a variable volume gaseous system.

16.3

PRACTICAL: SIMULATION

How do ideal gases behave in a fixed volume system?

Practical worksheet
16.3 How do ideal gases behave in a fixed-volume system?

Weblink
PhET Gas Properties simulator

Context

The behaviour of gases is summarised using the ideal gas equation:

$$PV = nRT$$

where P is pressure (kPa), V is volume (L), n is moles, R is the ideal gas constant (8.31 J mol^{-1} K^{-1}) and T is temperature (K). It is a combination of:

- Gay-Lussac's law
- Boyle's law
- Charles' law
- Avogadro's law.

In this practical, you will use a simulation to study how gases behave according to some of these laws in a fixed volume system.

Aim

To study the behaviour of gases using the PhET Gas Properties simulator

Instructions

Getting started

1 Click on the icon at the top of the page, or use the following link to access the PhET Gas Properties simulator: https://phet.colorado.edu/sims/html/gas-properties/latest/gas-properties_en.html

2 Select 'Ideal' to simulate gas behaviour at SLC. The instructions for using the simulator are shown in Figure 1.

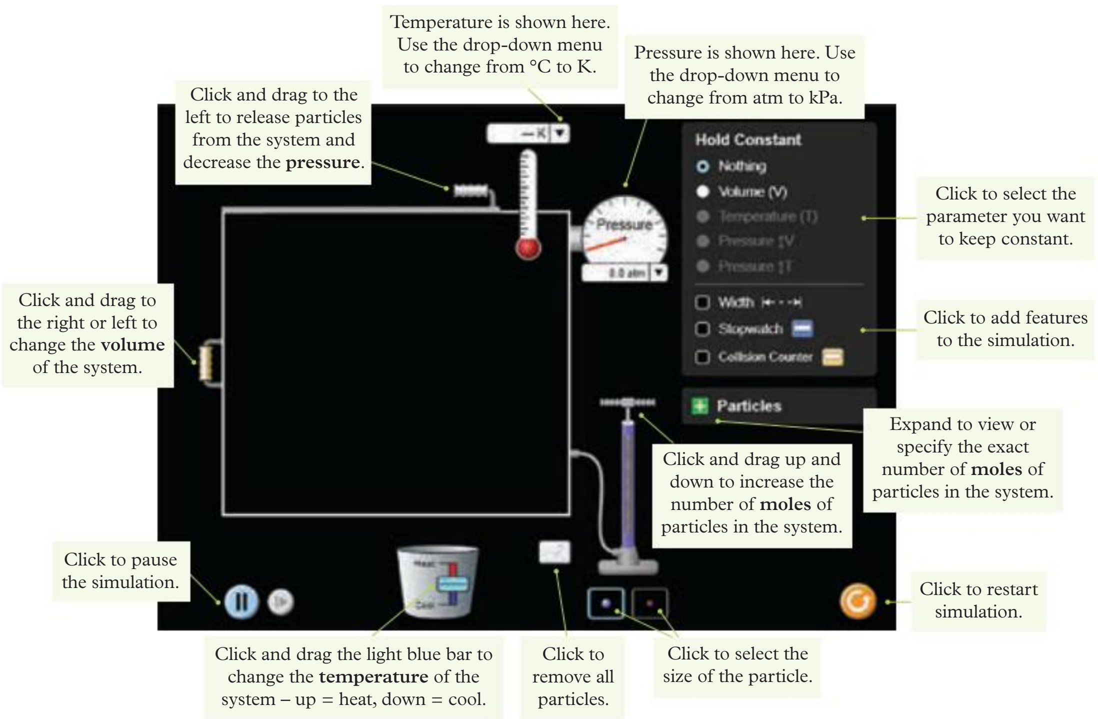


FIGURE 1 How to use the PhET Gas Properties simulator

3 Investigate the following gas laws using the simulator.
4 Download the practical worksheet in your obook pro and fill in the Results tables.

Part A – Gay-Lussac's law

Gay-Lussac's law describes the relationship between temperature and pressure. Here, you will change the temperature of the container and observe the effect on pressure.

1 Use the pump to add at least 300 particles to the system. Record the temperature (in K) and pressure (in kPa) in Results table 1.
2 Use the light blue bar to increase the temperature of the system by 20–50 K. Record the temperature and pressure in Results table 1.
3 Use the light blue bar to increase the temperature of the system by another 20–50 K. Record the temperature and pressure in Results table 1.
4 Repeat step 3 until you have five sets of temperature and pressure. Note: you can also drag the light blue bar down to decrease the temperature.
5 Reset the simulation.

Part B – Boyle's law

Boyle's law describes the relationship between pressure and volume. Here, you will change the volume of the system and observe the effect on pressure.

1 Use the pump to add at least 300 particles to the system. Add the 'Width' feature to the simulator. Record the width (in nm) and the pressure (in kPa) in Results table 2.
2 Use the handle to decrease the width of the container to 5.0 nm. Record the width and pressure in Results table 2.
3 Use the handle to increase the width of the container to 7.5 nm. Record the width and pressure in Results table 2.
4 Use the handle to increase the width of the container to 12.5 nm. Record the width and pressure in Results table 2.
5 Use the handle to increase the width of the container to 15.0 nm. Record the width and pressure in Results table 2.
6 Reset the simulation.

Questions

1 Create a graph to show the relationship between temperature and pressure. Describe your results.
2 Using the kinetic molecular theory, explain your result from Part A.
3 Create a graph to show the relationship between width (volume) and pressure. Describe your results.
4 Using the kinetic molecular theory, explain your result from Part B.
5 Charles' law describes the relationship between temperature and volume. You did not investigate Charles' law here.
   a Suggest a reason why.
   b Suggest one change you could make to the simulator to investigate Charles' law.

16.4 PRACTICAL: CONTROLLED EXPERIMENT

How can we measure the volume of gas produced in a reaction?

Video demonstration
16.4 How can we measure the volume of gas produced in a reaction?

Practical worksheet
16.4 How can we measure the volume of gas produced in a reaction?

Risk assessment
16.4 How can we measure the volume of gas produced in a reaction?

Context

Hydrogen peroxide decomposes very slowly into water and oxygen according to the chemical equation:

$$H_2O_2(aq) \rightarrow H_2O(l) + O_2(g)$$

To increase the rate of the chemical reaction, a catalyst is used. Some examples of catalysts are yeast, iron(III) chloride and potassium iodide.

In this practical, you will measure the volume of oxygen produced when hydrogen peroxide decomposes in the presence of a catalyst.

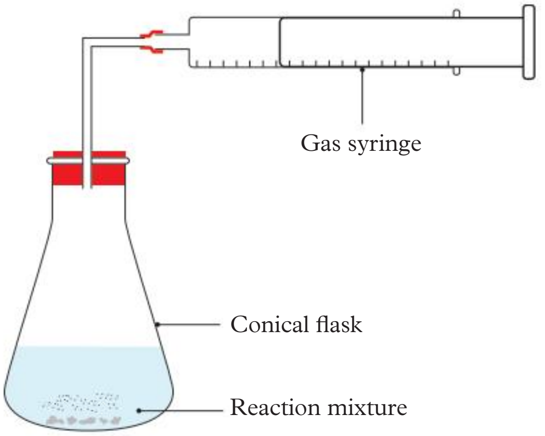


FIGURE 1 The experimental set-up for measuring the volume of a gas produced in a chemical reaction

Aim

To measure the volume of oxygen produced from the decomposition of hydrogen peroxide at SLC.

Materials

- 250 mL conical flask
- Rubber stopper with a gas outlet tube (and thermometer adaptor, if possible)
- Heatproof mat
- 10 mL plastic or gas syringe
- 50 mL measuring cylinder
- 50 mL of 30% v/v $H_2O_2(aq)$ (density = 1.11 g mL^{-1})
- Electronic balance
- Weigh boat
- 0.5 g potassium iodide (or iron(III) chloride)
- Thermometer

Method

1. Assemble the apparatus in Figure 1 and place the conical flask on the heatproof mat. Make sure that the gas syringe is pushed in all the way before beginning the experiment.
2. Calculate the theoretical volume of oxygen gas generated, assuming that the experiment is conducted at SLC. Based on the volume of oxygen, double check that the syringe is appropriate to measure this volume.
3. Use the measuring cylinder to measure 50 mL of 30% v/v H_2O_2 and transfer it into the conical flask. Record the temperature of the H_2O_2.
4. Place the weigh boat on the electronic balance and tare the balance so that it reads 0.00 g. Measure 0.5 g of KI into the weigh boat.
5. Add the KI into the conical flask and quickly replace the stopper. Allow the reaction to continue until no more bubbles are produced.
6. Record the final volume of the gas and the final temperature of the reacted H_2O_2.

Results

1. Record your results in a table.

Discussion

1 Compare the theoretical and experimental volume of oxygen gas produced. Explain any differences between the values.
2 Comment on the temperature of the experiment. Describe the impact that this has on the theoretical calculation of the volume of oxygen produced.
3 Evaluate whether the results are accurate. Justify your answer.
4 Evaluate the validity of the experiment. Justify your answer.

Inquiry: What if the concentration of the H_2O_2 was increased/decreased?

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify two materials or pieces of equipment that require a risk assessment.
4 Identify any sources of error that could affect your results. Suggest how you could change your method to minimise one of the errors.
5 Identify one assumption you made that could affect your results.
6 Discuss the results of your investigation.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors, suggestions for improvement and a statement of the volume of gas produced in the decomposition of hydrogen peroxide.

17.2A
PRACTICAL:
CONTROLLED EXPERIMENT

How is gravimetric analysis used to identify the mass of salt in a solution?

Video demonstration
17.2A How is gravimetric analysis used to identify the mass of salt in a solution?

Practical worksheet
17.2A How is gravimetric analysis used to identify the mass of salt in a solution?

Risk assessment
17.2A How is gravimetric analysis used to identify the mass of salt in a solution?

Context

Because salts are dissolved in a solution as ions, their mass and therefore concentration cannot be measured by simply filtering out the ions. Instead, they must first be precipitated from the solution as a solid.

In this practical, you will precipitate a metal sulfate salt from a sample of water using barium chloride then use gravimetric analysis and mass–mass stoichiometry to calculate the original mass and concentration of the metal sulfate.

Aim

To determine the mass and concentration of metal sulfate salt in a water sample, using a precipitation reaction and gravimetric analysis techniques

Materials

- 50 mL of metal sulfate salt solution (e.g. sodium sulfate, aluminium sulfate) (Note: The mass of the metal sulfate must not exceed 0.35 g or the $BaCl_2$ will no longer be in excess)
- 2 M HCl solution
- 0.1 M $BaCl_2$ solution
- 250 mL beakers
- Analytical balance
- Stirring rod
- 50 mL measuring cylinder
- Bunsen burner
- Heatproof mat
- Tripod
- Gauze mat
- Deionised water
- Filter funnel
- Filter paper
- Conical flask
- Watch glass
- Drying oven
- Deionised water in a wash bottle
- Plastic pasteur pipette
- 250 mL conical flask

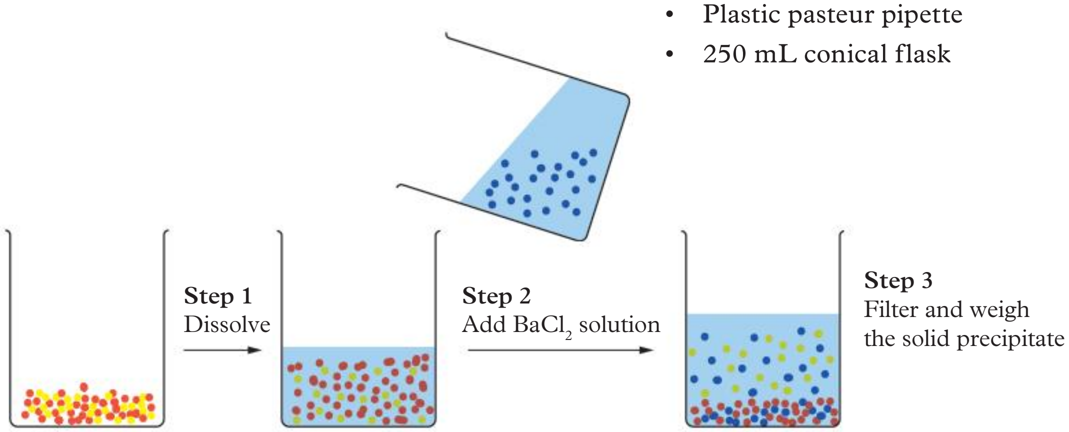


FIGURE 1 A salt can be precipitated from solution for gravimetric analysis.

Method

1 Collect a beaker containing a 50 mL sample of salt solution and note the exact mass of salt in it, as provided by your teacher. Add 20 drops of 2 M HCl and stir until everything is completely dissolved.

2 Use the Bunsen burner to heat the sample of salt, water and HCl. Stir the sample while heating and remove from the heat just before it boils.

3 Measure 25 mL of $BaCl_2$ solution in a measuring cylinder. This will be an excess amount of $BaCl_2$.

4 Very slowly add the $BaCl_2$ solution to the beaker containing the hot solution. A white precipitate will start to form. It should take you about 3 minutes to add the full volume into the beaker. This will ensure that large crystals form, which are easier to filter off and collect. Continually stir the contents of the beaker as you add the $BaCl_2$ solution.

5 Rinse down the sides of the beaker and stirring rod with deionised water. Let the precipitate solution sit for about 10 minutes while you set up the filtration.

6 Using an analytical balance, weigh a sheet of filter paper and filter paper + watch glass. Record the masses.

7 Set up the filtration using the filter paper with the known mass as shown in Figure 2.

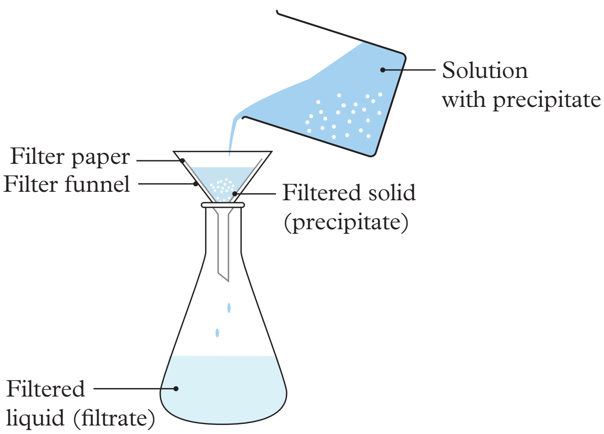


FIGURE 2 The solid precipitate can be filtered from the precipitate solution.

8 Pour the solution through the filter paper carefully, making sure you capture all the precipitate crystals in the filter paper. If the filtrate liquid looks to still have crystals in it, repeat the filtration using a new weighed piece of filter paper.

9 Carefully remove the filter paper containing the precipitate from the funnel. Lay it flat on a watch glass and place in the drying oven at 40°C.

10 Over the next couple of days, weigh the filter paper and watch glass together, and record the masses. Keep completing this step until a constant mass is achieved.

Results

1 Record your results in a results table.

2 Write a balanced chemical equation for the precipitation reaction between barium chloride and your metal sulfate salt.

3 Using the final mass of precipitate, calculate the amount (n) of precipitate formed in your experiment.

4 Using the stoichiometric ratio from the equation in Question 2, calculate the amount (n) of the metal salt in your solution.

5 Calculate the mass of salt in your 50 mL sample solution.

6 Calculate the concentration of salt in your 50 mL sample solution.

Discussion

1 Compare your answer with the correct mass of salt.

2 Identify and discuss any errors you made during this experiment.

3 Suggest a way you could change your method to minimise one of the errors.

4 Evaluate whether the results are accurate. Justify your answer.

5 Evaluate whether you could use a precipitation reaction and gravimetric analysis to identify an unknown salt.

Inquiry: Could you use a precipitation reaction and gravimetric analysis to identify an unknown salt?

1 Write a hypothesis for this investigation.
2 Identify the *independent* variable that you will change from the first method.
3 Identify two materials or pieces of equipment that require a risk assessment.
4 Assuming the unknown salt is a sulfate salt, suggest a precipitation reaction you could use to form a solid that you can then filter and analyse using gravimetric analysis.
5 Write a balanced chemical equation for the precipitation reaction in Question 4.
6 Discuss whether you will need to modify the method you used for gravimetric analysis of the known salt.
7 Outline the steps you would need to follow to determine the identity of the unknown salt.

Conclusion

Write a conclusion for your experiment that summarises your results, and includes an analysis of possible errors, suggestions for improvement and a statement of the mass and concentration of the salt in your sample solution.

17.2B

PRACTICAL: CASE STUDY

What can we do if soil salinity is too high?

Practical worksheet
17.2B What can we do if soil salinity is too high?

Context

The salt content or salinity of soil is very important in agriculture for seeds to germinate properly and grow into healthy plants. If the salinity is too high, plants may grow slowly, or even wilt and die. Water in the soil will remain in the soil instead of being absorbed into the plant's roots. Accumulation of elements such as sodium in the plant cells can also be harmful to the plant.

In this practical, you will examine the scenario below and play the role of a soil scientist to provide recommendations to a farmer.

You, a soil scientist, have been invited to a local farm to investigate a problem with the growth of their corn. The farmer tells you that the plants are much shorter and are not yielding much corn this season. You suspect this has to do with high salinity of the soil, since you recently visited a nearby farm that had soil containing a higher than normal salt content and was experiencing the same problems.

FIGURE 1 A sample of soil

Aim

To evaluate a situation where high soil salinity may be affecting plant growth and provide recommendations to address the problem

Questions

1. Explain why soil salinity is important in agriculture.
2. Outline how you could test the salinity of the soil. Include the materials and equipment required, and the steps you would take to perform the tests.
3. Identify the possible causes of increased soil salinity.
4. Discuss how human activity could be modified to prevent further increases in soil salt content.
5. Identify the other characteristics of the soil you could test to learn more about the health and quality of the soil.
6. Discuss the impacts of poor crop yield on the:
   - a farmer
   - b environment
   - c consumers.
7. Use the internet to research strategies to reduce the salinity of soil. Summarise this information and present it as a pamphlet you could distribute to the farmer and any neighbouring farms experiencing the same problem. In your pamphlet, make sure you include the following information, written in a way that people without scientific knowledge can understand:
   - a description of soil salinity and its impact on crops
   - the causes of increased soil salinity
   - two or three strategies that the farmer could use to decrease soil salinity, including a brief background on how they work, a list of materials or equipment required, and the method to be used.

ANSWERS

You can find the full worked solutions to the answers in your **Student obook pro**

Chapter 1: Chemistry toolkit

1.1 Overview of VCE Chemistry

GROUNDWORK

1A A hypothesis is a testable statement that predicts how the independent variable will affect the dependent variable in an investigation.

1B Qualitative data is data described in words, phrases or categories (e.g. colour of the solution). Quantitative data is data described using numbers, quantities or other numerical values (e.g. concentration of NaCl).

1C Student answers will vary but examples include: title, units, columns, *x*- and *y*-axes, independent variable, dependent variable.

1D A controlled variable is one that is kept the same during an experiment. An independent variable is what you change – it is the variable being investigated. A dependent variable is what you measure in an experiment.

1.1 CHECK YOUR LEARNING

1 Areas of Study 1 and 2: one of the assessment tasks outlined in the list (e.g. a report of laboratory or fieldwork activity); Area of Study 3: response to an investigation question with links to sustainability (Unit 1) or a report of a student-adapted/designed investigation (Unit 2).

2 Student answers will vary.

3 **a–e** Student answers will vary.

4 Student answers will vary.

1.2 Aboriginal and Torres Strait Islander knowledge, cultures and histories

1.2 CHECK YOUR LEARNING

1 Student answers will vary.

2 Aboriginal and Torres Strait Islander Peoples have developed and refined their own chemical knowledge over thousands of years; it is highly sophisticated and specialised for the specific location where they live.

3 Student answers will vary but should include how plants are used as medicine, and how natural materials are used and modified for a particular purpose.

4 Student answers will vary.

1.3 Developing aims, questions and hypotheses

1.3 CHECK YOUR LEARNING

1 The scientific process is a cycle in which the results are used as feedback to perform additional tests or make new hypotheses.

2 Aim: a statement of what is to be investigated. Hypothesis: a testable statement that predicts how the independent variable affects the dependent variable in the investigation, including an explanation.

3 **a** Student answers will vary but should include a clear, concise statement about what they are investigating and provide enough scope for experimentation.

b Student answers will vary but should include a statement of what is to be investigated.

c Student answers will vary but should include a testable statement that predicts how the independent variable will affect the dependent variable, and include a scientific explanation for the prediction.

1.4 Planning and conducting investigations

1.4 CHECK YOUR LEARNING

1 Provides a useful record of primary and secondary data for scientific investigations

2 Primary data: raw data collected directly from investigations or experiments. Secondary data: data that has already been collected, analysed and interpreted.

3 **a** Similarity: Both can involve testing an independent variable to observe its effects on the dependent variable. Difference: Controlled experiments are typically conducted in a lab, whereas fieldwork is undertaken at a specific location.

b Similarity: Both involve looking at information to solve a problem. Difference: Case studies involve solving problems using knowledge you already have or information that has been given to you. Literature reviews require you to look for secondary data to answer a question.

c Similarity: Both involve investigation of a real system or scientific phenomena using visualisation tools and can be used to make predictions; Difference: Modelling requires you to create a model, whereas simulation involves using existing models to understand behaviour of systems.

4 Student answers will vary but should include a written question that clearly explains the chemical issue being investigated and the experimental approach being used.

1.5 Safety in chemistry

1.5 CHECK YOUR LEARNING

1 Chemists often work with dangerous goods and hazardous substances; therefore, it is important to follow safety rules and conduct risk assessments for experimental investigations in the laboratory.

2 Student answers will vary but should include an evaluation of each section of the risk assessment before making a final risk judgment of the laboratory-based activity.

3 Similarities: potential hazards and standard handling procedures are listed. Differences: hazard warning symbols and a classification of the type of hazardous substance are provided for GHS substances classified as hazardous.

4 Student answers will vary but should include who has prepared the risk assessment, the title of the investigation, who is conducting the investigation, the date and location, a list of materials, procedure, equipment and chemicals used, and a final risk judgment with a signature.

1.6 Ethical understanding

1.6 CHALLENGE

1 Student answers will vary.

2 Student answers will vary but should identify issues such as the impact or harm of the issue (on living things or the environment), integrity, safety and the needs of society in the present or future.

3 Student answers will vary but could include political, legal, economic or social factors.

4 Student answers will vary.

5 Student answers will vary.

6 Student answers will vary.

1.6 CHECK YOUR LEARNING

1 To make sure you the broader implications of conducting the investigation are considered, reporting of results is honest and all contributors to the ideas are acknowledged.

2 Student answers will vary but should include a discussion of the similarities and differences.

3 Student answers will vary but should include an explanation of arguments for and against the ethical issue, as well as justification for their opinion or decision on the issue.

1.7 Generating, collating and recording data

1.7 CHECK YOUR LEARNING

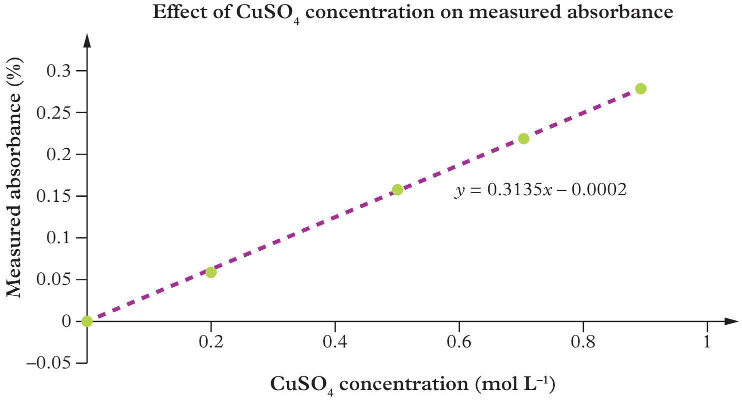
1 A title, labelled axes, the IV plotted on the *x*-axis, the DV plotted on the *y*-axis, appropriate axis scales, a key (if applicable)

2 Column graphs display categorical data, with bars that are plotted vertically or horizontally but not touching. Histograms show the frequency of numerical data as a percentage; each column width can be a single data value or a data interval.

3 **a** As a control.

b Dependent variable: measured absorbance. Independent variable: $CuSO_4$ concentration

c Student answers will vary but should present the data using a line graph.



As $CuSO_4$ concentration increases, the measured absorbance increases.

d 0.6 mol L^{-1} (1 sig fig)

1.8 Evaluating data and investigations

1.8 CHECK YOUR LEARNING

1 Systematic errors

2 It reduces the impact of random errors in the measurement process. By conducting additional trials, more data can be obtained and therefore the data can be averaged.

3 **a** Accuracy: how close a value is to the true value. Precision: a measure of how close the data values in a set are to each other.

b Repeatability: the data values can be produced again (by the same experiment and under the same conditions). Reproducibility: the same data values can be produced again (under slightly different conditions).

c Mistakes: personal errors that should not be included in reporting or data analysis. Errors: can be random (unpredictable, occur because of an error in the measurement process) or systematic (consistent and repeatable, are caused by a problem with the method or equipment).

4 The data obtained is not accurate because the values are not close to the accepted boiling point of water. The values in the data set lack precision.

5 **a** 4.08

b 36.0

c 2.90

6 Student answers will vary.

1.9 Constructing evidence-based arguments and conclusions

1.9 CHECK YOUR LEARNING

1 So that the results obtained can be shared and evaluated by scientists before the new knowledge is added to the existing body of scientific knowledge.

2 Use scientific terminology and/or supporting figures, make evaluative statements, use formal language, avoid first person.

3 Student answers will vary but should include an analysis of the experimental results, errors, inconsistencies and outliers that may have affected the results, scientific explanations that link the results to the question and aim and suggestions for improvement.

1.10 Communicating

1.10 CHECK YOUR LEARNING

1 **a** Identify and understand your target audience, use appropriate language, science literacy and convention, essential content and stylistic elements to enhance presentation.

b Using stylistic elements (e.g. diagrams, photos, graphs, tables) so that the content is clearly conveyed, easily understood and enhances written content

2 Scientific language and literacy should be used throughout a poster so that a summary of the research is communicated clearly with the intended audience.

3 Student answers will vary but should include a discussion of the similarities and differences.

4 Results: presents the raw data collected during the investigation. Discussion: examines whether the data supports or refutes the hypothesis, analyses the data and compares to expected results, identifies ambiguities and further questions that may arise and explains issues with the data.

5 Student answers will vary but should include important information and be written for the intended audience.

1.11 Sustainability

1.11 CHECK YOUR LEARNING

1 Atom economy, catalysis, design for degradation, design for energy efficiency, designing safer chemicals, prevention of waste and use of renewable feedstocks

2 A linear economy removes resources from the Earth to make products that are then disposed of, whereas a circular economy is a sustainable approach that focuses on using and reusing resources repeatedly at different stages of production and consumption.

3 Student answers will vary but should include a discussion about which of the

seven principles of green chemistry were important when considering replacing the hazardous substance.

1.12 Balancing chemical equations

1.12 CHECK YOUR LEARNING

1 The law of conservation of mass states that in an isolated system, during a chemical reaction, mass is conserved; atoms cannot be created or destroyed. The mass of reactants and products remains constant, and the number of atoms present does not change.

2 **a** $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

b $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

c $Ag_2CO_3(s) \rightarrow Ag_2O(s) + CO_2(g)$

d $C_2H_4(g) + 2O_2(g) \rightarrow 2CO(g) + 2H_2O(g)$

e $P_4(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$

f $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

g $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$

h $Ca(CN)_2(s) + 2HCl(aq) \rightarrow 2HCN(aq) + CaCl_2(aq)$

i $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

j $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

3 Student answers will vary.

1.13 Preparing for assessment

1.13 CHECK YOUR LEARNING

1 Use writing time effectively, note down ideas once writing time has commenced, use writing space provided, return to difficult questions later rather than dwelling on them, eliminate incorrect answers for multiple-choice questions, beware of careless mistakes, draw large, clear diagrams when required.

2 It is important to have a work–life balance; to have time to study, participate in extra-curricular activities and rest.

3 Student answers will vary but should include a goal that is clearly defined, can be measured, and is achievable, worthwhile and time-based.

4 **a** Derive: to manipulate or give a new relationship (i.e. obtain an equation). Calculate: to find a numerical answer (i.e. when solving a mathematical problem).

b Label: to show exactly where a feature is; Identify: only requires you to determine a characteristic.

c Establish criteria: to determine measures to conduct an evaluation; Evaluate: to apply a weighted criteria to the relative strengths and weakness of the arguments.

5 Student answers will vary.

Chapter 1 review

MULTIPLE CHOICE

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

SHORT ANSWER

9 Student answers will vary.

10 Use connectives to show cause and effect, use scientific terminology and/or supporting figures, use supporting facts and/or quotes from experts or other researched sources, make evaluative statements, use formal language and avoid third person.

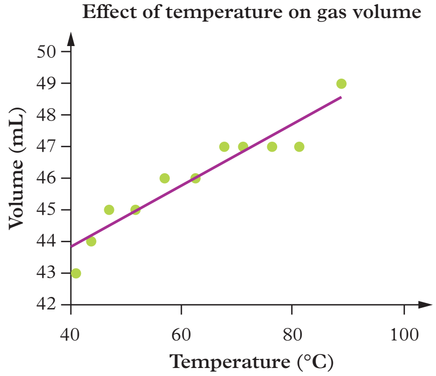
11 **a** Dependent variable = volume (mL). Independent variable = temperature (°C)

b To determine the effect of temperature on the volume of a gas

c If temperature increases, then the volume of a gas increases because as the particles gain more kinetic energy, they will collide with the walls of the container with more force and cause an increase in volume.

d Student answers will vary but should present the data as a line graph.

Effect of temperature on gas volume



Students should note that as temperature increases, the volume increases.

12 Student answers will vary but should include an analysis of the experimental results, errors, inconsistencies and outliers that may have affected the results, scientific explanations that link the results to the question and aim and suggestions for improvement.

13 Student answers will vary but should include who has prepared the risk assessment, the title of the investigation, who is conducting the investigation, the date and location, a list of materials, procedures, equipment and chemicals used, and a final risk judgment assigned with a signature.

Chapter 2: Elements and the periodic table

GROUNDWORK

2A An element is a pure substance that consists of only one type of atom.

2B Protons and neutrons are in the nucleus, electrons orbit outside the nucleus.

2C The periodic table is an organisational tool in which elements are arranged in order of increasing atomic number.

2.1 Elements, isotopes and ions

2.1 CHECK YOUR LEARNING

1 **a** Positive

b Neutral

c Negative

2 Protons and neutrons are in the nucleus. Electrons orbit the nucleus in electron shells.

3 The nucleus contains protons and neutrons. Neutrons are neutral; therefore, the nucleus has an overall positive charge due to its protons.

4 **a** $^{14}_{7}N$

b $^{23}_{11}Na$

c $^{82}_{36}Kr$

d $^{25}_{12}Mg$

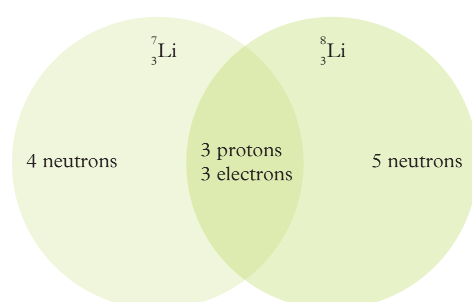
5 Approximately 1837, because the mass of an electron is approximately 1837 times smaller than the mass of a proton.

6 **a** $^{16}_{8}O^{2-}$: 8 protons, 8 neutrons, 10 electrons

b $^{27}_{13}Al$: 13 protons, 14 neutrons, 13 electrons

c $^{79}_{35}Br^{-}$: 35 protons, 44 neutrons, 36 electrons

7



2.2 The periodic table

2.2 CHECK YOUR LEARNING

1 The 3d subshell contains 5 orbitals and can hold a maximum of 10 electrons; therefore, there are 10 groups in the d-block.

2 Metallic character describes how readily an atom loses a valence electron. Non-metallic character describes how readily an element accepts electrons.

3 The core charge increases across the period, which causes electrons to be held more tightly to the nucleus, resulting in a smaller atomic radius.

4 **a** Sodium
  b Phosphorus
  c Bromine

5 They are both in group 1 and have one electron in their outer shell.

6 K > Mg > Si > F

7 K^+ $1s^2 2s^2 2p^6 3s^2 3p^6$, Ar $1s^2 2s^2 2p^6 3s^2 3p^6$, Cl^- $1s^2 2s^2 2p^6 3s^2 3p^6$. They all have the same electron configuration.

8 Zn^{2+} $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$, Cu^+ $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$. They both have the same electron configuration.

9 Answers should discuss any three of:
- atomic radius increases
- electronegativity decreases
- ionisation energy decreases
- metallic character increases

2.3 Critical elements

2.3 CHECK YOUR LEARNING

1 Student answers will vary; for example:
- Helium – used in balloons and aircrafts, but its primary use is to cool substances to extremely low temperatures and create controlled environments.
- Indium – used to make LCDs, LEDs, infrared lasers (for fibre optic cables) and some types of solder and glass.

2 Any of:
- Recycling indium from electronic devices is extremely difficult because it is usually fused with other elements/polymers making it very difficult to separate.
- The amount of indium that can be recovered is usually too small to make it worthwhile at current prices.

3 It enables the endangered elements to be recovered and recycled so that critical elements in limited supply can be conserved.

4 Student answers will vary but should include at least one possibility that could be initiated to encourage people to recycle their old electronic devices.

5 Student answers will vary but should include at least one possibility that could be developed to encourage manufacturers of electronic devices to recover endangered elements from the devices they collect through their recycling program.

Chapter 2 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 **a** The tendency for an atom to attract electrons when in a chemical bond
  b Describes how readily an atom loses a valence electron
  c The amount of energy required to remove one valence electron

12 **a** 4+
  b 1+
  c 2+
  d 4+
  e 5+

13 Chromium and copper do not follow the Aufbau principle because $3d^4$ and $3d^9$ are unstable electron configurations in neutral atoms. One electron from the 4s subshell in chromium and copper moves to occupy the 3d subshell. This results in $3d^5$ and $3d^{10}$ electron configurations.

14 **a** [Ne] $3s^2 3p^2$
  b [Ar] $4s^1 3d^{10}$
  c [Ar] $4s^2 3d^{10} 4p^4$
  d [Ar] $4s^2 3d^6$

15 They have the same number of valence electrons; therefore, they have similar chemical properties.

16 They have the same number of electron shells.

17 A re-liquefier can be used that captures and condenses 95% of the helium gas that can then be reused.

18 Despite both being in group 1 of the periodic table, sodium has 3 electron shells while lithium has 2 electrons shells. The presence of more shells of electrons while the core charge remains constant makes the atomic radius larger.

19 Fluorine has a higher core charge than nitrogen (core charge F = 7+ and N = 5+). Increasing core charge causes electrons to be held more tightly to the nucleus, resulting in greater attraction between the nucleus and the valence shell of electrons.

20 Chlorine-35 has 17 protons, 18 neutrons and 17 electrons. Chlorine-37 has 17 protons, 20 neutrons and 17 electrons. Both chlorine isotopes have the same number of electrons, so they will have the same chemical properties. Chlorine-37 has a greater mass because it has 20 neutrons, whereas chlorine-35 has only 18 neutrons.

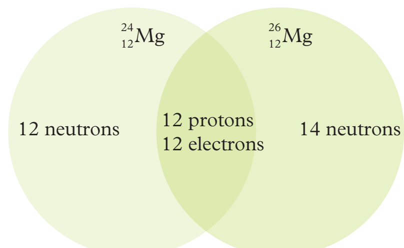
21 **a** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
  b $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$
  c $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$
  d $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^7$

22 **a** Li
  b Mg
  c Cl

23 **a** 24 protons, 30 neutrons, 21 electrons
  b 23 protons, 28 neutrons, 20 electrons
  c 14 protons, 16 neutrons, 10 electrons
  d 35 protons, 46 neutrons, 36 electrons

24 **a** Br, As, V, K
  b The core charge increases while the electron shielding remains constant. This causes the electrons to be attracted more strongly to the nucleus, resulting in a smaller atomic radius.

25 Student answers will vary but should include that a circular economy is one strategy that can be used for managing scarce or endangered resources. Instead of disposing of critical elements after their use, the circular economy focuses on recovering these resources.

26


27 Vertical trends are explained by differences in the shielding effect. Down a group, increasing shield effect causes larger atomic radii, decreasing ionisation energy and decreasing electronegativity.
Horizontal trends are explained by differences in core charge. Across a period, increasing core charge causes smaller atomic radii, increasing ionisation energy and increasing electronegativity.

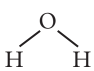
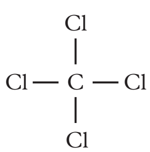
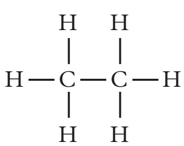
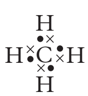
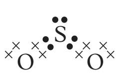
Chapter 3: Covalent substances

GROUNDWORK

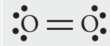
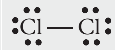
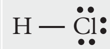
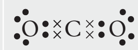
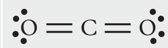
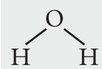
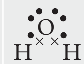
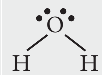
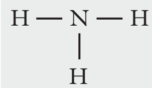
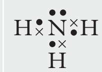
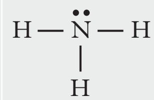
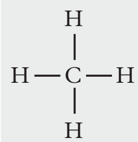
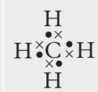
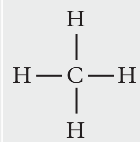
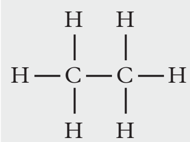
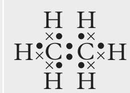
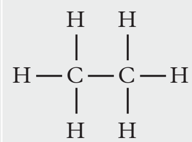
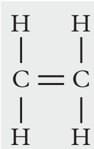
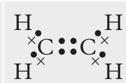
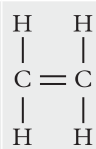
3A Non-metals located on the left-hand side of the periodic table.
3B By sharing electrons
3C The ability of an atom that is bonded to another atom to attract bonding electrons to itself
3D A characteristic of matter that can be measured or observed without changing the chemical composition of the substance (e.g. hardness, boiling point)

3.1 Covalent compounds

3.1 CHECK YOUR LEARNING

1 A bond formed when two or more non-metal atoms share electrons
2 Non-metal atoms form bonds so that each atom has eight electrons in its outer shell by gaining, losing or sharing electrons and becoming stable (except hydrogen and helium).
3 Diatomic molecule: made up of two atoms of the same or different elements; polyatomic molecule: made up of more than two atoms of the same or different elements
4 **a** 1
b 2
c 1
d 3
5 **a** 
b 
c 
d H—F
6 **a** 
b 
c 
d 
7 **a** XM_3
b X_2
c XYM_3
d Y_2

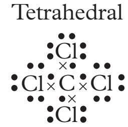
8

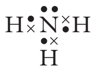
Molecular formula	Structural formula	Lewis structure	Valence structure
H_2	H—H	H:H	H—H
O_2	O=O		
Cl_2	Cl—Cl		
N_2	N≡N		
HCl	H—Cl		
CO_2	O=C=O		
H_2O			
NH_3			
CH_4			
C_2H_6			
C_2H_4			

9 Lewis structures show bonded electrons as dots and valence structures show them as lines. H_2, CH_4 and C_2H_6 only have single covalent bonds between atoms; C_2H_4 has one double covalent bond between carbon atoms.
10 Student answers will vary but should include three advantages and three disadvantages such as those outlined in Table 2 on page 9.

3.2 Shapes of molecules

3.2 CHECK YOUR LEARNING

1 **a** Atoms arranged in straight line, central atom bonded to one or two atoms
b Atoms arranged in a V-shape; central atom bonded to two atoms
c Atoms arranged in a triangular shape; central atom bonded to three atoms with one lone pair present
d Atoms arranged in a triangular pyramid shape; central atom bonded to four atoms
2 Valence electron pairs (bonded and unbonded) in an atom of a molecule repel one another because of their negative charge and arrange themselves to be as far apart as possible.
3 Unbonded valence electron pairs (along with bonded valence electron pairs) are arranged as far apart as possible, thus determining the shape of the molecule.
4 **a** Linear
b Pyramidal
c Linear
d Linear
e Tetrahedral
5 **a** 
b 

c 

d 

e 

6 a Tetrahedral
b Linear
c Pyramidal
d V-shape or bent
e Linear

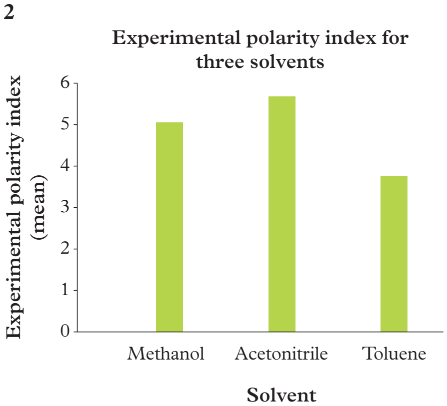
7 Molecules with double or triple bonds tend to be linear, except molecules that contain unbonded valence electron pairs (e.g. SO_2, which is bent or V-shaped).

3.3 Polar and non-polar characteristics

3.3 SKILL DRILL

1 Methanol 5.1, acetonitrile 5.7, toluene 3.8

2



3 Low accuracy: the results for toluene are higher than expected when compared with the theoretical polarity index. Low precision: the result of 6.5 for acetonitrile is significantly different from the other two results.

4 Acetonitrile has a potential outlier at 6.5.

3.3 CHECK YOUR LEARNING

1 The shape of a molecule determines the direction of each bond dipole in the molecule and thus the polarity of the molecule.

2 O–H; greatest difference between electronegativity of its atoms

3 a N–F: fluorine, H–C: carbon, O–S: oxygen, P–O: oxygen, Cl–P: chlorine, N–C: nitrogen
b N – partial positive, F – partial negative; H – partial positive, C – partial negative; O – partial negative, S – partial positive; P – partial positive, O – partial negative; Cl – partial negative, P – partial positive; N – partial negative, C – partial positive
c P–O
d H–C and N–C

4 a Polar
b Polar
c Non-polar
d Polar
e Polar
f Polar

5 a Polar
b Polar
c Polar
d Non-polar
e Polar

6 Polar bond: covalent bond with unequal sharing of electrons between two atoms; polar molecule: has one or more polar covalent bonds with the charge distributed unevenly. A non-polar molecule can have polar bonds if the bond dipoles cancel out.

7 Non-polar; contains polar covalent bonds but the charge is distributed evenly

3.4 The relative strength of bonds

3.4 CHECK YOUR LEARNING

1 a Present in all molecules and atoms
b Present in polar molecules
c Present in polar molecules that contain an H bonded to an F, O or N atom

2 CH_3Cl

3 a Dipole–dipole attraction
b Dipole–dipole attraction
c Dispersion forces
d Dispersion forces
e Hydrogen bonding

4 $CH_3OH > CHCl_3 > HCN > CO_2 > H_2$

5 Hydrogen bonding (strong attraction between highly electronegative F, O or N atoms and partially positive H atom) > dipole–dipole attractions (attraction between permanent dipoles) > dispersion forces (weak attraction between temporary dipoles)

6 The lone pairs of the O atom in propanone can form hydrogen bonds with the H atoms of the water molecules, but propanone lacks an H bonded to F, O or N atoms, so cannot hydrogen bond with itself.

7 Example question to answer *before*: 'Is there a permanent bond dipole?'; example question to answer *after*: 'How many hydrogen bonds can the molecule form?'

3.5 Physical properties of molecular substances

3.5 CHALLENGE

a i CH_2Cl_2
ii SO_2
iii SF_2
iv OCl_2
v H_2O_2
vi CH_3OH

b i CH_2Cl_2
ii SO_2
iii SF_2
iv NCl_3
v H_2O_2
vi $CH_3CH_2CH_2OH$

3.5 CHECK YOUR LEARNING

1 They do not contain charged particles.

2 Cl_2 −34.6°C, H_2O 100°C, CH_4 −161.6°C, HCl −85.0°C

3 F_2: non-polar, has weak dispersion forces between molecules; HF polar, has strong hydrogen bonding between molecules

4 Melting point: temperature at which a solid becomes a liquid; boiling point: the temperature at which a liquid becomes a vapour

5 a $HF > HCN > CO_2 > O_2$
b ii > iii > i
c ii > i > iii > iv

6 HF forms strong hydrogen bonds with itself = highest boiling point. HCl, HBr and HI form dipole–dipole attractions = lower boiling point. HI > HBr > HCl due to size of the halogen.

3.6 The structure and bonding of diamond and graphite

3.6 REAL-WORLD CHEMISTRY

1 Similarities: contain carbon atoms covalently bonded in a network lattice. Differences: natural diamonds have impurities whereas laboratory grown diamonds do not.

2 Laboratory-grown diamonds are more pure than natural diamonds.

3 Laboratory-grown diamonds are harder and are more useful in industries such as electronics, lasers and computers.

4 Laboratory-grown diamonds are high-quality, pure carbon diamonds, whereas natural diamonds are often flawed or weak because of structural defects.

3.6 CHECK YOUR LEARNING

1 Carbon can form four covalent bonds with a variety of other atoms, including other carbon atoms.

2 Both contain only strong covalent bonds between carbon atoms in a network lattice, which require more energy to break than intermolecular bonds.

3 **a** Diamond: very hard because it only contains strong covalent bonds between carbon atoms in the covalent network lattice. Graphite: soft because it contains weak dispersion forces between layers of carbon atoms.

b Diamond: does not conduct electricity because it does not contain delocalised electrons. Graphite: contains delocalised electrons that can carry a charge.

c Both have high thermal conductivity. Diamond: heat passes through the network lattice. Graphite: heat passes through the layers.

4 **a** Diamond can conduct heat away from sensitive electronic parts and can keep heat-sensitive parts cool.

b Graphite is soft and slippery because the layers slide over one another reducing the friction between moving machinery parts.

Chapter 3 review

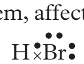
MULTIPLE CHOICE

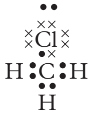
1	B	**2**	C	**3**	A	**4**	B	**5**	D
6	A	**7**	C	**8**	C	**9**	D	**10**	C

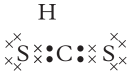
SHORT ANSWER

11 Hydrogen bonding occurs between polar molecules that have a H bonded to a F, O or N. Dipole–dipole attraction occurs between polar molecules with permanent dipoles. Dispersion forces occur between non-polar molecules due to temporary dipoles

12 Electron pairs (including non-bonding electrons) will repel each other to minimise negative charges between them, affecting molecular shape.

13 **a** 

b 

c 

d 

14 A double covalent bond is stronger than a single covalent bond, so requires more energy to break.

15 Molecules that contain atoms with different electronegativities have an unequal sharing of electrons in the bonding electrons. This creates polar covalent bonds in the molecule.

16 Graphite contains delocalised electrons, which can move freely throughout a layer of carbon atoms and carry electrical charge. Naphthalene does not contain any free electrons so cannot conduct electricity.

17 Diamond does not contain delocalised electrons; the electrons are bonded tightly to carbon atoms and so are not able to move freely throughout the network lattice.

18 Long straight-chain molecules have stronger dispersion forces than smaller, more compact molecules with the same number of electrons. The closer the molecules can get to one another, the more dispersion forces between them.

19 **a** Linear
b Linear
c Tetrahedral
d Linear
e Pyramidal
f Linear
g Linear
h Pyramidal

20 **a** Non-polar: a, b, f; Polar: c, d, e, g, h
b Dispersion forces: a, b, f; Dipole–dipole attractions: c, e, g; Hydrogen bonding: d, h

21 **a** **i** N–H – nitrogen
ii C–H – carbon
iii O–H – oxygen
iv H–H – identical
v Cl–O – oxygen

b N – partial negative, H – partial positive; C – partial negative, H – partial positive; O – partial negative, H – partial positive; no partial charges; O – partial negative, Cl – partial positive

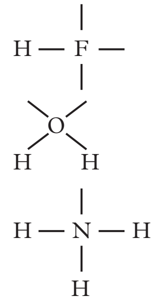
c **iii** O–H
d **v** Cl–O

22 **a** O_2, C_2H_4
b N_2, HCN
c H_2S
d N_2, O_2, C_2H_4
e HCN, H_2S, HCl
f HCN

23 Incorrect; only true if the diatomic molecules are non-polar

24 **a** $CH_3CH_2OH > CH_3C{=}OCH_3 > CH_3CH_2CH_3$
b $I_2 > Br_2 > Cl_2$
c $Cl_2 > CO_2 > N_2 > H_2$

25 Chloroethane and chlorohexane are non-polar molecules. The strength of the dispersion forces increases as the size of the molecule increases. Chlorohexane has a higher molar mass; therefore, it has stronger dispersion forces and a higher boiling point.

26 **a** 

b HF

27 **a** Both are polar, so have dipole–dipole attractions. However, $CHBr_3$ is a bigger molecule; therefore, it has stronger dispersion forces than $CHCl_3$.

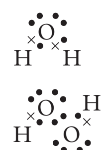
b Both are polar. Dipole–dipole attractions in HCl < hydrogen bonding in HF.

c Br_2 is non-polar; therefore, it has weak dispersion forces. ICl is polar; therefore, it has dipole–dipole attractions, which are stronger than dispersion forces.

28 H_2 – linear, dispersion forces, low melting point and boiling point

HCl – linear, dipole–dipole attraction, low melting point and boiling point

CCl_4 – tetrahedral, dispersion forces, low melting point and boiling point

29 **a** 

b Covalent bonding

c Both H_2O and H_2O_2 have hydrogen bonding.

30 **a** **i** PF_3, **ii** NH_3, **iii** CH_3Cl, **iv** Both are non-polar, **v** PCl_3, **vi** HF, **vii** CH_3OH

b **i** PF_3, **ii** NH_3, **iii** CH_2Cl_2, **iv** O_2, **v** PCl_3, **vi** HF, **vii** CH_3OH

31 Temporary dipoles: exist in all molecules and result in dispersion forces; permanent dipoles: exist in molecules due to the different electronegativities of atoms, resulting in dipole-dipole interactions

32 Strong hydrogen bonding between water molecules creates a surface that is stretched tight like a skin and is called surface tension. Surface tension allows objects to float on its surface.

33 Both contain only strong covalent bonds between carbon atoms in a network lattice, which require more energy to break than intermolecular bonds.

34 **a** I_2 is a larger molecule, with more electrons, than HI. Therefore, I_2 has stronger intermolecular forces than HI, which require more energy to break.

- **b** HF has strong hydrogen bonding. which requires more energy to break than the dispersion forces present in F_2.
- **c** HF contains a H atom bonded to an F atom which is the most electronegative halogen; therefore, it can form stronger dipole–dipole attractions than the other hydrogen halides.
- **d** The boiling point of halogens increases as you move down the group because the size of the molecules increase; therefore, they have stronger dispersion forces.

35 In water, the intramolecular bonding (covalent bonding) between the hydrogen and oxygen atoms is very strong compared with the hydrogen bonding between H_2O molecules; therefore, it requires a lot of energy to break the covalent bonding.

Chapter 4: Reactions of metals

GROUNDWORK

4A Lustrous, malleable, ductile, generally high melting point, good conductors of heat and electricity, dense and hard

4B It becomes positive.

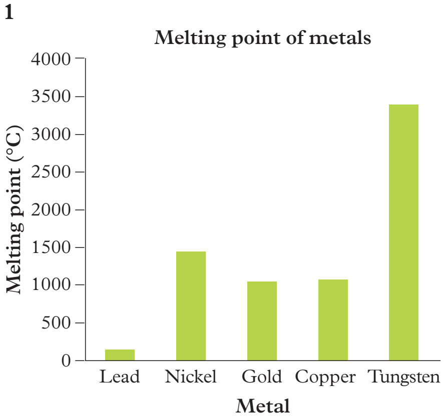
4C Sodium is a metal in group 1 of the periodic table. Elements in group 1 have the lowest first ionisation energy (the amount of energy required to remove one valence electron).

4D They have a range of properties that make them useful such as electrical and heat conductivity and high melting points.

4.1 Properties of metals

4.1 SKILL DRILL

1



2 Lead is in the p-block of the periodic table (group 14) and elements in this group usually have lower melting points than the transition metals (nickel, gold, copper and tungsten).

3 Tungsten is a transition metal and transition metals generally have higher melting points than other metals.

4.1 CHECK YOUR LEARNING

1 The electrostatic attraction that occurs between positive metal (cations) and negative delocalised electrons.

2 $1s^2 2s^2 2p^6 3s^2 3s^2 3p^6 4s^2$; +2

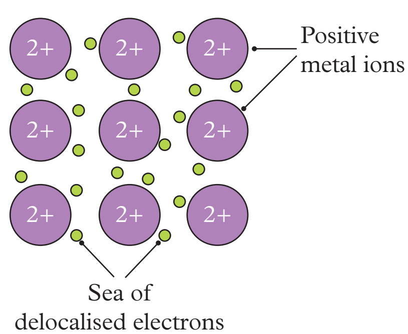
3
- **a** Metals are good conductors of heat due to the delocalised electrons within the metallic lattice. These electrons can vibrate and transfer kinetic energy (heat) throughout the lattice.
- **b** Metals are good conductors of heat because they transfer the heat away from your skin quickly. Wood is a heat insulator because it transfers the heat away from your skin more slowly.
- **c** Metals are malleable because of the delocalised electrons, which allow metal cations to slide over each other when a force is applied without breaking the metallic bonding.
- **d** Metals and our bodies are good electrical conductors; therefore, if you touch someone who is still in contact with live wires, an electrical current can flow from them to you.

4
- **a** Gold can be bent/hammered into shape (malleable) and be heated to its melting temperature so that it becomes molten.
- **b** It tells us that the particles do not experience repulsion when a force is applied because the gold sheet can be hammered into a shape without breaking. When cations are forced to slide over each other,metallic bonding remains intact because of the sea of delocalised electrons.

5 Lithium has a lower number of protons and delocalised electrons and a smaller ionic radius; therefore, it has weaker metallic bonding. Calcium has more protons and delocalised electrons, and a larger ionic radius; therefore, it has stronger metallic bonding.

6 Student answers will vary but should discuss the properties of the metals chosen and how they relate to their use.

7



4.2 Reactivity series of metals

4.2 CHECK YOUR LEARNING

1 Scientists have collected and used experimental data on how different metals react with water, acids and oxygen to determine metal reactivity series.

2 The presence of hydrogen gas can be confirmed by holding a lit splint (or match) over the reaction vessel. If an audible 'pop' sound is produced, then hydrogen gas is present.

3 The reactivity of metals increases going down a group and decreases moving from left to right across the table. Caesium is the most reactive because it is in group 1. Scandium is less reactive than caesium because the reactivity of transition metals is generally lower than that of group 1. Tin is the least reactive because reactivity of metals decreases moving left to right across a period.

4 Student answers will vary but should include how reactive the metal is to oxygen and water by referring to the reactivity series of metals.

5 This is to ensure that they do not get exposed to oxygen and water because they are highly reactive.

4.3 Metal recycling

4.3 CHECK YOUR LEARNING

1
- **a** Mining → Refining using Bayer process → Smelting using Hall–Heroult process
- **b** Collecting scrap → Sorting → Crushing → Remelting → Secondary casting → Repurposing

2 In the smelting process, aluminium oxide and carbon react to form aluminium metal and carbon dioxide. This can be expressed through the following reaction:

$2Al_2O_3 + 3C \rightarrow 4Al + 3CO_2$

3 Old scrap

4 Aluminium has an extensive range of uses and maintain their core chemical properties through repeated recycling and reprocessing.

5 New scrap is scrap taken from the manufacturing and/or fabrication of aluminium products; it may be safely recycled if the composition is known. Old scrap has been used by the consumer and is discarded or disposed; it cannot be safely recycled because its composition is unknown, and it may be contaminated.

6
- **a** Rolled product
- **b** Milled product
- **c** Castings

7 This statement is correct. In the aluminium recycling process, paper, plastic and other non-aluminium recycling is removed, so that it can be repurposed.

Chapter 4 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 a Hardness and density
b Conductor of heat
c Lustrous
d Hardness

12 a Cs > K > Mg > Fe > Pb > Au
b Cs and K
c Au

13 The amount of energy required to remove one valence electron in the gaseous state. If the first ionisation energy is low, the metal is more reactive because it can lose valence electrons readily during chemical reactions.

14 a i Metal solids consist of tightly packed solid crystals. If one of these crystals is viewed through an electron microscope, a three-dimensional arrangement of atoms is seen.
ii Metals are malleable, ductile, lustrous, and good conductors of electricity and heat because they have delocalised electrons within the metallic lattice.
b Metals have a low ionisation energy; therefore, metal atoms tend to easily lose their valence (outer shell) electrons and form cations (positive ions).

15 a A large amount of energy is needed to break the strong metallic bonds between the metal cations (positive ions) and delocalised electrons.
b Metals are good conductors of electricity because electrons are free to move around the metallic lattice completing an electrical circuit.
c Delocalised electrons allow metal cations to slide over each other when a force is applied without breaking the strong metallic bonds; the electrostatic attraction between the cations (positive ions) and delocalised electrons remains intact.

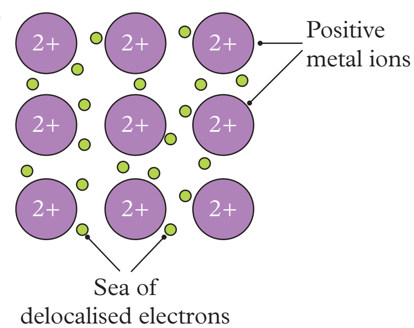
16 Goal 12 'Responsible consumption and production'. Recycling metals ensure that resources are consumed and produced sustainably so that the needs of people in the present and future are met.

17 Student answers will vary but should include that a linear economy uses a 'take–make–dispose' model where resources are extracted from the Earth to make products that will be thrown away.

18 a Lithium and sodium are both group 1 metals and have lower melting temperatures than iron, chromium, copper and zinc which are transition metals.
b One of: denser, stronger and harder.
c One of: conductors of heat and electricity, malleable and ductile.
d Student answers will vary but should include one of the following:
- good conductors of heat due to the delocalised electrons which can vibrate and transfer kinetic energy (heat) throughout the metallic lattice
- good conductors of electricity due to the delocalised electrons which can send an electrical charge throughout the metallic lattice
- malleable and ductile due to the delocalised electrons, which allow metal cations to slide over each other when a force is applied without breaking metallic bonds throughout the metallic lattice.

19 A circular economy is a sustainable approach to managing resources that involves a continuous cycle of using and reusing resources from natural raw materials. Aluminium recycling conserves natural resources of bauxite and saves up to 95% of the energy required to extract aluminium from its bauxite ore.

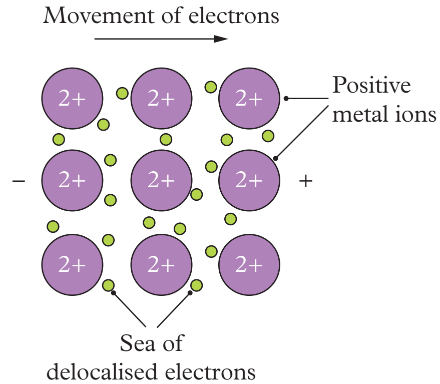
20



21 Student answers will vary but should include a variety of metals reacting with oxygen, water and acid.

22 Student answers will vary.

23 a Student answers will vary. Diagram provided shows positive metal ions with a 2+ charge (e.g. calcium or magnesium).



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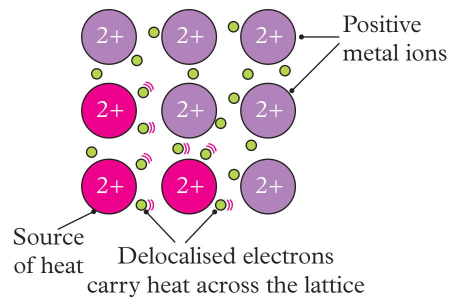
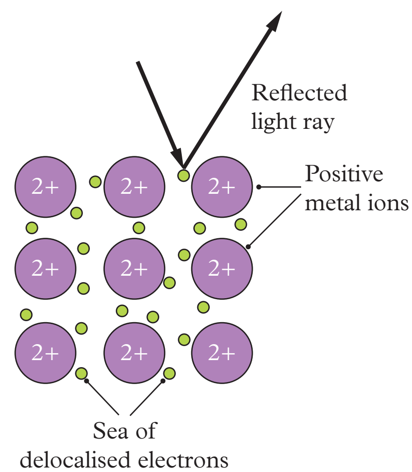


c Student answers will vary. Diagram provided shows positive metal ions with a 2+ charge (e.g. calcium or magnesium).



Chapter 5: Reactions of ionic compounds

GROUNDWORK

5A Cations are positively charged ions that are formed by atoms losing electrons.
Anions are negatively charged ions that are formed by atoms gaining electrons.

5B Ionic bonding is one of the three types of atomic bonding and involves electrostatic attraction between positive cations and negative anions.

5C Ionic compounds are generally soluble in water because most can be pulled apart by the highly polarised partially charged ends of water molecules. Water molecules can move between ions and separate ionic bonds to form free mobile ions.

5D Ionic bonds are stronger than both covalent and metallic bonds.

5.1 Properties of ionic compounds

5.1 SKILL DRILL

1 Student 1's suggestion was based on opinion from their subjective past experiences.

Student 2's suggestion was based on evidence from external creditable sources.

Student 3's suggestion was based on anecdote from objective data from past experiments.

2 Student 1 stated that ionic compounds are lustrous; this is a non-scientific idea because there are no free moving particles in solid ionic compounds to reflect light.

Student 2 stated that ionic compounds have high melting points; this is a valid scientific idea because solid ionic compounds are held together by strong ionic bonds that require a lot of energy/heat to break during melting.

Student 3 stated that ionic compounds can conduct electricity; this is partially true because ionic compounds can conduct electricity in either the aqueous or the molten state because of the existence of free-moving ions.

5.1 CHECK YOUR LEARNING

1 Ionic bonds are formed from the electrostatic attraction between positive cations and negative anions.

2 Main group elements have the same number of valence electrons. Groups 1 and 2 elements have a tendency to lose their valence electron to have a full outer shell, whereas non-metal elements have a tendency to gain electrons to have a full outer shell.

3 Solid ionic compounds form crystal lattices that are held in place by ionic bonds. Ions in the lattice are arranged in a regular repeating pattern so the ions are only surrounded by their opposite charges.

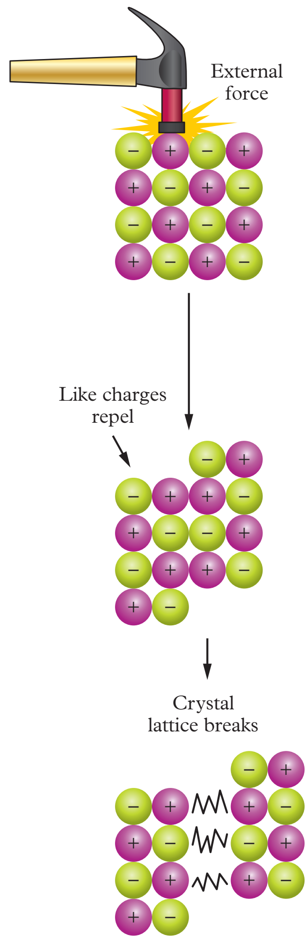
4 CsCl – cation caesium, anion chlorine
ZnS – cation zinc, anion sulfur
TiO_2 – cation titanium, anion oxygen
CaF_2 – cation fluorine, anion calcium

5 Solid ionic compounds are arranged in rigid crystal lattice in which the charged ions are unable to move and hence unable to conduct electricity. In liquid molten ionic charged ions can move around and slide past each other, allowing molten ionic compounds to conduct electricity.

6 Both solid metal and solid ionic compounds are arranged in crystal lattice formation. However, solid metals have a lustrous appearance whereas solid ionic compounds do not. This is because metallic crystals contain delocalised electrons which can reflect light, whereas ionic crystals do not have free moving electrons.

Metallic compounds are malleable because of the sea of negative electrons holding the positive cation lattice of the metallic compound when forces is applied. Ionic compounds are brittle and non-malleable because solid ionic compounds are arranged in a crystal lattice in which the charged ions are unable to move.

7



5.2 Formation of ionic compounds

5.2 CHECK YOUR LEARNING

1 Step 1. Write the electron configurations.

Step 2. Determine how many electrons need to be lost or gained to make each atom stable.

Step 3. Draw both elements next to each other as electron shell diagrams.

Draw an arrow from cation to anion to indicate the transfer of electron(s) and label it with e^-.

- More than one electron can be transferred.
- The number of electron donators and electron acceptors can be varied.

Step 4. Draw the final ions to the right including the charge.

2 All electrons and charges need to be accounted for. The positive charges of the cations and the negative charges of the anions need to be balanced and cancel out because ionic compounds must be electrically neutral.

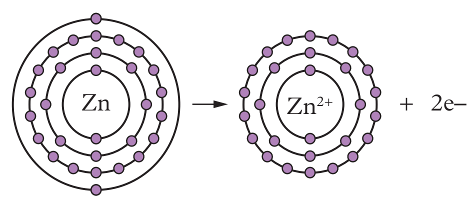
3 **a** Lose 2 electrons (Pb^{2+}) or lose 4 electrons (Pb^{4+})

b Gain 2 electrons (Se^{2-})

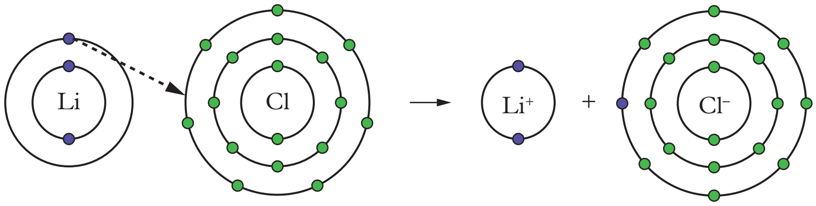
c Lose 2 electrons (Ba^{2+})

d Lose 3 electrons (Sb^{3+})

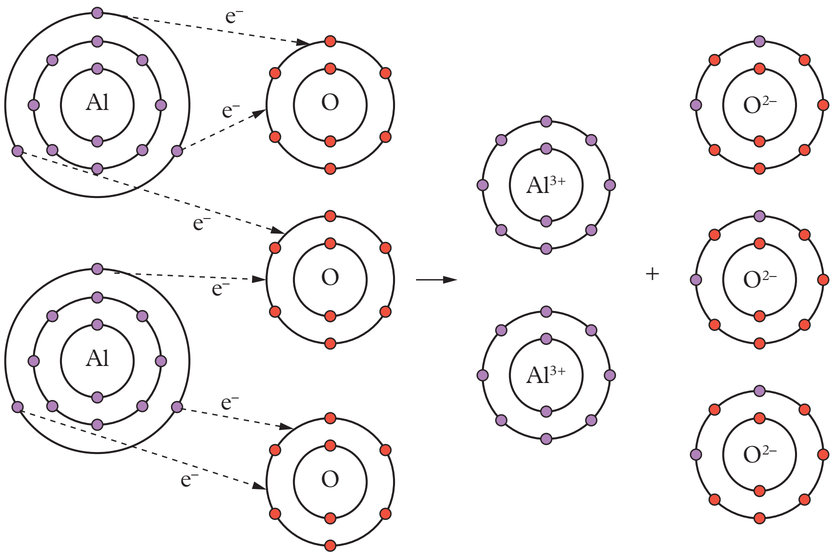
4 Student examples will vary but should recognise that when a transition metal atom donates electrons to oxygen to form an ionic compound, the resulting metal ion will have a smaller radius than the original atom. This is due to fewer shells being occupied because of the loss of valence electrons.



5 a



b



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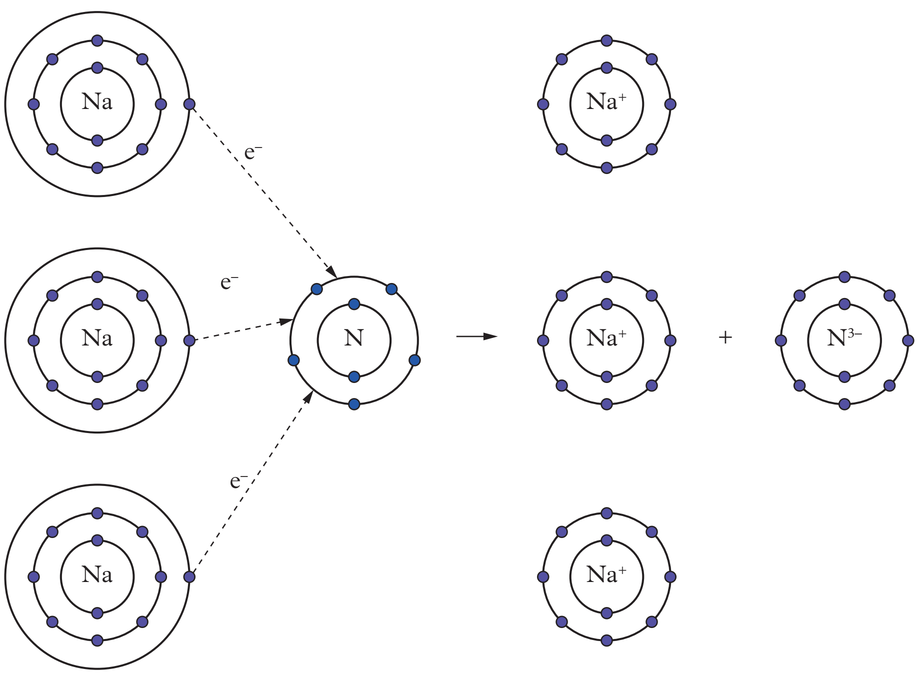


5.3 Formulas and naming of ionic compounds

5.3 CHALLENGE

1 Chromium(II) sulfate
2 Calcium phosphate
3 Iron(II) sulfate octahydrate

5.3 CHECK YOUR LEARNING

1 Simplest whole-number ratios of atoms are used to write ionic compound formulas because the actual number of ions in a sample varies depending on the size of the sample without any changes in chemical properties. The empirical formula can also enable scientists to determine the charges of transition cations which can have different charges.
2 The valency of the transition metal is included after the name of the metal followed by the charge number written in Roman numerals
3 **a** CrO_2
  b $NaHSO_4$
  c AlI_3
4 **a** Calcium carbide
  b Ammonium sulfate
  c Iron(III) sulfide
5 Monoatomic ions are formed from one atom, whereas polyatomic ions are formed from two or more atoms per ion.
6

Steps	Example
1 Place the cation first with its correct charge.	Al^{3+}
2 Place the anion second with its correct charge.	Al^{3+} NO_3^-
3 Combine the positive and negative ions so the overall positive and negative charges balance.	Al^{3+} + $(NO_3^-) \times 3$
4 Write the number of each ion needed to achieve a neutral charge with a subscript after the ion.	$Al(NO_3)_3$

5.4 Precipitation reactions

5.4 CHECK YOUR LEARNING

1 A precipitate is the insoluble, solid ionic compound formed by the reaction of two soluble ionic solutions.

2 Spectator ions are ions that stay in aqueous state throughout the precipitation reaction and do not change.

3 Dissociation is the breaking up of the ionic lattice structure when dissolved into an aqueous solution.

Student examples will vary but must include a solid reactant and two aqueous ionic products. For example: $KNO_3(s) \rightarrow KNO_3(aq) \rightarrow K^+(aq) + NO_3^-(aq)$

4 Balanced equations include all reactants and products regardless of state changes. Ionic equations only include the ions that change states during a chemical reaction, i.e. they do not include spectator ions. Ionic equations are also shorter than balanced equations.

5 a $MgI_2(aq) + Na_2CO_3(aq) \rightarrow MgCO_3(s) + 2NaI(aq)$

b $Hg(NO_3)_2 + Na_2S(aq) \rightarrow HgS(s) + 2NaNO_3(aq)$

c $2NaOH(aq) + CuSO_4(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$

6 Student answers will vary. For example, the mnemonic SNAPE to remember ions that are always soluble.

- Sodium
- Nitrate
- Ammonium
- Potassium
- Ethanoate

Chapter 5 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Potassium only has 1 valence electrons in its outer shell and only needs to lose 1 electron to achieve a full outer shell. If it loses 2 electrons, its outer shell will not be full. Similarly, chlorine only has one place left in its valence shell to achieve a full outer shell. If it gains 2 electrons, it will have an outer shell that is not full.

12 a KF

b No. Magnesium fluoride has an empirical formula of MgF_2. This means it has an atom ratio of 1:2, which requires a different arrangement from KF, which has an atom ratio of 1:1.

c Solid potassium fluoride ions exist in a rigid crystal lattice and cannot move, whereas while molten potassium fluoride ions can move and slide past each other. This ability for charged particles to move enables the conduction of electricity.

13 Double displacement reactions are chemical reactions in which both sets of ions swap or exchange partners in solution to form new products. All precipitation reactions are double displacement reactions because they involve ion swapping the aqueous reactants. However, only double displacement reactions forming principates are considered precipitation reactions.

14 a CrP

b Fe_2O_3

c $Pb(NO_3)_2$

d $(NH_4)_4C$

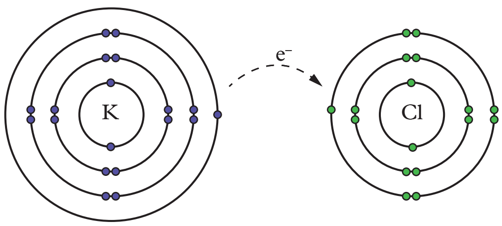
15 $AuClO_4$

16 A yellow precipitate of lead iodide will be produced.

$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

17 a $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

b



18 a $NiSO_4$

b NH_4Cl

c $Al_2(HPO_4)_3$

d $Fe(NO_3)_3$

e Ag_2S

f CaC_2

19 Substance B – it has a high melting temperature, does not conduct electricity as a solid (at temperatures lower than melting temperature) but does conduct electricity when melted (at temperatures higher than melting temperature)

20 a Ammonium bromide

b Lead(IV) sulfate

c Iron(III) nitrate

d Copper(I) phosphide

e Aluminium sulfite

f Potassium carbonate

21 a $2Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2CrO_4(s)$

$2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$

b $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

$Al_2(SO_4)_3(aq) + 3Ca(OH)_2(aq) \rightarrow 2Al(OH)_3(s) + 3CaSO_4(aq)$

c $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$

d $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

$Ba(NO_3)_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + 2LiNO_3(aq)$

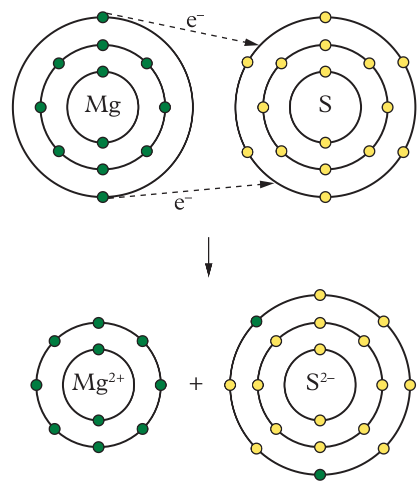
22 a Potassium iodide ions have a 1+ and 1– charge respectively, whereas magnesium oxide ions have a 2+ and 2– charge respectively. Magnesium oxide will have a higher melting point because of the stronger ionic bonds between the higher charged ions.

b All ions involved have a 1+ or 1– charge. Lithium fluoride will have the higher melting point because of the smaller ions, which allow for smaller distances between the ions in the ionic lattice solid state.

c All ions involved has a 2+ or 2– charge. Magnesium oxide will have the higher melting point because of the smaller ions, which allow for smaller distances between the ions in the ionic lattice solid state.

d Both ionic compounds contain sodium cations and a 1– anion. Sodium chloride will have the higher melting point because of the smaller chloride ions, which allow for smaller distances between the ions in the ionic lattice solid state.

23



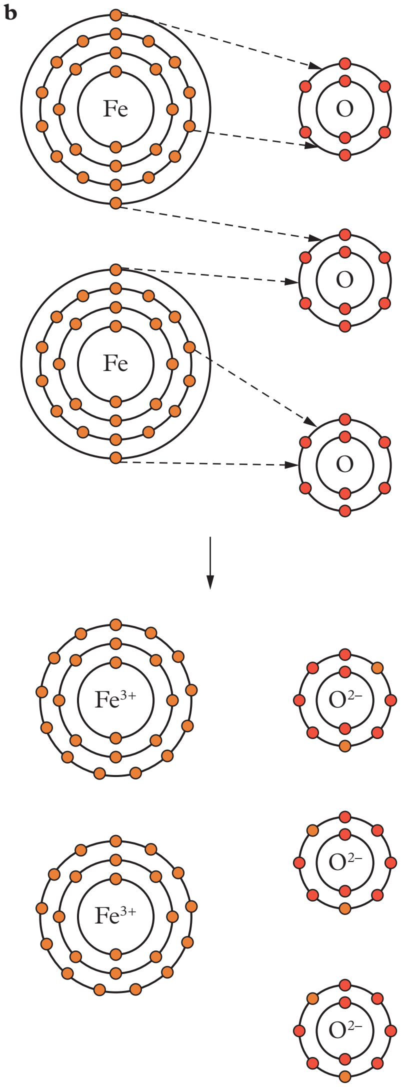
24 a $2NaOH(aq) + CaCl_2(aq) \rightarrow Ca(OH)_2(s) + 2NaCl(aq)$

b $CuBr_2(aq) + (NH_4)_2CO_3(aq) \rightarrow CuCO_3(s) + 2NH_4Br(aq)$

c $3K_2SO_4(aq) + 2Fe(NO_3)_3(aq) \rightarrow Fe_2(SO_4)_3(aq) + 6KNO_3(aq)$

25 a Fe_2O_3

b



26 Student answers will vary.

Chapter 6: Separation and identification of the components of mixtures

GROUNDWORK

6A The polarity of a molecule depends on molecular shape and the polarity of each individual bond.

6B Dispersion forces, dipole–dipole attraction, hydrogen bonding

6C The formation of intermolecular forces depends on molecular weight (dispersion forces), polarity (dipole–dipole attraction) and the configuration of intramolecular bonds and types of atoms (hydrogen bonding).

6.1 Polar and non-polar character in chromatography

6.1 SKILL DRILL

1 Some larger polar molecules may have low solubility because of the large non-polar sections of the molecules.

2 Data collected was worded descriptions of observations and not numerical/quantifiable values.

3 Student answers will vary, but may include melting point, boiling point and solubility under different conditions. A sample response can include measuring the temperature at which each molecular compound boils.

6.1 CHALLENGE

A polar mobile phase will be able to ionise and dissolve the sodium chloride and form dipole–dipole attraction hydrogen bonds with butanol. A non-polar stationary phase will retain the non-polar hexane.

6.1 CHECK YOUR LEARNING

1 a A solvent that moves over or through the stationary phase during chromatography. It carries the sample over the stationary phase to be separated. Without the mobile phase, the sample will not move.

b A substance that does not move during chromatography. It competes with the mobile phase to separate out components of the sample. Without the stationary phase, the sample will be dissolved in the mobile phase.

c The attraction of components of the sample to the mobile or stationary phase. The differences between different components and its affinity for mobile versus stationary phase allows for the separation of mixtures in chromatography.

d Components of the sample desorb (unstick) from the stationary phase and return to the mobile phase if they have an affinity for the mobile phase. The different rates of absorption and desorption of the components allows for their separation from a mixture.

e Intermolecular forces are attractive forces that can hold molecules together, which influence their affinity for the mobile and stationary phases. The 'like dissolves like' rule is used to predict whether two molecular compounds will be soluble based on polarity. Non-polar solutes tend to only dissolve in and have high affinity for non-polar solvents and polar solutes tend to only dissolve in and have high affinity for polar solvents.

2 A polar component of a mixture is more attracted to a polar mobile phase (e.g. water) than a non-polar mobile phase (e.g. hexane) because it can form dipole–dipole attractions with water molecules.

3 Ethanol: stronger affinity for the polar mobile phase because of its hydroxyl group and a small hydrocarbon chain. Ethyl ethanoate: a larger molecule that does not contain a hydroxyl group, will have lower affinity for the polar mobile phase.

4 Both polar and non-polar substances contain dispersion forces. However, polar substances can additionally form dipole–dipole attractions. A non-polar stationary phase can be used with a polar mobile phase to separate solutes with different polarities from a sample mixture.

5 Water is an extremely polar molecule that can form all three types of intermolecular forces. Water is also inexpensive and readily available and can dissolve a wide range of ionic compounds.

6 a Component B – can form strong dipole–dipole attractions to the polar mobile phase because of its polarity.

b Component C – the least polar, meaning that it is also the most non-polar; more likely to interact with the non-polar stationary phase and form dispersion forces than the other components.

c Component B – can form strong dipole–dipole attractions and adsorb most readily into to the polar stationary phase.

d Component C – the least polar, meaning that it is also the most non-polar; will readily desorb from a polar stationary phase because it has higher affinity for the non-polar stationary phase.

7 Dipole–dipole attractions; therefore, should be polar.

8 The 'like dissolves like' rule predicts whether two molecular compounds will be soluble based on its polarity. Polar molecules will interact with other polar molecules and form dipole–dipole attractions rather than interact with non-polar molecules. Non-polar molecules only interact with other non-polar molecules via dispersion forces.

6.2 Chromatography

6.2 REAL-WORLD CHEMISTRY

1 By using the beef DNA as a control, scientists were able to see whether there were any components of the burgers' DNA which fell outside the beef DNA range.

2 By releasing more details about the types of conditions under which the tests were performed, e.g. how the samples were collected, the type of mobile and stationary phases that were used.

3 By using different techniques and ensuring the calibration of equipment.

6.2B CHALLENGE

1 Valid because the retention times for all three components were recorded

2 Change the mobile phase to a mixture that is less polar to better separate component B

3 The blue dye will have a lower retention time because it is more polar and have a lower affinity to the non-polar stationary phase than the yellow dye.

6.2 CHECK YOUR LEARNING

1 Step 1. Measure the distance from the origin to the solvent front.

Step 2. Measure the distance from the origin to the middle of each component.

Step 3. Divide the distance to each component by the distance to the solvent front.

2 The mobile phase is the solvent that moves through the stationary phase during paper or thin-layer chromatography. The components of a sample separate based on their relative affinity for the mobile versus stationary phases.

3 **a** R_f(green) = 0.081; R_f(purple) = 0.24; R_f(blue) = 0.66

b Blue: has the highest R_f, travelled the furthest

c Green: has the lowest R_f, travelled the least

4 Green: probably polar because of its high affinity for the polar stationary phase; would experience dipole-dipole attractions with the stationary phase and only a small amount of dispersion forces with the mobile phase.

Purple: a slightly higher R_f and is likely to have mostly polar molecules with some non-polar components; will probably form dispersion forces with the mobile and stationary phase with a small amount of dipole–dipole attraction with the polar stationary phase.

Blue: likely to be more non-polar as it has a higher R_f; will form mainly dispersion forces with the mobile phase and some dispersion forces with the stationary phase.

5 **a** R_f(A1) = 0.15 – Indigotine;
R_f(A2) = 0.40 – Erythrosine;
R_f(B1) = 0.38 – Fast green FCF;
R_f(B2) = 0.88 – Patent blue V;
R_f(C1) = 0.13 – Brilliant blue FCF;
R_f(C2) = 0.65 – Tartrazine

b The components may look the same to the naked eye but are different due to their different R_f values, caused by their different intermolecular interactions with the mobile and stationary phases.

6 **a**

Team	R_f		
	1	2	3
1	0.56	0.79	0.21
2	0.32	0.59	0.87
3	0.56	0.79	0.21
4	0.56	0.79	0.21
5	0.32	0.59	0.87
Crime scene note	0.32	0.79	0.87

b Teams 1, 3 and 4; Teams 2 and 5.

c Teams 2 and 5

d The components for the pens are unique because of the different R_f values.

e The chromatography were completed under known conditions so the retardation factors of each component can be compared to a list of retardation factors of known substances performed under the same conditions.

Chapter 6 review

MULTIPLE CHOICE

1	D	**2**	B	**3**	D	**4**	A	**5**	A	**6**	C
7	C	**8**	B	**9**	A	**10**	A	**11**	D	**12**	D

SHORT ANSWER

13 Hydrogen bonding; the highly electronegative nitrogen, oxygen or fluorine atoms cause hydrogen to gain a very strong partially positive charge that attracts the lone pairs of neighbouring nitrogen, oxygen or fluorine atoms.

14 Non-polar molecules can still experience dispersion forces that result from the formation of temporary dipole from the shared electrons.

15 The mobile phase is the solvent that moves over or through the stationary phase during chromatography.
The differences between different components and their affinity for the mobile or stationary phase allows for the separation of mixtures in chromatography. Components with a higher affinity for the mobile phase will travel faster and have a higher tendency to desorb (unstick) off the stationary phase. Components with a higher affinity for the stationary phase will travel slower and have a higher tendency to adsorb (stick) to the stationary phase.

16 By comparing the retardation factor of the impurities with the retardation factors for known substances under the same conditions, the impurities can be identified.

17 Unknown samples can only be identified if they can be compared with known standards that have undergone chromatography under the same conditions.

18 An R_f value of 0 means that the component did not move from the origin, whereas an R_f value of 1 means that the component moved as fast as the solvent front. Both scenarios render the results invalid.

19 Two components that overlap can be separated by two-dimensional chromatography, which involves turning the chromatogram by 90° and performing TLC again with a different mobile phase.

20 If the components of the column chromatography elute from the column at the same time, they can be run through column chromatography again with a different mobile and/or stationary phase to separate them.

21 **a** R_f(green) = 0.15; R_f(purple) = 0.53

b Standard dye C

c Standard dye B

d Standard dye Be Only one of the three standard dyes was present in the sample. The sample also contained a dye that did not match any of the three standard dyes tested.

22 **a** R_f(green) = 0.30; R_f(yellow) = 0.45; R_f(blue) = 0.85

b The mobile phase is polar because it consists of water and ethanol. The stationary phase is paper, making it non-polar.

c Standard dye A

d Standard dye B

e Standard dye B

f Two-dimensional chromatography: rotating the chromatogram by 90° and using a different mobile phase to further separate the compounds. Or test more standard dyes under the same conditions and compare their retardation.

23 **a** A – 14.11 cm; B – 12.58 cm; C – 9.18 cm; D – 5.44 cm

b Mobile phase – polar; Stationary phase – non-polar

c Component A

d Component D

24 a Time how long the solute took to travel down the column (retention time) and collect the solute from the bottom of the column as it passes through.

b Qualitative because the retention time is used to identify each component. This can also be used as a quantitative analysis if each component is collected and weighed separately.

c Green

d Blue solute

e Blue; likely to be polar

f Green

25 By decreasing the polarity of the mobile phase, the sample will have a weaker affinity to the mobile phase and move through the column slower, resulting in a larger R_t value and a better separation.

26 a If the amino acids were analysed in a TLC analysis, then the amino acid with the highest R_f value will be serine, followed by threonine, tyrosine and alanine, because serine has a small polar R group.

b Separated components can be visualised and the chromatogram can be analysed.

c Serine and threonine may overlap due to their similarity in size and polarity.

d After running an initial TLC analysis, select a different mobile phase and turn the chromatogram 90°. This will produce a two-dimensional chromatogram, which can be used to further separate the compounds.

27 Student answers may vary.

UNIT 1 Area of Study 1 Checkpoint

MULTIPLE CHOICE

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

SHORT ANSWER

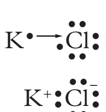
1 a C and D

b B and E

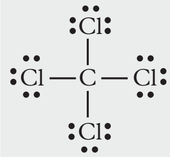
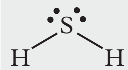
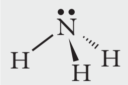
c C and F

2 a $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

b Two potassium atoms lose one electron each to form two positive potassium ions (or cations). Two chlorine gas atoms accept one electron each to form two negative chloride ions (or anions).



3

Molecule	Structural formula	Polar/ non-polar	Type of intermolecular forces
CCl_4		Non-polar	Dispersion forces
H_2S		Polar	Dipole–dipole forces
NH_3		Polar	Hydrogen bonding

4 a Polar

b 0.08

c Yellow – has moved the furthest on the chromatogram, likely due to strong intermolecular forces between the yellow solute and the mobile phase.

5 a Full equation: $H_2SO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + H_2(g)$

Ionic equation: $2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + H_2(g)$

b Full equation: $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

Ionic equation: $Ca^{2+}(aq) + 2OH^-(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

6

Name	Formula
Ammonium acetate	NH_4CH_3COO
Copper(II) hydrogen carbonate	$Cu(HCO_3)_2$
Aluminium sulfite	$Al_2(SO_3)_3$
Magnesium hydride	MgH_2
Lead(IV) oxide	PbO_2
Iron(III) hydrogen phosphate	$Fe_2(HPO_4)_3$

7 a i E

ii H

iii L

iv G

v B

b i C_3F, ionic bonding

ii ED_4, covalent bonding

c Metallic bonding, i.e. an alloy mixture

d Bohr electron shell configuration = 2,8

Subshell configuration = $1s^2\ 2s^2\ 3p^6$

Chapter 7: Quantifying atoms and compounds

GROUNDWORK

7A The number of protons and neutrons in an atom

7B Atoms with the same number of protons but different numbers of neutrons.

7.1 Relative isotopic mass

7.1 SKILL DRILL

1 a 1 proton, 1 electron, 1 neutron

b 6 protons, 6 electrons, 6 neutrons

2 1.1669

3 Relative isotopic mass (RIM) is measured on a standard scale using carbon-12. The relative isotopic mass of the carbon-12 isotope is assigned a value of 12 exactly.

4 2.9926

7.1 CHECK YOUR LEARNING

1 The mass of an isotope atom, relative to the mass of an atom of carbon-12 (exactly 12)

2 There are no units.

3 Mass numbers are only a rough calculation of an atom's mass. They are a whole-number approximation of its relative isotopic mass.

4 Numerically they are equal; the only difference is that the isotope's mass is measured in daltons and the relative isotopic mass has no units.

5 It is more stable than other carbon isotopes carbon-13 or carbon-14, can be cheaply and widely obtained, is safer and has low toxicity to (most) living organisms and is abundant and found easily in nature.

7.2 Relative atomic mass

7.2B CHALLENGE

1 35.5 g mol^{-1}
2 75%
3 25%

7.2 CHECK YOUR LEARNING

1 Separates isotopes in an element sample
2 The number of isotopes in the element sample, relative isotopic mass of each isotope and relative abundance of each isotope
3

Isotope	Relative abundance (%)	% abundance
1	42.1	20.25
2	100	48.10
3	65.8	31.65
Total	**207.9**	**100.00**

4 Percentage abundance – the abundance of each isotope where the sum of all the percentage abundances equals 100% exactly. Relative abundance - the abundance of each isotope where the abundance of the most abundant isotope is set to a value of 100%. The sum of all the relative abundances of all the isotopes will be greater than 100%.
5 Exact values may differ slightly due to the numbers that students read from the graph.

Isotope (from *x*-axis)	Relative abundance (from *y*-axis) (%)	% abundance
^{46}Ti	11	8.03
^{47}Ti	10	7.30
^{48}Ti	100	72.99
^{49}Ti	8	5.84
^{50}Ti	8	5.84

6 183.9 (4 sig fig)

7.3 Avogadro's constant and molar mass

7.3 CHECK YOUR LEARNING

1 The mass in grams of one mole of substance
2 **a** 30.01 g mol^{-1}
b 78.12 g mol^{-1}
c 137.32 g mol^{-1}
3 **a** 0.196 mol (3 sig fig)
b 3.66 mol (3 sig fig)
c 0.00520 mol (3 sig fig)
4 **a** 17.3 g (3 sig fig)
b 167 g (3 sig fig)
c 176 g (3 sig fig)
5 **a** 3.0 mol (2 sig fig)
b 5.0 mol (2 sig fig)
c 25.0 mol (3 sig fig)
6 **a** 18×10^{23} (2 sig fig)
b 30×10^{23} (2 sig fig)
c 151×10^{23} (3 sig fig)
7 3.00 g of water (3 sig fig)
8 120.0 g of sodium chloride

7.4 Percentage composition and empirical formulas

7.4 CHECK YOUR LEARNING

1 %(H in HCl) = 2.74%
%(Cl in HCl) = 97.26%
2 %(C in $C_6H_8O_7$) = 37.5%
%(H in $C_6H_8O_7$) = 4.17%
%(O in $C_6H_8O_7$) = 58.33%
3 **a** CH
b C_5H_{11}
c P_2H_5
4 **a** C_2H_4O
b $C_4H_8O_2$
5 **a** $C_3H_5F_3$
b $C_3H_5F_3$
6 **a** CHO_2
b $C_2H_2O_4$
7 **a** CH_2O
b $C_6H_{12}O_6$
c There is not enough information given to identify the exact compound. Many molecules have the molecular formula $C_6H_{12}O_6$ (e.g. fructose, glucose, galactose).

Chapter 7 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Mass number is the number of protons and neutrons found in atom; relative atomic mass is the weighted average of the relative isotope masses of the element on the ^{12}C scale.
12 Mass numbers are only a rough calculation of an atom's mass. Each isotope has its own isotopic mass. This is because protons and neutrons have very slightly different masses, electrons also have mass, which is not included in mass number, and the binding energy of the nucleus (and other factors) make an isotope's relative isotopic mass slightly different from its mass number.
13 The number of isotopes in the element sample, relative isotopic mass of each isotope and relative abundance of each isotope.
14 **a** 10 mol (2 sig fig)
b 48.0 mol (3 sig fig)
c 6.70×10^{-2} mol (3 sig fig)
15 **a** 180 g mol^{-1}
b 169.9 g mol^{-1}
c 158 g mol^{-1}
16 **a** 0.175 mol (3 sig fig)
b 18.4 mol (3 sig fig)
c 6.80×10^{-5} mol (3 sig fig)
17 46 (2 sig fig)
18 The abundance of ^{121}Sb = 55%.
The abundance of ^{123}Sb = 45%
19 **a** FeC_3O_6
b $Fe_2C_6O_{12}$
20 **a** $LiPF_3$
b $LiPF_3$
21 **a** 3.840 g (4 sig fig)
b $CaSO_4.2H_2O$
22 Student answers will vary but should refer to the use of mass spectrometry in order to determine the relative abundance of each isotope. They should indicate that relative atomic mass is the sum of each relative isotopic mass multiplied by its individual abundance.
23 No. The number of protons in an atom is constant and this identifies the element. but the number of neutrons can vary, resulting in different isotopes of the same element. The number of neutrons does not affect the element's physical properties.
24 The volume of the cube is $2 \times 2 \times 2 = 8$ cm^3
Find the mass, and calculate the number of moles ($n = m/M$). Use the number of moles to find the number of atoms by multiplying by Avogadro's constant. Divide the volume of the 8 cm^3 cube by the number of atoms to find the volume of one atom of zinc.

Chapter 8: Families of organic compounds

GROUNDWORK

8A It is important in chemistry to ensure that safer and more sustainable products and processes are designed to minimise risks and hazards, the amount of materials and energy consumed and unwanted wastes. This can be achieved through a circular economy where resources are used, recovered and reused.
8B Covalent
8C The shape of a molecule helps determine its polarity, which determines the intermolecular forces present and its physical properties (e.g. melting point, boiling point, solubility).

8.1 Grouping hydrocarbon compounds

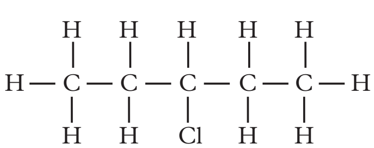
8.1 CHECK YOUR LEARNING

1 An organic compound that contains one or more carbon atoms bonded to hydrogen atoms
2 Student answers will vary but should include the functional group and how it affects the properties of the compound.
3 Compounds with the same molecular formula, number and type of atoms, but arranged differently
4 **a** C_6H_{14}
b C_2H_5COOH
c $C_8H_{17}OH$
d C_3H_8
e C_3H_6
5 $CH_3CH_2CH_2CH_2CH_2OH >$ $CH_3CH_2CH_2CH_2CH_2Cl >$ $CH_3CH_2CH_2CH_2CH_3 >$ $CH_3CH_2CH{=}CHCH_3$
6 Carboxylic acids have higher boiling points than other hydrocarbon families because two molecules of a carboxylic acid form two hydrogen bonds with each other.
7 **a** Alcohol
b Haloalkane
c Alkane
d Haloalkane
e Carboxylic acid
f Alkene
8 Student answers will vary but should include two structural isomers with the same molecular formula but different structures, then discuss how the molecular shape affects the packing (and therefore, boiling point) of molecules.
9 Alcohols contain the polar hydroxy (OH) group, which forms strong hydrogen bonds with water. Alkanes are non-polar and are insoluble in water.

8.2 Representing organic compounds

8.2 CHECK YOUR LEARNING

1 So that chemists can communicate, name and identify organic compounds in a consistent way
2 **a** Octane
b 2,4-Dimethylpentane
c 4,5-Difluorohexanoic acid
d Oct-3-ene
3 **a**



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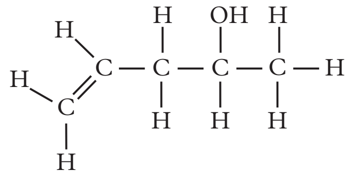
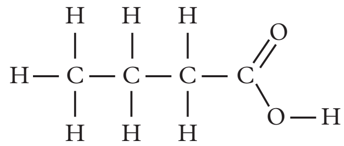


c



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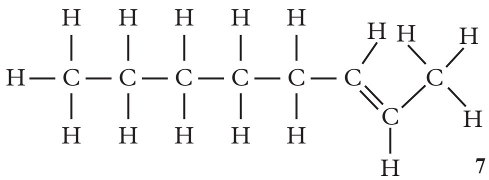

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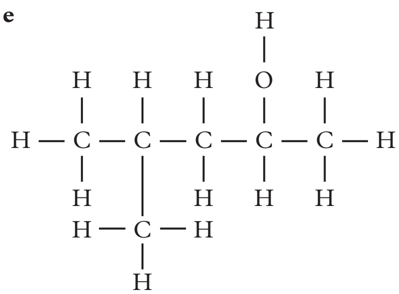
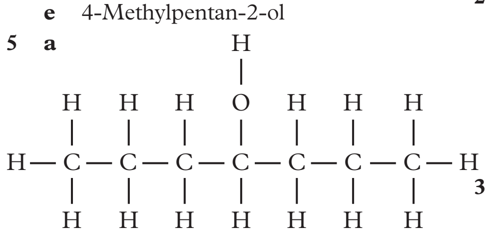
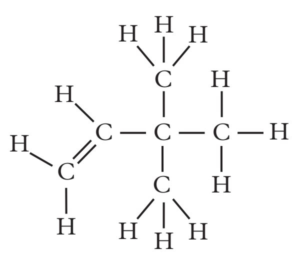


4 **a** 3-Chloropentane
b Pent-4-en-2-ol
c Butanoic acid
d Oct-2-ene
e 4-Methylpentan-2-ol
5 **a**



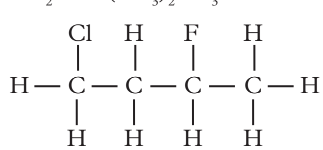
$CH_3(CH_2)_2CH(OH)(CH_2)_2CH_3$

b $CH_2CHC(CH_3)_3$



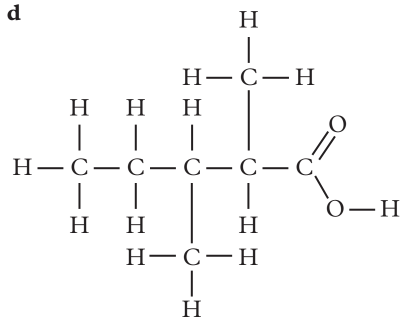
$CH_2CHC(CH_3)_2CH_3$

c



$CH_2ClCH_2CHFCH_3$

d



$CH_3CH_2CH(CH_3)CH(CH_3)COOH$

6 Molecular formulas only show the number and types of atoms in the molecule. Structural formulas show the bonds between all atoms as lines, and semi-structural formulas show the positions of all the atoms and functional groups, but not the bonds.
7 The molecular formulas for the two molecules are identical ($C_4H_8O_2$) and therefore, cannot be used to distinguish between the molecules.

8.3 Renewable sources of organic compounds

8.3 REAL-WORLD CHEMISTRY

1 They can all be used to produce plastic backsheets; however, fossils fuels are a non-renewable resource (not sustainable), while cotton rags and castor beans are renewable resources (more sustainable and long-lasting).
2 The use of cotton rags and castor beans is a shift away from a linear economy; however, it is not at the stage of a circular economy because resources are not used and reused resources repeatedly during production and consumption.
3 Use of renewable feedstocks, designing safer chemicals, design for degradation
4 They ensure that renewable resources are used instead of non-renewable resources, which are extracted from the Earth to produce a product.

8.3 CHECK YOUR LEARNING

1 A mixture of hydrocarbons made from decomposed plants and animals, found and mined from the Earth's crust.
2 Wood and wood processing wastes, agricultural crops, waste materials, animal manure and human sewage and recycled products
3 Plant-based biomass is material from plants that contains stored chemical energy from the Sun.
4 Fossils fuels are non-renewable resources that are available in limited supply and their use produces chemicals that are toxic to the environment.

Plant-based biomass products are renewable resources that can be grown at the same or a faster rate than they are consumed and are not toxic to the environment.

5 Student answers will vary but should include that a linear economy uses non-renewable, finite resources that are extracted from the Earth to make products that will be thrown away; however, in a circular economy, renewable resources are used sustainably.

6 Renewable feedstocks are made from sustainable (e.g. plant-based) materials that can easily break down in the environment into harmless products that do not accumulate in and harm the environment.

8.4 Organic compounds in everyday life

8.4 REAL-WORLD CHEMISTRY

1 Azo direct dyes have been linked to cancer, mutations and negative reproductive effects in humans and azo reactive dyes to increased allergy risk. They don't biodegrade and they accumulate in the marine food chain, which affects humans who consume food from the sea.

2 Azo dyes give long-lasting, bright and vibrant colours to the fabrics and manufacturers are able to mass produce clothing at low cost.

3 Design for degradation, designing safer chemicals, prevention of wastes

4 They help to protect the oceans, seas and marine resources.

8.4 SKILL DRILL

1 The conclusion does not use scientific terminology/supporting figures, supporting facts/quotes from experts or provide data that is evaluated.

2 Industries will be affected, which in turn will affect the people's livelihoods. For example, industries may not be able to access existing alternative chemicals that do not bioaccumulate.

3 Any decision needs to rely on the most current scientific evidence because it can affect society (humans and the environment) for many years.

8.4 CHECK YOUR LEARNING

1 A compound that contains carbon

2 Natural organic compounds come from plants and animals. Synthetic organic compounds are produced by chemical reactions in the laboratory.

3 **a** It contains five electronegative O atoms and one electronegative N atom, which can form strong hydrogen bonds with water molecules.

b Glyphosate is soluble in water; therefore, it can be absorbed by invasive and toxic weeds and pests.

4 Student answers will vary but should include a discussion of the risks and benefits of using organic chemicals and state if the benefits outweigh the risks.

5 All food additives and cosmetics are tested to ensure they are non-toxic and safe for ingestion or when applied on skin. Pesticides are sprayed into the environment and do not come in direct contact with humans in the concentrations that organic compounds in food and cosmetics do.

6 Many organic chemicals are toxic to humans and the environment. Therefore, biodegradable organic products should be designed so that they break down into harmless products and do not bioaccumulate in the environment.

Chapter 8 review

MULTIPLE CHOICE

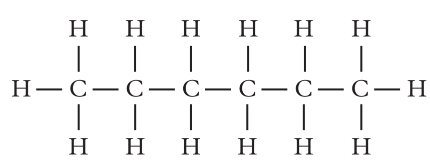
1	A	**2**	A	**3**	A	**4**	B
5	B	**6**	C	**7**	C	**8**	D

SHORT ANSWER

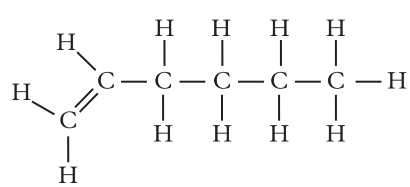
9 Carbon can form single, double and triple bonds. It can be connected to other carbons in chains of different lengths and in different shapes, as well as to other atoms.

10 Fluoromethane contains carbon bonded to a fluorine atom, which can form strong hydrogen bonds between the molecules. Hydrogen bonds require a large amount of heat energy to break them compared with the weak dispersion forces between methane molecules.

11 **a** Hexane



Hex-1-ene



b Hexene has a lower boiling point because the double bond in hexene affects the packing of the molecules and the dispersion forces between them.

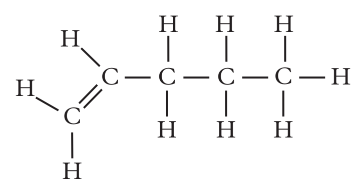
12 1-Chloropropane has three carbon atoms in the carbon chain; therefore, the chlorine atom could be found on carbon-1 or carbon-2. Chloroethane has only two carbon atoms in the carbon chain; therefore, the chlorine atom can only be found in one position on the carbon chain.

13 Ethanol contains a hydrogen atom bonded to an oxygen atom, which form strong hydrogen bonds between the molecules that require a large amount of heat energy to break.

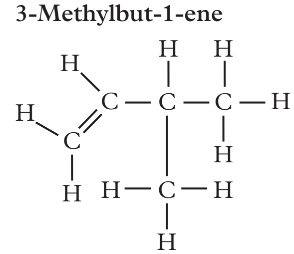
14 In carboxylic acids, the position of the carboxyl functional group cannot change. It can only be at the end of the carbon chain because to form one double bond to an oxygen and a single bond to a hydroxyl group, the carbon must be an end carbon.

15 Together, they ensure that renewable resources are used sustainably, instead of using non-renewable resources which are in limited supply.

16 **a** Pent-1-ene



3-Methylbut-1-ene

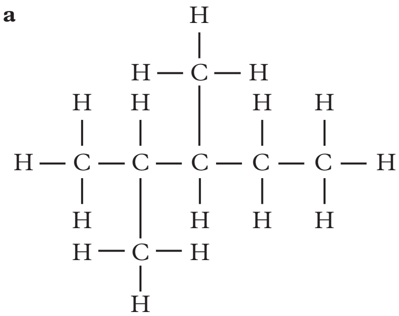


b 3-methylbut-1-ene contains a branch in its structure, which means the molecules cannot pack as tightly and the dispersion forces between the molecules will be reduced.

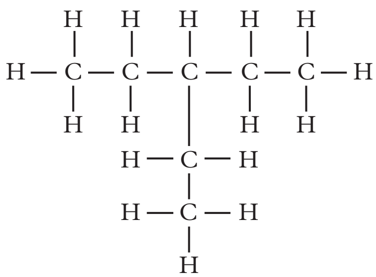
17 **a** Alkene
b Alkane
c Alcohol
d Haloalkane
e Alkene
f Carboxylic acid

18 **a** 2,2-Dimethylpentane
b 2-Chloropentane
c Hex-3-ene
d 2,3-Dimethylbutane

19 a



b



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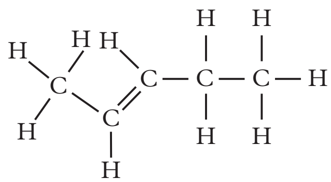

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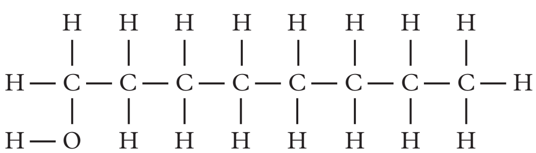

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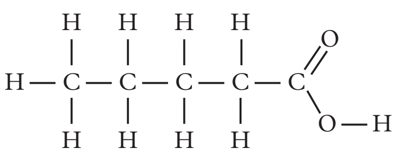

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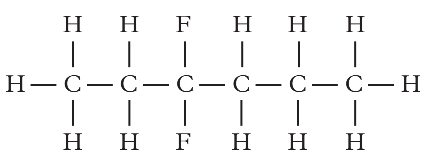


20 a (a) C_7H_{16}, (b) C_7H_{16}, (c) C_5H_{10}, (d) $C_8H_{18}O$, (e) $C_5H_{10}O_2$, (f) $C_6H_{12}F_2$

b (a) $CH_3CHCH_3CHCH_3CH_2CH_3$,
(b) $CH_3CH_2CHC_2H_5CH_2CH_3$,
(c) $CH_3CH{=}CHCH_2CH_3$,
(d) $CH_2OH(CH_2)_6CH_3$,
(e) $CH_3(CH_2)_3COOH$,
(f) $CH_3CH_2CF_2(CH_2)_2CH_3$

21 Hydroxy OH, alkane C–C, alkene C=C

22 a 3,3-Dimethylpent-1-ene

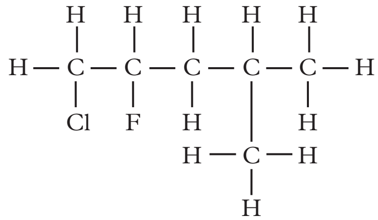
b 2-Chloro-3,4-dimethylpentane

c 4-Fluorohex-2-ene

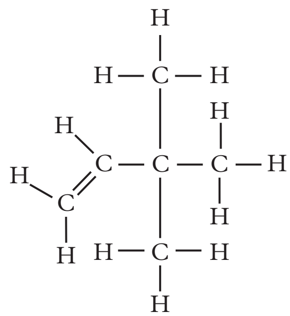
d 2-Methylbutan-2-ol

e But-3-en-2-ol

23 a



b



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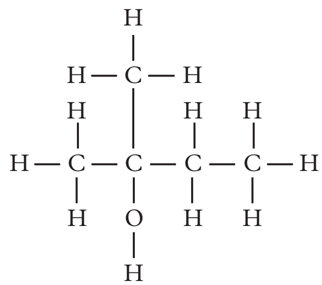

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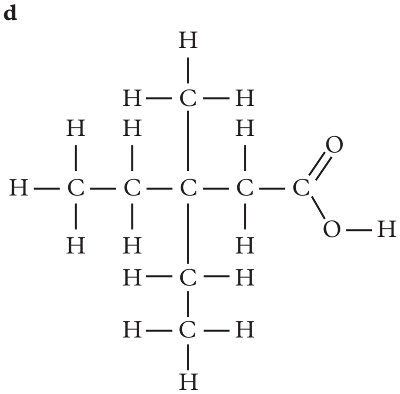

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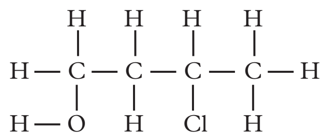
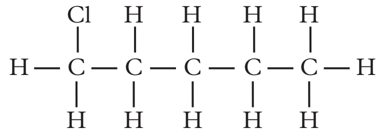


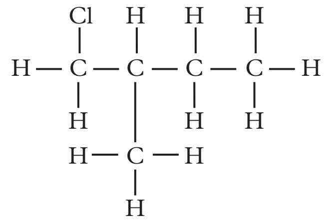
24 a (a) $C_6H_{12}ClF$, (b) C_6H_{12}, (c) $C_5H_{12}O$, (d) $C_8H_{16}O_2$, (e) C_4H_9ClO

b (a) $CH_2ClCHFCH_2CHCH_3CH_3$,
(b) $CH_2{=}CHC(CH_3)_2CH_3$,
(c) $CH_3CCH_3OHCH_2CH_3$,
(d) $CH_3CH_2C(CH_3)(CH_2CH_3)CH_2COOH$,
(e) $CH_2OHCH_2CHClCH_3$

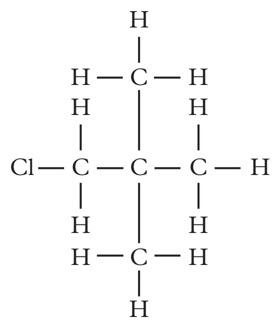
25 a Three possible isomers are shown below



1-Chloropentane



1-Chloro-2-methylbutane



1-Chloro-2,2-dimethylpropane

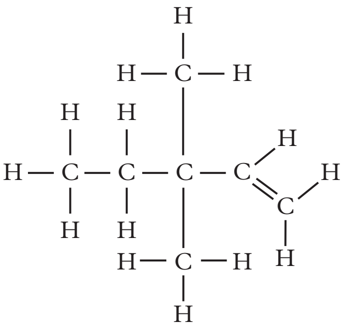
b 1-Chloropentane, 2-chloropentane, 3-chloropentane, 1-chloro-2-methylbutane, 2-chloro-2-methylbutane, 2-chloro-3-methylbutane, 1-chloro-2,2-dimethylpropane, 1-chloro-3-methylbutane

26 Ethanol and octanol are polar because they both contain a hydroxyl group (OH). However, ethanol is more soluble than octanol because it has a shorter non-polar hydrocarbon chain than octanol and can be more easily surrounded by water molecules.

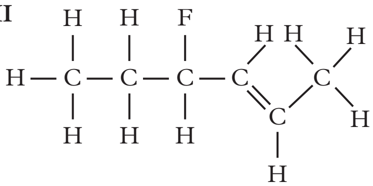
27 Propanoic acid is more soluble than propane because it has a polar carboxyl group (COOH) with two electronegative oxygen atoms, which can form strong hydrogen bonds with water. Propane is non-polar; therefore, it is insoluble.

28 a

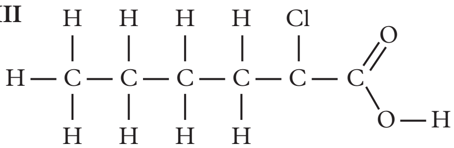
I



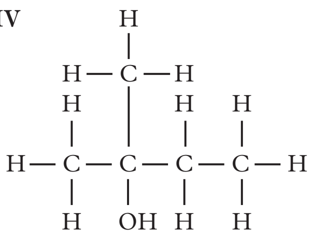
II



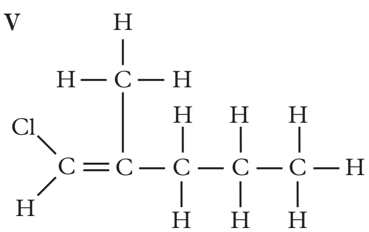
III



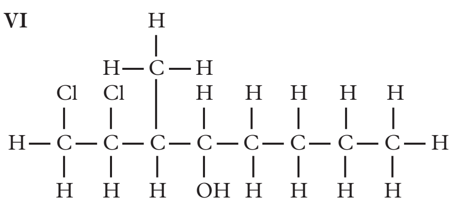
IV



V



VI



b I incorrect, II incorrect, III incorrect, IV incorrect, V incorrect, VI correct

c I – the double bond takes priority over the alkyl groups; therefore, carbon-1 begins at the double bond

II – the double bond takes priority over the halogen functional group; therefore, carbon-2 begins at the double bond

III – all carboxyl groups are located on an end carbon, so it is not assigned a number

IV – Should be numbered from the other end of the chain to obtain the lowest numbers possible.

V – the halogen function group takes priority over the alkyl group so is named first

d I 3,3-dimethylpent-1-ene,
II 4-fluorohex-2-ene,
III 2-chlorohexanoic acid,
IV 2-methylbutan-2-ol,
V 1-chloro-2-methylpent-1-ene

29 Student answers will vary.

30 a Organic molecules that have different structures and contain different functional groups have different physical properties (e.g. boiling point).

b Pentane has only dispersion forces between its molecules; therefore, it has a low boiling point. Methanoic acid has a polar carboxyl functional group, which forms two hydrogen bonds with carboxyl groups; therefore, it has a higher boiling point.

c Ethanol has a higher boiling point than fluoroethane because it has strong hydrogen bonding between its molecules, whereas fluoroethane has weaker dipole–dipole attractions between its molecules.

d The boiling point of propanol would be between those of ethanol and methanoic acid.

31 In a circular economy, renewable resources such as plant-based biomass are used sustainably to produce isododecane, whereas in a linear economy, non-renewable finite resources such as fossil fuels are extracted from the Earth to make products that will be thrown away.

32 Using biomass to produce isododecane ensures the sustainable consumption of resources.

33 Chlorpyrifos accumulates in the environment and is toxic to many living things, including humans. Another insecticide (e.g. glyphosate) could be used instead because it can break down in the environment and does not bioaccumulate.

Chapter 9: Polymers and society

GROUNDWORK

9A An organic compound that contains one or more atoms of carbon covalently linked to atoms of other elements, most commonly hydrogen, oxygen, or nitrogen.

9B A polymer is a natural or synthetic substance composed of very large molecules formed by one or multiple repeating units of simpler molecules called monomers.

9C Biomass is plant-based material used as fuel to produce heat or electricity. Renewable feedstock is a natural resource that can replenish itself in a limited time, preferably within several months.

9D Moving towards a more circular economy could deliver many benefits for the consumer and the environment such as:

- reducing pressure on the environment
- improving the security of the supply of raw materials
- increasing competitiveness
- stimulating innovation
- boosting economic growth and creating jobs.

9.1 Addition and condensation polymerisation reactions

9.1 CHECK YOUR LEARNING

1 Because monomers are the simpler chemical units that join together via covalent bonds to form polymers.

2 Each glucose monomer has multiple hydroxyl (OH) groups that react to form water as a by-product and –O– bonds between monomers. The –O– bond links form either the starch or cellulose polymer (this depends on the monomer of glucose).

3 Monomers used in condensation polymerisation contain functional groups, such as COOH and OH groups, that can lose a water molecule. Monomers used in addition polymerisation reactions contain double-bonded carbons that can be opened up to form a single carbon–carbon covalent bond with other monomers.

4 a Reaction I = addition polymerisation. Reaction II = condensation polymerisation.

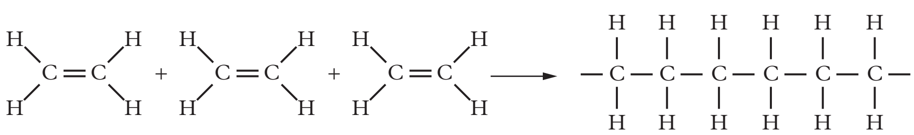
b In addition polymerisation, monomers react to form a polymer without the formation of by-products. For example, in the addition polymerisation of ethene, the carbon–carbon double bond of one ethene molecule was broken to connect with the second ethene molecule. All double bonds became single bonds in the final polymer product and no by-products were formed. In condensation polymerisation, monomers react with each other to form larger structural units while releasing smaller molecules as a by-products such as water or methanol.

5 Polypeptides are an example of natural condensation polymers that are formed by condensation polymerisation with amino acid monomers. In this reaction, the NH_2 of one amino acid monomer reacts with the COOH of another monomer to produce water (the by-product). A covalent bond then forms between the outer carbon and nitrogen atoms of the monomers, linking together the polypeptide polymer.

The synthetic condensation polymer, polyethylene terephthalate (PET), is also formed by condensation polymerisation; however, the monomers and linking bonds produced differ between the two. In the PET reaction, the carboxylic acid and hydroxyl functional groups on monomers react in condensation reactions to form water and –COO– bonds that link polymer chain.

6 Two or more different monomer units are required for an addition copolymer to be formed. By having a combination of different monomers, different properties are transferred to the polymer.

7



9.2 Formation of addition polymers

9.2 REAL-WORLD CHEMISTRY

1 Heat resistance comes from acrylonitrile. Impact strength and toughness come from butadiene. Mouldability and rigidity come from styrene.

2 Toy bricks need to be strong and tough upon impact and handling, a property that comes from butadiene. They also need to be mouldable while manufactured and rigid when Lego product is formed, a property that comes from styrene. On moderate heating and change of temperature, toy bricks also need to be resistant to heat, a property that comes from acrylonitrile.

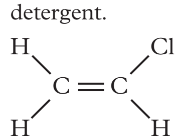
3 Increasing the proportion of each monomer would strengthen the corresponding properties. For example, if the proportion of acrylonitrile is increased, the ABS polymer will become more heat resistant. If butadiene is increased, the ABS polymer will become tougher. If styrene is increased, it becomes more rigid.

9.2 CHECK YOUR LEARNING

1 The double carbon–carbon bond must break before the monomers can be combined to make addition polymers. Breaking the bond makes it possible to covalently bond to another monomer.

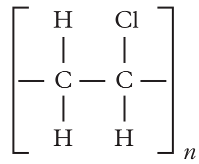
2 PVOH is very water soluble so once in contact with water, the active ingredient in dishwashing tablets can be released. It is also colourless and odourless, meaning it does not affect the appearance and properties of the detergent.

3

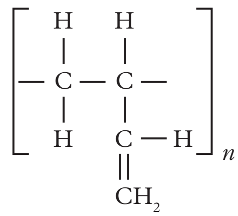


4 a Molecules I and II can undergo addition polymerisation reactions with the presence of the carbon–carbon double bond. Molecule III cannot undergo addition polymerisation reactions.

b (I)



(II)



5 1 344 328 g mol^{-1}

6 a In PVCA, vinyl chloride brings durability, which makes PVCA very tough and durable. Vinyl acetate has UV-resistant properties, which makes PVCA stronger in this aspect.

b When the percentage of vinyl chloride increase, PVCA becomes tougher and more durable. When the percentage of vinyl acetate increases, the UV-resistant property of PVCA increases.

c PVCA is tough and durable. It is durable while being manufactured, which allows the printing of card number and information to the card. It is also tough, which allow the card to be strong and resist breaking.

9.3 Linear and cross-linked addition polymers

9.3 CHECK YOUR LEARNING

1 Thermoplastic polymers are linear polymers. They have no strong chemical bonding such as covalent bonding between polymer chains. There are only hydrogen bonding, dipole–dipole attractions or dispersion forces to attract the linear polymers chains together.

2 Thermosetting polymers are cross-linked polymers. They have strong covalent bonds between polymer chains, which are known as cross-links

3 Elastomers have occasional cross-links between the chains but far fewer than thermosetting polymers. This means that chains can move across and past each other but can also return to their original position once the impact is removed.

4 Both are polymers used frequently in manufacturing. However, thermoplastic polymers have low softening points

and are resistant to impact, whereas thermosetting polymers have much higher softening and melting points and are much more heat resistant.

5 Thermoplastic polymers are linear polymers, which means different polymer chains are joined by intermolecular forces such as hydrogen bonds instead of strong covalent bonds. Intermolecular forces are easier to break at lower temperatures. Therefore, they are ideal for 3D printing because the temperature at the 3D printing nozzle is sufficient to break the intermolecular forces to soften the material. Soft material can also be easily re-shaped into the desired product.

6 Thermosetting polymers have excellent heat and electrical insulator properties. They have strong covalent bonds between polymer chains, making the polymer have higher softening and melting points. They can also be moulded into different thicknesses and tolerances depending on the application need.

7 Elastomers have occasional cross-links between the chains but far less than thermosetting polymers. They can move across and past each other but can also return to their original position once the impact is removed. Therefore, cross-links stop elastomers from melting when heated, but the small number of links allows it to be re-formed, making it a perfect material for the inner tube in a bike tyre.

9.4 Features of linear addition polymers

9.4 REAL-WORLD CHEMISTRY

1 A copolymer is made of two monomers. A designer polymer is a copolymer made with a specific purpose in mind where properties of the targeted polymer are first identified before choosing the type of monomers.

2 HEMA is hydrophilic and can absorb several times its weight in water. This makes a contact lens soft, comfortable and compatible with the biology of the eye. Methyl methacrylate meant lenses were transparent and lightweight.

3 Hydrogel polymers can absorb water at several times of its weight. This makes it a perfect material for nappies and sanitary pads because they can absorb a lot of fluids.

9.4 CHECK YOUR LEARNING

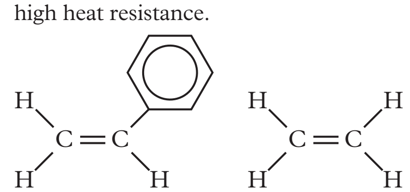
1 As chain length increases, the molecular mass of the polymer increases, which results in the intermolecular forces between the polymer chains increasing. This means polymers with longer chains are stronger and have higher melting and boiling points.

2 The more branches a polymer has, the more difficult it is for chains to pack in closely together, which weakens intermolecular forces. Weaker intermolecular forces result in lower strength of a polymer.

3 Structures with amorphous regions are less able to pack closely together. These structures have weaker intermolecular forces between the chains and are therefore less rigid, weaker and often transparent materials.

4 Both LDPE and HDPE are formed from the polymerisation of ethylene molecules. However, LDPE forms rapidly, which allows for the formation of branches off the main linear chain, whereas HDPE has limited branching because of the low formation temperature. Additionally, the chains of LDPE cannot pack together tightly and dispersion forces between the chains are weaker, whereas HDPE polymers are more rigid as chains can pack closely together. LDPE is flexible and has excellent ductility but low tensile strength. HDPE is a moderately robust and stiff plastic that is strong and has high heat resistance.

5



6

Advantages/ Disadvantages	LDPE	HDPE
Flexibility	Less crystalline structure therefore more flexible	High crystalline structure therefore tougher and more rigid
Chemical resistance	Resistant to most alcohols, acids and bases. Low to no resistance to oxidising agents and organic solvents	Superior resistance to organic solvents, alcohols, acids and bases
Strength	Weaker than HDPE because of the branching and the weaker intermolecular forces	Strong as a result of the linear structure and stronger intermolecular forces
Heat resistance	Density drops at higher temperatures but retains strength and flexibility over a wide range of temperatures	Useful over 100°C
Transparency	Highly transparent due to low crystalline structure	Low transparency due to increased level of the crystalline structure

7 a Nylon does not have great heat-resistant properties. Nylon is also wrinkly and does not maintain the shape of the jacket.

b Nylon is a copolymer. It is an abrasion-resistant fabric. It is resistant to stains, heat, UV rays and chemicals. Polyester is elastic, which makes the jacket stretchy. Polyesters are very resilient and durable. The fabric is water resistant. Polyesters are also resistant to mould and mildew.

Polyurethane is resistant to water, resilient and strong, and performs well in harsh environments. It is also resistant to mould and fungus, which makes it suited as the inner layer of jackets.

Teflon has excellent thermal insulation properties, which means the wearer can maintain body heat, which is great for a jacket used in colder weather.

c Nylon or polyester provides an outer layer of protection for the jacket. It makes an abrasion-resistant outer shell.

Polyurethane and Teflon both ensure the breathability of the jacket and make sure water evaporates in rainy weather. Polyurethane keeps the body temperature constant and makes the jacket suitable to be used in cold weather.

Teflon is a more rigid layer, which gives protection against wind and water.

9.5 Categorising plastics

9.5 SKILL DRILL

1 Compostable bioplastics are an example of a circular economy because after use, they can be decomposed by living organisms such as bacteria or fungi and disintegrate back into the Earth. It addresses the green chemistry principles by using materials that do not consume the Earth.

2 Growing crops such as sugar cane can make polymers for films, drink containers, fuel tanks, injection moulded parts and tubes. It is a green chemistry principle because the materials are biodegradable and can disintegrate back into the Earth. However, if farms are used for plastics rather than for growing food, a food crisis might occur, so farmlands need to be allocated sustainably.

3 Recycling fossil-fuel plastics by chemical and mechanical recycling makes use of the plastics again but the plastic is often made into lower-quality polymers or in lower amount. It is an example of a circular economy because it reuses already formed plastics to make the most of the resources. It also relates to the green chemistry principles because it decreases the need to use new resources.

9.5 CHECK YOUR LEARNING

1 Biodegradable plastics can be decomposed by living organisms and under the right conditions will break down and disintegrate back into the Earth. The process may take months or years depending on factors such as chemical structure and surrounding conditions.

2 Mechanical recycling is a process used to recover plastic waste by mechanical processes that include sorting, washing, drying, grinding, compounding, re-granulating or pelletising.

3 Depending on the material to be recycled and the intended use of the plastic, different categories are needed to recycle plastics. Mechanical recycling is used to recover plastic waste by mechanical processes. The order and number of times each of these processes will occur depends on the type of plastic. In chemical recycling, polymers are broken down into monomers and re-form into new polymers for new use. Organic recycling includes breaking down biodegradable by living organisms such as bacteria.

4 Student answers will vary.

5 Student answers will vary but should include using renewable feedstocks, eliminating waste and avoiding emission of harmful materials.

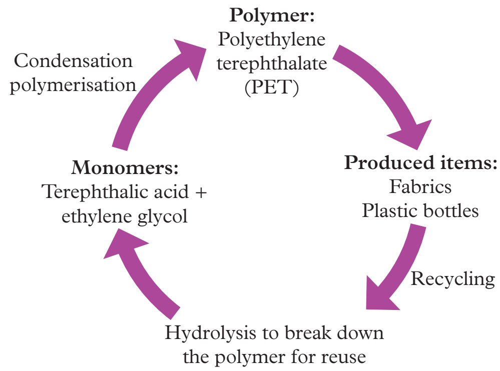
6 Chemical and mechanical recycling can both produce products of high quality. However, products formed from chemical recycling can range from being the same, similar or lower quality as the original polymer plastic. This is because the materials are split into monomers or feedstock that can then be transformed into secondary raw materials. All thermoplastics can be mechanically recycled with often limited impact on the quality of the recycled plastics. Plastics that can be mechanically recycled to high-quality raw materials include polyethylene (PE), polypropylene (PP) and polyethylene terephthalate (PET).

7 Mechanical and chemical recycling are examples of a circular economy because they reuse materials and minimise the need to use more new materials for the same purpose. They decrease the consumption of new materials and reduce waste.

8 Organic recycling is an example of a circular economy because the biodegradable plastics are broken down by living organisms. Then the products disintegrate back into the Earth and can become part of the food chain. New renewable resources such as food can be made again with the nutrition and made into new plastics.

9.6 Innovations in polymer manufacturing

9.6 CHALLENGE



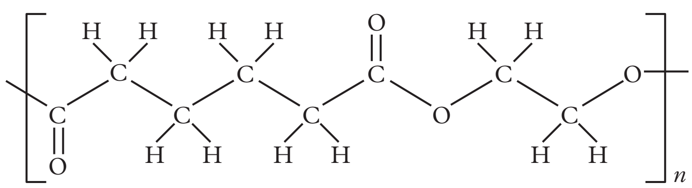
9.6 CHECK YOUR LEARNING

1 A condensation reaction is a reaction between a carboxyl and a hydroxyl functional group. They react to form a –COO– group and water as the by-product.

2 A hydrolysis reaction is the opposite of a condensation reaction because instead of forming an ester group, an ester group is broken with the addition of water. The carboxyl side of the ester group gains the –OH from water to form a carboxyl group and the hydroxyl side gains the –H from water to form a hydroxyl group.

3 An advantage of mechanical recycling is that it is a relatively simple process and requires less time and equipment. However, after mechanical recycling, PET can only be reused for similar purposes and does not produce useful products such as oil and energy. In Cat-HTR recycling, hydrolysis produces the original monomers, which can be put back into the plastic manufacturing stream to create new plastics.

4



5 Student answers will vary but should include that PET is durable, is soft and can be drawn and spun into fibres, allowing it to be easily handled and manufactured into different shapes.

6 It is not an example of a circular economy. *Ideonella sakaiensis* breaks down PET by hydrolysis reactions. The bacteria complete their hydrolysis reactions through the use of two enzymes that help break up the polymer into its monomers of terephthalic

acid and ethylene glycol. Since these products are used by the bacteria the products cannot go back into the plastic manufacturing process.

7 Both methods address the sustainable goal of responsible consumption and production. Mechanical recycling is the method of recycling PET by mechanical sorting. The sorted material can then be used for new purposes. Cat-HTR recycling is the process of performing hydrolysis reaction to break down PET into the monomers, which can then be put back into the plastic production cycle. They are two different methods but they both can relieve the reliance on new fossil fuels.

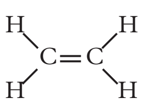
Chapter 9 review

MULTIPLE CHOICE

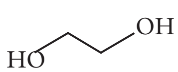
1	A	**2**	D	**3**	B	**4**	A	**5**	A
6	C	**7**	B	**8**	C	**9**	C	**10**	A

SHORT ANSWER

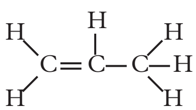
11



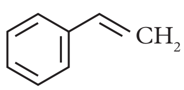
Ethene



Ethylene glycol



Propene



Phenylethene

Ethene, propene and phenylethane can be used in addition polymerisation because they have double carbon–carbon bonds. Ethylene glycol can be used in condensation polymerisation because of the two hydroxyl groups.

12 Less branching means that the chains can pack in closer and the intermolecular forces between chains are stronger. Therefore, the polymer is stronger.

13 An addition copolymer is formed by the addition polymerisation of at least two monomers. Each monomer must have a double carbon–carbon bond. On polymerisation, the double bonds are broken and connect to the next molecule to form polymers.

14 Monomers must have at least two functional groups per molecule. The functional group of one monomer reacts with a functional group on another monomer and produces a by-product, which breaks off.

15 Biodegradable plastics can be decomposed by living organisms and under the right conditions will break down and disintegrate back into the Earth. A fully compostable substance must be able to break down fully without leaving any traces. Fully compostable substances must also not have any toxic elements in their structure. Some biodegradable plastics cannot completely break down, and can leave behind residues, or microplastics that accumulate in the food chain.

16 Increasing branching makes the polymer less organised because it is harder for the polymer chain to pack together. Better packing brings the chains closer together, increases the dispersion forces, and strengthens the material. Therefore, increasing branching decreases the crystallinity of a polymer. The crystallinity of I is high because of the lack of branches and the close packing. The crystallinity of II is low because the branches make it hard to pack tightly.

17 Polymer chain I has a high melting point and polymer chain II has a low melting point. Polymer chain I has fewer branches and packs together better, therefore has greater intermolecular forces. It takes more energy to break the bonds and therefore the melting point is higher. Polymer chain II has more branches and cannot pack closely; therefore, it has weaker intermolecular forces. It takes less energy to weaken/break the intermolecular forces and therefore the melting point is lower.

18 a Reaction I is addition polymerisation and reaction II is condensation polymerisation.

b Propene nitrile and phenylethene (styrene)

c Polypeptide

d Copolymers are manufactured for a specific use. Each monomer brings a desired property to the polymer so a material can meet specific needs. Some of the unwanted properties of a particular monomer can also be overcome by using another monomer.

19 The manufacturing of fossil fuels to make products such as plastic bottles, followed by usage and direct disposal is an example of a linear economy. The process consumes fossil fuels and exhausts the resources because there is no reuse of the products synthesised.

20 The recycling of plastics enables materials from the products to be diverted back into the manufacturing process, forming a cycle of reuse. This is achieved through mechanical recycling or chemical recycling where polymers are turned back into monomers and re-manufactured into new plastics.

21 In a condensation polymerisation, monomers lose a small molecule such as water or methanol to form functional groups such as ester groups or peptide bonds. In addition polymerisation reactions, double carbon bonds are converted to carbon–carbon single covalent bond without releasing any small molecules.

Monomers with carboxyl and hydroxyl groups with the potential to lose a small molecule in the polymerisation reaction are needed for condensation. A catalyst is also required.

In addition polymerisation reactions, unsaturated molecules are required. UV light or high temperature and high pressure is usually needed to catalyse the reaction.

Similarities are that both reactions can form large molecules. Differences are that a small molecule is lost in condensation reactions but not in addition reactions. The functional groups needed in the monomers are also different.

22 a $312\,500\,g\,mol^{-1}$ (4 sig fig)

b $6\,480\,000\,g\,mol^{-1}$ (3 sig fig)

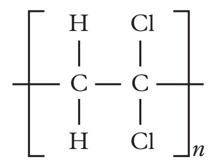
c $7\,100\,000\,g\,mol^{-1}$ (2 sig fig)

23 a Addition: I, III

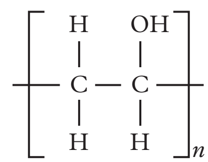
Condensation: II, V

Neither: IV

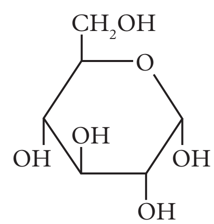
b I



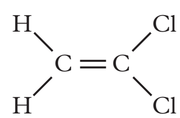
III



24 a I



II



b I is a condensation polymer.

II is an addition polymer.

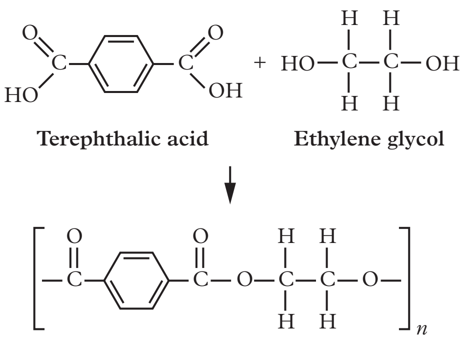
25 i Polymer b is a thermoplastic polymer.

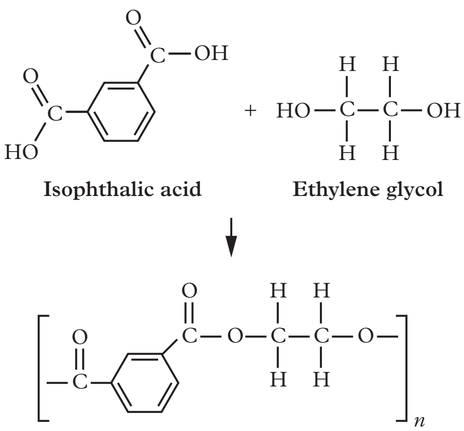
ii Polymer a is a thermosetting polymer.

iii Due to cross-links, the thermosetting polymer has a higher melting point than thermoplastic polymer. The is because greater intermolecular forces are provided by the covalent bond. Thermoplastic polymers are easier to recycle because they require less energy to break down to smaller monomers because of their weaker bonds compared with breaking down all the cross-linking in thermosetting polymers.

iv Polymer a is a high-density polymer, which is rigid and resistant to a wide range of conditions. Polymer a could be used for fuel tanks, boat parts, 3D printing filaments and strong packaging materials. Polymer b is a low-density polymer, which is likely to be soft. It can be used for soft packing materials, bottles and waterproof layers.

26 a Terephthalic acid is a symmetrical molecule because the carboxyl groups are directly opposite eachother but isophthalic acid is not symmetrical. When terephthalic acid reacts with ethylene glycol, the polymer they they form maintains a linear shape. However, when isophthalic acid reacts with ethylene glycol, the polymer is kinked because of the placement of the carboxyl groups on the monomer.





b Isophthalic acid would make the structure more amorphous because it is kinked. The kink makes it harder for the chains to pack together and therefore they are less organised and crystalline.

c Because of the kink and the polymer being less tightly packed, using isophthalic acid would decrease the intermolecular forces between polymers and hence decrease the melting point.

27 Recycling plastic is only the starting point to help us achieve the United Nations Sustainable Development Goal 12. It alleviates the burden of breaking down plastics and our reliance on the original source, fossil fuels. However, to achieve the goal, more efforts need to be made to improve the recycling method such as research on chemical reactions that do not require high temperature and pressure to break down polymers. Other methods such as decreasing the use of packing from companies also need to be taken to increase public awareness.

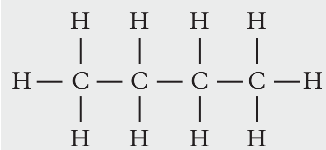
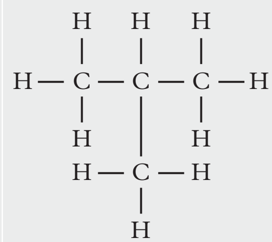
UNIT 1 Area of Study 2 Checkpoint

MULTIPLE CHOICE

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

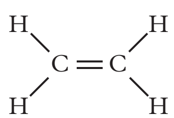
SHORT ANSWER

1 a i A, D, G and J
ii C_nH_{2n+2}
b i J
ii

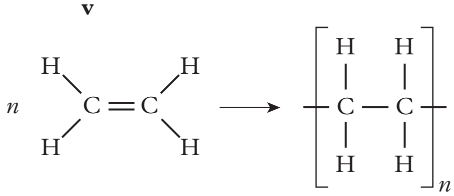
Name: butane	Name: 2-methylpropane
Structural formula: 	Structural formula: 

iii An isomer contains the same ratio of atoms (molecular formula) but they have different structural arrangements of those atoms.

c i C
ii ethene
iii



iv $C_2H_4 + Br_2 \rightarrow CH_2BrCH_2Br$
v

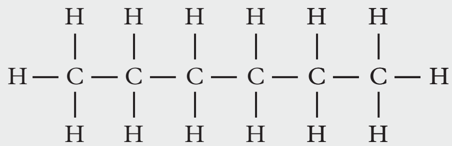
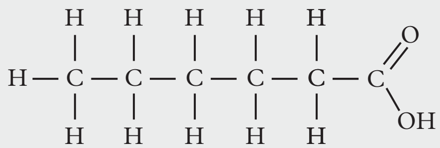
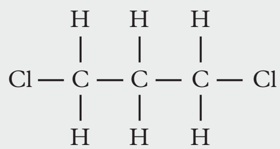
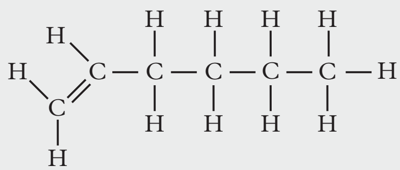
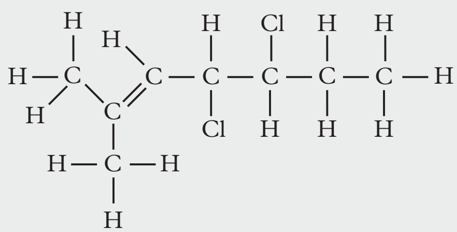


vi Polyethene

2 a C_9H_9N
b 2.2×10^{-22} g per molecule (2 sig fig)
c 1.2×10^{-4} mol (2 sig fig)

3 a 1.03 kg (3 sig fig)
b Ethane is a non-polar molecule that has weak dispersion forces between each molecule, which do not require much energy to disrupt. Therefore, the boiling point is low. Ethanol is a polar molecule. Ethanol molecules are held together by stronger hydrogen bonds, which require more energy to break. Therefore, ethanol has a higher boiling point.
c i Fermentation
ii Student answers will vary but should explain that using plant-based biomass addresses United Nations Sustainable Development Goal 12 'Responsible consumption and production' and the green chemistry principle 'Use of renewable feedstocks'.

4

Name	Structural formula	Semi-structural formula
Hexane		$CH_3CH_2CH_2CH_2CH_2CH_3$
Hexanoic acid		$CH_3(CH_2)_4COOH$
1,3-Dichloropropane		$CH_2ClCH_2CH_2Cl$
Hex-1-ene		$CH_2{=}CH(CH_2)_3CH_3$
2-Methyl-4,5-dichlorohept-2-ene		$CH_3C(CH_3){=}CHCHClCHClCH_2CH_3$

5 **a** CH_3NSO
b $C_2H_6N_2S_2O_2$
c 32% (2 sig fig)

UNIT 1 review

MULTIPLE CHOICE

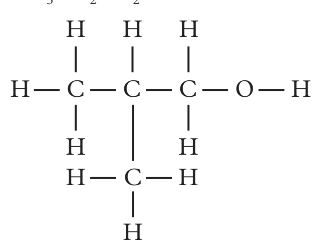
1	A	**2**	D	**3**	D	**4**	A	**5**	B
6	D	**7**	A	**8**	C	**9**	B	**10**	B

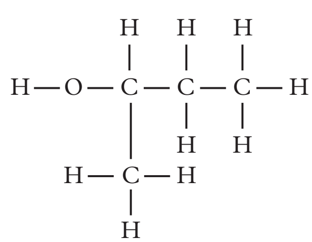
SHORT ANSWER

1 **a** $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$
b **i** Electronegativity increases
ii Atomic radius decreases
c **i** Student answers will vary but may include heat conductivity, electrical conductivity, malleability, ductility, lustre, hardness.
ii Student answers will vary but may include reactivity with oxygen, reactivity with HCl.
iii $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
iv Place two reactants in a test tube, look for bubbling and place a lit split into the opening of the test tube. If you hear a pop, this is a sign that a reaction has occurred.

2 **a** V-shaped/bent
b Bond X: hydrogen bonding (intermolecular)
Bond Y: covalent bonding (intramolecular)
c Deionised water does not contain any ions. This means there are no free charged particles to conduct electricity. However, water from the tap contains impurities that may contain dissolved free ions. These dissolved charged particles enable tap water to conduct electricity.

3 **a** Ionic bonding is one of the three types of atomic bonding that involve electrostatic attraction between positive cations and negative anions. Ionic compounds have high melting points because the electrostatic attractions between the ions are strong and takes a large amount of energy to overcome.
b Solid ionic compounds are arranged in a rigid crystal lattice where the charged ions are unable to move, and therefore unable to conduct electricity. Liquid molten ionic compounds are no longer in a rigid structure. Instead, the charged ions can move around and slide past each other. This enables molten ionic compounds to conduct electricity.
c **i** $Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3(aq)$
ii $Na^+(aq)$ and $SO_4^{2-}(aq)$
iii C yes, D no

4 **a** $CH_3CH_2CH_2CHOH$
b


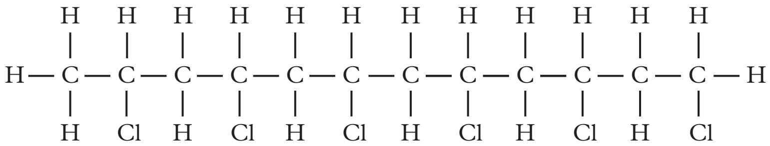


c **i** C_2H_6O

ii 8.38×10^{23} atoms (3 sig fig)

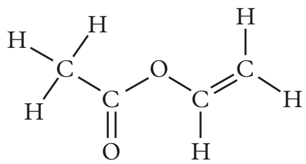
d Because of dispersion forces being closely related to molecular size, butan-1-ol will have a higher boiling point than smaller unbranched alcohols. In the same way, butan-1-ol will have a lower boiling point than larger unbranched alcohols.

5 **a**



b Thermoplastic polymers only have weak intermolecular forces between polymer chains and will soften/melt when heated. Thermosetting polymers have strong covalent bonds between polymer chains and will not soften when heated. Instead, they will burn when enough heat is being applied.

c **i**



ii $C_4H_6O_2$

d Polyethylene can be broken back down into its ethene monomers. The monomers can then be reused. This is an example of a circular economy because the polyethylene can be used and reused.

Chapter 11: Water as a unique chemical

GROUNDWORK

11A Hydrogen bond, dipole–dipole interaction, dispersion force

11B A hydrogen covalently bonded to a fluorine, oxygen or nitrogen atom

11C The presence of stronger intermolecular forces between molecules increases their melting and boiling points because it takes more energy to break apart the molecules.

11.1 The three states of water

11.1 CHECK YOUR LEARNING

1 Reservoirs that collect river water

2 It contains 3.1–3.8% sodium chloride with low concentrations of other salts.

3 Biological water is water found in the body, including fluid in body tissues. The largest volume of biological water that you will encounter is urine, which could be purified for drinking.

4 Liquid water is more accessible to us. Solid water is only accessible to us when ice melts and flows through rivers.

5 Student answers will vary, but should discuss distillation and desalination as key methods to purify water.

11.2 Anomalous properties of water

11.2 SKILL DRILL

1 Student answers will vary, but examples include: the hotplate, liquid HCl and ethanol.

2 Student answers will vary, but an example is: Only use HCl in the fume hood.

3 Student answers will vary, but examples include: avoiding wasting materials and being honest when recording data.

11.2 CHECK YOUR LEARNING

1 Water can form hydrogen bonds with itself, which require a large amount of heat energy to break, while hydrogen sulfide cannot.

2 Hydrogen telluride has more electrons available to form temporary dipoles and dispersion forces than hydrogen selenide. More heat energy is required to overcome the greater number of dispersion forces.

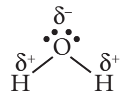
3 Melting point: the temperature at which a solid turns into a liquid, the energy required to weaken intermolecular bonds; Boiling point: the temperature at which a liquid turns into a gas, depends on the energy required to break intermolecular bonds.

4 Similarity: Water molecules in liquid water and ice form hydrogen bonds with each other.

Difference: Water molecules in liquid form only 2–3 hydrogen bonds with other water molecules. The distance between molecules is relatively short and the density is higher, meaning that liquid water occupies a relatively small volume. Water molecules in ice have less kinetic energy, allowing them to form four stable hydrogen bonds around each water molecule. The hydrogen bonds are held further apart in a three-dimensional hexagonal lattice, so the density is lower and the volume is higher.

5 **a**

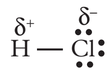
I



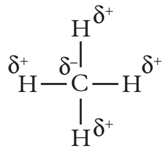
II No dipoles

H — H

III



IV



b I hydrogen bonding

II dispersion forces

III dipole–dipole attractions

IV dispersion forces

6 Highest: water; lowest: hydrogen gas

7 The specific heat capacity of sand is much lower than that of water ($0.84\,J\,g^{-1}\,K^{-1}$ compared to $4.18\,J\,g^{-1}\,K^{-1}$). Sand will increase in temperature more quickly when heat is applied.

11.3 Vaporisation of water

11.3 CHECK YOUR LEARNING

1 The energy required to change the physical state of one mole of a substance.

2 $40.7\ kJ\ mol^{-1}$

3 Water has strong hydrogen bonds between molecules, which require more energy to break than other molecules that do not form hydrogen bonds.

4 Similarity: both indicate the amount of heat energy required to change the physical state of one mole of a substance; Difference: latent heat of fusion is the heat energy required to change a solid to a liquid, whereas latent heat of vaporisation is the heat energy required to change a liquid into a gas.

5 **a** **i** solid

ii liquid

iii gas

b **i** melting of solid ethanol into liquid ethanol

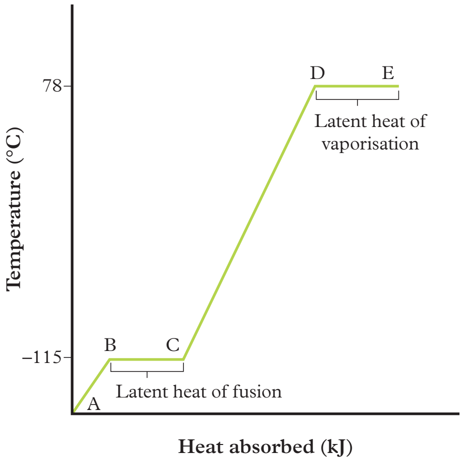
ii boiling of liquid ethanol to form ethanol vapour

c **i** Heating of solid ethanol – the ethanol molecules vibrate faster in their fixed positions and gain kinetic energy as the intermolecular forces become weaker.

ii Heating of liquid ethanol – they gain kinetic energy and the temperature increases as the intermolecular forces become weaker.

iii Heating of ethanol vapour – the ethanol molecules gain more kinetic energy and temperature continues to rise as the intermolecular forces become weaker.

d



6 Student answers will vary but should include the steps required to generate a heating curve, the materials/equipment required and specify the independent and dependent variables. Methods may be similar to Practical 11.3.

Chapter 11 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Water can form hydrogen bonds with itself, which require a large amount of heat energy to break, while hydrogen sulfide cannot.

12 Water molecules can form strong hydrogen bonds with other water molecules, which can absorb a lot of heat energy before they break.

13 Seawater contains a high concentration of sodium chloride, which can harm cells when ingested.

14 The hydrogen bonds between the H_2O molecules weaken and the molecules start to flow over each other.

15 All the remaining hydrogen bonds between H_2O molecules break, and the molecules separate.

16 Allows oceans and rivers to remain at a stable temperature because a lot of heat energy needs to be absorbed before the temperature rises.

17 Similarity: specific heat capacity and latent heat are physical properties of chemical substances. They are both related to the ability of a substance to absorb heat energy.

Difference: Specific heat capacity is the amount of energy required to raise the temperature of 1 g of a substance by 1°C or 1 K. It indicates the ability of a material to absorb heat energy and is measured in $J\,g^{-1}\,K^{-1}$. Latent heat is the energy required to change the physical state of one mole of a substance. It is measured in $kJ\,mol^{-1}$.

18 Hydrogen fluoride can form hydrogen bonds with other hydrogen fluoride molecules, which require more energy to break than the dispersion forces between elemental fluorine molecules.

19 The high heat capacity of water allows it to absorb large amounts of heat before the temperature of the water changes. Because of this, heat generated by the gaming PC can be absorbed by the water to keep the PC cool.

20 a Vegetable oil because it has the smallest heat capacity and will increase in temperature when a smaller amount of heat is applied.

b Student answers may vary, but examples include: ethanol (which is flammable) and hotplates (which can reach very high temperatures).

c Ethanol

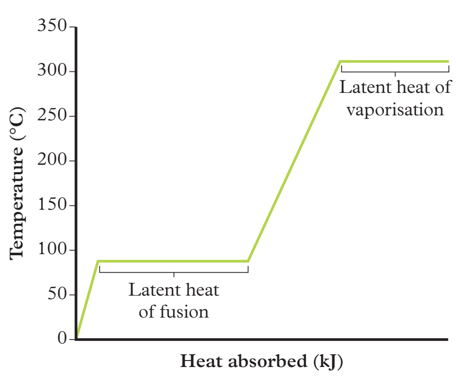
21 a Ice has a lower density than liquid water because the water molecules are able to form four hydrogen bonds. This causes them to become arranged in a lattice structure with greater distance between each water molecule. Lower density substances will float on top of higher density substances, therefore ice floats on water.

b Ice formed in lakes during winter will form a layer above the liquid water. Most of the water underneath will not freeze. This allows fish and other aquatic life to survive underneath the ice layer.

22 a i 85°C

ii 310°C

b



23 The atomic mass of xenon is greater than that of argon, allowing it to form more dispersion forces. More energy is required to break the greater amount of dispersion forces, so the latent heat of vaporisation of xenon is higher than argon.

24 Student answers will vary but should include the steps required to determine specific heat capacity, the materials/equipment required and the independent and dependent variables.

25 Student answers will vary but should include the steps required to determine the latent heat of vaporisation of a substance, the materials/equipment required and the independent and dependent variables. Student methods should include heating the substance beyond 80°C until it has all vaporised.

26 Student answers will vary but should indicate whether they agree or disagree with the statement and provide justifications.

Chapter 12: Acid-base reactions

GROUNDWORK

12A Acids are molecules that donate protons. whereas bases are molecules that accept protons.

12B The pH scale is a logarithmic scale used to compare the acidity of a sample based on its hydronium concentration.

- pH 1 to <7: acidic
- pH 7: neutral
- pH >7 to 14: basic

12C Conjugate acid–base pairs are formed by the gain or loss of protons and differ by one proton.

12.1 The Brønsted–Lowry theory of acids and bases

12.1 CHALLENGE

1 $H_2PO_4^-(aq) + NH_3(aq)\ NH_4^+(aq) + HPO_4^{2-}(aq)$

$HPO_4^{2-}(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + PO_4^{3-}(aq)$

2 $HCO_3^-(aq) + H_2S(aq) \rightarrow CO_2(g) + H_2O(l) + S^{2-}(aq)$

3 $HSO_4^-(aq) + HCO_3^-(aq) \rightarrow CO_2(g) + H_2O(l) + SO_4^{2-}(aq)$

12.1 CHECK YOUR LEARNING

1 Acids are hydrogen ion (H^+) or proton donors and bases are hydrogen ion (H^+) or proton acceptors.

2 Conjugate acid–base pairs refer to an acid and base that differ by one proton.

3 a NO_2^-

b SO_3^{2-}

4 **a** HCOOH

b $HIO_3$5 Amphiprotic species can act as an acid or base (i.e. donate or accept protons), depending on its surroundings. An example of amphiprotic species is water.

$H_2O(l) + HCl(aq) \rightarrow H_3O^+(aq) + Cl^-(aq)$ Water is acting as a base

$H_2O(l) + NH_3(aq) \rightarrow OH^-(aq) + NH_4^+(aq)$ Water is acting as an acid

6 **a** $RbOH(aq) + H_2O\ (l) \rightarrow Rb^+(aq) + OH^-(aq)$

b $H_2CO_3(aq) + H_2O(l) \rightarrow \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$

$HCO_3^-(aq) + H_2O(l) \rightarrow \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$

12.2 Strong and weak acids and bases

12.2 CHALLENGE

$c(NaOH) = 0.10\,M$ (or 40 g L^{-1}) (2 sig fig)

12.2 CHECK YOUR LEARNING

1 Student answers may vary but may include: $H_2PO_4^-$, HPO_4^{2-}, HCO_3^-

2 A strong acid will readily donate its proton(s) to a base and completely dissociate in water, e.g. $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

A weak acid will not readily donate its proton(s) to a base and only partially dissociates in water — the incomplete ionisation can be represented by the reversible arrow in its chemical equation.e.g. $HCN(aq) + H_2O(l) \rightarrow CN^-(aq) + H_3O^+(aq)$

3 A strong base will ready accept proton(s) from an acid and will completely dissociate in water.

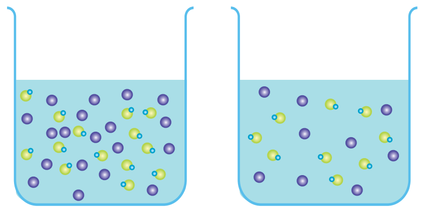
e.g. $NaOH(aq) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq)$

A weak base will not readily accept proton(s) from an acid and will only partially dissociate in water — the incomplete ionisation can be represented by the reversible arrow in its chemical equation. e.g. $NH_3(aq) + H_2O(l) \rightleftharpoons \rightarrow NH_4^+(aq) + OH^-(aq)$

4 Concentration is a quantitative measure of the amount of particles within a given volume. Something that is concentrated has a large amount of particles in a given volume while something that is diluted has only a small amount of particles in a given volume.

5 The strength of an acid or a base is a measure of how readily a substance donates or accepts a proton and how readily it dissociates in water. On the other hand, concentration is a quantitative measure of the number of acid or base particles in a given volume.

6 **a**



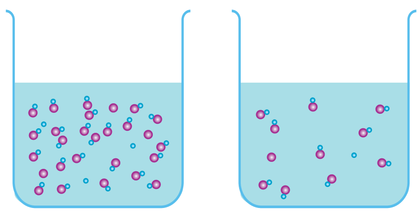
Concentrated strong base   Dilute strong base

Sodium hydroxide NaOH (strong base)

Conjugate acid H_2O

Na^+ OH^- H^+

b



Concentrated weak acid   Dilute weak acid

Ammonia NH_3 (weak base)

Conjugate acid NH_4^+

H^+

12.3 Neutralisation reactions to produce salts

12.3 SKILL DRILL

Student answers will vary.

12.3 CHECK YOUR LEARNING

1 **a** Water + metal salt

b Water+ metal salt + carbon dioxide

2 • Collect gas and bubble through limewater — the presence of a milky white precipitate indicates carbon dioxide.

• Hold a lit match next to the mouth of the reaction vessel; the match will extinguish in the presence of carbon dioxide

3 Both reactions will produce water and metal salts. However, acid reactions with metal carbonates also produce carbon dioxide but acid reactions with metal hydroxides do not.

4 **a** $2H_3PO_4(aq) + 3Ca(OH)_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6H_2O(l)$

$H^+(aq)+ OH^-(aq) \rightarrow H_2O(l)$

b $MgCO_3(s) + 2CH_3COOH(aq) \rightarrow (CH_3COO)_2Mg(aq) + CO_2(aq) + H_2O(l)$

$MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(s) + CO_2(aq) + H_2O(l)$

c $2LiHCO_3(aq) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(l) + 2CO_2(g)$

$HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

d $NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq) \rightarrow NH_4Cl(aq)$ (molecular equation)

$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$ (ionic equation)

e $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

5 $Mg_2O_8Si_3 \cdot H_2O(aq) + 4HCl(aq) \rightarrow 2MgCl_2(aq) + H_4O_8Si_3 \cdot H_2O(aq)$

12.4 The pH scale

12.4 CHALLENGE

1 pH = 7.4 (2 sig fig)

2 The answer calculated (pH =7.4) suggests that the blood sample is closer to neutral and less basic than what the diagram indicated.

12.4 CHECK YOUR LEARNING

1 Water is an amphiprotic substance that can donate and accept protons. Sometimes a small number of water molecules will dissociate and transfer protons between each other, which results in hydronium (H_3O^+) and hydroxide (OH^-) ions.

$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

2 **a** Two of: hydrochloric acid, stomach acid, lemon, vinegar, apple, tomato, banana.

b Two of: caustic soda, ammonia, soap, small intestine, baking soda.

c Two of: milk, distilled water, blood.

3 The pH scale is a logarithmic scale measuring a sample's acidity by its hydronium concentration. The pOH scale is also a logarithmic scale, but it measures a sample's alkalinity (how basic) by its hydroxide concentration. Both use the logarithmic scale and add up to 14.

4 pH = 11

5 $[H^+] = 7.59 \times 10^{-4}$ M (3 sig fig)

6 pH = 13.6 (3 sig fig)

pOH = 0.40 (3 sig fig)

7 pH = 6.52 (3 sig fig)

pOH = 7.48 (3 sig fig)

12.5 Indicators and measuring pH

12.5 REAL-WORLD CHEMISTRY

1 The pH of wine is below 7; therefore, wine is acidic.

2 Wine production involves the fermentation of grape juice by selective microbes that live between the pH

range of 3.5–4.0. If the pH level of wine during production deviates from this range, then other unwanted microbes can survive and ruin the wine.

3 Litmus paper and universal indicator only give a rough estimate of pH and should not be used in the wine making process, which requires more precise pH control.

12.5 CHECK YOUR LEARNING

1 Red litmus paper turned blue, which indicates that this solution is basic.

2 The chemical compound anthocyanin gives a coloured pigment that is sensitive to pH. When the pH is low, the pigment is stable and appears red; when the pH is high, the pigment fades into a blue colour.

3 Natural indicators and litmus paper can only provide a rough guidance on whether a sample is acidic or basic and is the most limited in terms of accuracy and precision. Universal indicators can be used to determine acid and base strength based on colour change and are more accurate and precise than litmus paper, especially when the pH of a substance is near neutral. The pH meter is the only method mentioned so far that gives quantitative results; it is far more accurate and precise than the qualitative methods described.

4 **a** Methyl violet, thymol blue
b Thymol blue, bromothymol blue, phenolphthalein
c Alizarin yellow

5 Students answers may vary but should consider all controlled variables and how to measure dependent variables in a consistent manner.

12.6 Acid-base reactions in society

12.6 SKILL DRILL

1 The measured levels of atmospheric CO_2 has steadily increased over time in the last five decades. In 1960, the level of CO_2 in the atmosphere was under 325 ppmv and by the mid-2000s it had risen to levels above 375 ppmv.

2 The measured levels of carbon dioxide in the seawater have steadily increased over the last two decades.
In 1990, the level of CO_2 in the seawater was near 325 µatm and by the mid-2000s it had risen to levels above 350 µatm.

3 As the level of atmospheric carbon dioxide increases, so does the dissolved carbon dioxide in the seawater. This is due to the increased availability of atmospheric carbon dioxide to be dissolved into the water. When carbon dioxide is dissolved in seawater, it reacts with water to form carbonic acid, which has decreased the pH of the seawater over time, making it more acidic.
$CO_2(aq) + H_2O\ (l) \rightleftharpoons H_2CO_3(aq)$
formation of carbonic acid

12.6 CHECK YOUR LEARNING

1 Ocean acidification is the decrease of ocean pH level, increasing its acidity.

2 Shell-building animals and corals build their shells from free carbonate and calcium ions available in seawater ($Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$). Increased ocean acidity means increased carbonic acid concentration in the water, which limits the formation of calcium carbonate. This is because free carbonate ions tend to react and interact with free hydrogen ions (supplied by the carbonic acid) rather than form calcium carbonate.

3 Student responses may vary but can include combinations of the following:
- Carbon dioxide is added to the carbon–oxygen cycle by: volcano eruptions, burning of fossil fuels, released from oil, respiration, decomposition of dead animals.
- Carbon dioxide is removed from the carbon–oxygen cycle by: photosynthesis, formation carbonates in the sea, formation of acid rain, formation of calcium carbonate from coral and marine animals.

4 Normal rain is slightly acidic with a pH of approximately 5.6 from the atmospheric carbon dioxide reacting with water vapour to form carbonic acid (weak acid).
Acid rain, contains high levels of either nitric or sulfuric acid with a pH usually between 4.2 and 4.4. Acid rain is caused by the additional atmospheric sulfur dioxide or nitric oxide (usually released from burning of fossil fuels) in the air reacting with water vapours to form sulfuric acid, nitrous acid and nitric acid.

5 Student answers will vary but should discuss Goal 13 'Climate action' and Goal 14 'Life below water' because they are most relevant to the problem of ocean acidification.
Other goals that may contribute to discussion are: Goal 6 'Clean water and sanitation', Goal 7 'Affordable and clean energy', Goal 11 'Sustainable cities and communities', Goal 12 'Responsible consumption and production'.

Chapter 12 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 **a** $H_2CrO_4(aq) + H_2O(l) \rightarrow HCrO_4^-(aq) + H_3O^+(aq)$
$HCrO_4^-(aq) + H_2O(l) \rightarrow CrO_4^{2-}(aq) + H_3O^+(aq)$
b $PH_3(aq) + H_2O(l) \rightarrow PH_4^+(aq) + OH^-(aq)$
c $HS^-(aq) + H_2O(l) \rightarrow H_2S(aq) + OH^-(aq)$
$HS^-(aq) + H_2O(l) \rightarrow S^-(aq) + H_3O^+(aq)$

12 **a** $HBr(aq) + H_2O(l) \rightarrow Br^-(aq) + H_3O^+(aq)$
$HOI(aq) + H_2O(l) \rightarrow OI^-(aq) + H_3O^+(aq)$
b $HBr(aq) + H_2O(l) \rightarrow Br^-(aq) + H_3O^+(aq)$
- Acid: $HBr(aq)$
- Base: $H_2O(l)$
- Conjugate acid: $H_3O^+(aq)$
- Conjugate base: $Br^-(aq)$

$HOI(aq) + H_2O(l) \rightleftharpoons \rightarrow OI^-(aq) + H_3O^+(aq)$
- Acid: $HOI(aq)$
- Base: $H_2O(l)$
- Conjugate acid: $H_3O^+(aq)$
- Conjugate base: $IO^-(aq)$

c In the $HBr(aq) + H_2O(l) \rightarrow Br^-(aq) + H_3O^+(aq)$ reaction, the strong base (HBr) completely underwent complete 'dissociation' as it broke apart into smaller ions. The water itself wasn't reacting but rather facilitating the dissociation. In the $HOI(aq) + H_2O(l) \rightarrow OI^-(aq) + H_3O^+(aq)$ reaction, the weak acid HOI underwent partial 'ionisation' in water — this involved the dissociation of the ions before new ions are formed.

13 **a** **i** $[H^+] = 0.000\,10$ M (2 sig fig)
ii $[OH^-] = 1.0 \times 10^{-10}$ M (2 sig fig)
b $HCOOH(aq) + H_2O(l) \rightarrow \rightleftharpoons HCOO^-(aq) + H_3O^+(aq)$

14 This is a base as the pH is greater than 7.
$[H_3O^+] = 8.91 \times 10^{-9}$ M (3 sig fig)
$[OH^-] = 1.11 \times 10^{-6}$ M (3 sig fig)

15 $[H_3O^+] = 1.51 \times 10^{-5}$ M (3 sig fig)
$[OH^-] = 6.61 \times 10^{-10}$ M (3 sig fig)

16 $[H_3O^+] = 0.0040$ M (2 sig fig)

17 **a** 2 M (1 sig fig) **b** 5×10^{-15} M (1 sig fig)
c 14

18 **a** Acid: H_2O, Base: NH_3, Conjugate base: OH^-, Conjugate acid: NH_4^+
b Acid: H_2SO_4, Base: H_2O, Conjugate base: HSO_4^-, Conjugate acid: H_3O^+
c Acid: CH_3COOH, Base: H_2O, Conjugate base: CH_3COO^-, Conjugate acid: H_3O^+

19 **a** $MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$
$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

b $LiOH(aq) + CH_3COOH(aq) \rightarrow CH_3COOLi(aq) + H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

c $2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$
$2H^+(aq) \rightarrow H_2(g)$

20 Solution A: $[H_3O^+] = 3.0$ M, $[OH^-] = 3.3 \times 10^{-15}$ M (2 sig fig)
Solution B: $[H_3O^+] = 0.0024$ M, $[OH^-] = 4.2 \times 10^{-12}$ M (2 sig fig)

21 Student answers will vary but experimental design should consider all controlled variables and how to measure dependent variables in a consistent manner.

Chapter 13: Redox reactions

GROUNDWORK

13A First ionisation energy and therefore reactivity increases going down and from right to left across the periodic table.

13B Ionic bond

13C $Na_2CO_3(aq) + Ca(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + CaCO_3(s)$

13.1 Oxidising and reducing agents

13.1 CHALLENGE

1 $CH_3OH(aq) + H_2O(aq) \rightarrow HCOOH(aq) + 4e^- + 4H^+(aq)$

2 $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 7H_2O(l) + 2Cr^{3+}(aq)$

3 $3CH_3OH(aq) + 2Cr_2O_7^{2-}(aq) + 16H^+(aq) \rightarrow 3HCOOH(aq) + 11H_2O(l) + 4Cr^{3+}(aq)$

13.1 CHECK YOUR LEARNING

1 Reduction is the gain of electrons and oxidation is the loss of electrons. These reactions occur simultaneously.

2 Valence shell electrons are transferred. The oxidising agent gains electrons and the reducing agent loses electrons.

3 Oxidation half-equation: electrons on the product side of the equation; Reduction half-equation: electrons on the reactant side of the equation.

4 **a** I lithium is oxidised, sulfur is reduced
II magnesium is oxidised, fluorine is reduced
III manganese is oxidised, silver is reduced

b I sulfur is the oxidising agent, lithium is the reducing agent
II fluorine is the oxidising agent, magnesium is the reducing agent
III silver is the oxidising agent, manganese is the reducing agent

c I oxidation half-equation: $Li(s) \rightarrow Li^+(aq) + e^-$
reduction half-equation: $S(g) + 2e^- \rightarrow S^{2-}(aq)$
II oxidation half-equation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$
reduction half-equation: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$
III oxidation half-equation: $Mn(s) \rightarrow Mn^{2+}(aq) + 2e^-$
reduction half-equation: $Ag^+(aq) + e^- \rightarrow Ag(s)$

d I $2Li(s) + S(g) \rightarrow Li_2S(s)$
II $Mg(s) + F_2(g) \rightarrow MgF_2(s)$
III $Mn(s) + 2AgNO_3(aq) \rightarrow Mn(NO_3)_2(aq) + 2Ag(s)$

5 **a** The oxidised species is H_2 and the reduced species is Cl_2.
b The oxidised species is Fe and the reduced species is Cu^{2+}.
c The oxidised species is Pb and the reduced species is Ag^+.

6

6	a	oxidation:	$H_2(g) \rightarrow 2H^+(aq) + 2e^-$
		reduction:	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
	b	oxidation:	$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
		reduction:	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
	c	oxidation:	$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$
		reduction:	$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

7 **a** $MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + 2H_2O(l)$
The oxidation number has decreased, so reduction has occurred.

b $BiO_3^-(aq) + 6H^+(aq) + 2e^- \rightarrow Bi^{3+}(aq) + 3H_2O(l)$
The oxidation number has decreased, so reduction has occurred.

c $SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$
The oxidation number has increased, so oxidation has occurred.

d $NO_3^-(aq) + 3H^+(aq) + 2e^- \rightarrow HNO_2(aq) + H_2O(l)$
The oxidation number has decreased, so reduction has occurred.

8 Student answers will vary but should include reactions that demonstrate a gain of oxygen or a loss of hydrogen.

13.2 The reactivity series of metals

13.2 REAL-WORLD CHEMISTRY

1 $6CO_2 + 24H^+ + 24e^- \rightarrow C_6H_{12}O_6 + 6H_2O$

2 $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

3 Glucose is the reducing agent. Oxygen is the oxidising agent.

13.2 SKILL DRILL

1 Aim: To determine the reactivity of metals by measuring the volume of hydrogen gas produced when they are reacted with HCl
Hypothesis: Calcium will be more reactive than magnesium or aluminium because it has a higher period number (further down in group 2) than magnesium and a lower group number than aluminium.

2 IV: metal being tested. DV: volume of hydrogen gas

3 Materials: 1 M HCl, calcium, magnesium and aluminium strips of a similar mass and surface area.
Equipment: beakers, sensitive balance (mass), conical flask, measuring cylinder (or graduated/volumetric pipette) for measuring acid, timer, gas syringe to collect gas

4 Quantitative because volume is recorded

13.2 CHECK YOUR LEARNING

1 For metals, the reactivity decreases as you go across a period. For non-metals, the reactivity increases going across a period.

2 For metals, the reactivity increases as you go down a group. For non-metals, the reactivity decreases going down a group

3 Metals have fewer valence electrons and are more likely to lose them in oxidation than gain more. Non-metal elements undergo reduction and gain electrons to complete their valence shell.

4 **a** No reaction will occur.
b A reaction will occur to form zinc nitrate and solid nickel.
$Ni(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Ni(s)$
c A reaction will occur to form iron(II) nitrate and solid copper.
$Cu(NO_3)_2(aq) + Fe(s) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$
d No reaction will occur.

5 **a** Lithium is a stronger reducing agent than tin. Lithium is higher in the metal reactivity series and it is in Group 1 in the periodic table.
b Zinc is a stronger reducing agent than iron. Zinc is higher in the metal reactivity series.
c Manganese is a stronger reducing agent than chromium. Manganese is higher in the metal reactivity series.

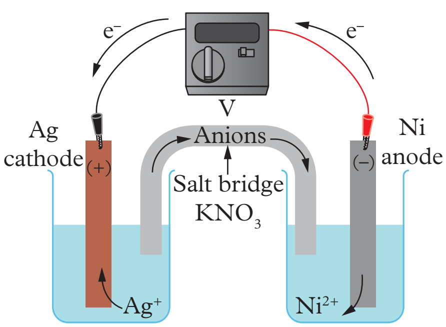
13.3 Redox reactions in society

13.3 REAL-WORLD CHEMISTRY

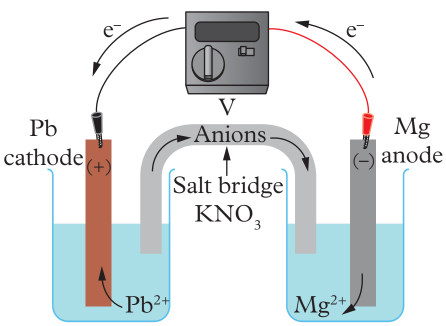
1 $4Fe(s) + 3O_2(aq) \rightarrow 2Fe_2O_3(s)$
2 Paint forms a barrier over the iron structure to protect it from corrosion.
3 To protect the iron in the structure from being exposed, therefore preventing its corrosion

13.3 CHECK YOUR LEARNING

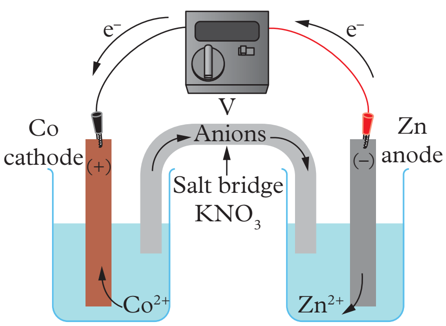
1 Corrosion is the degradation of a metal to form a more stable compound. This involves a loss of electrons, so it is an oxidation reaction.
2 When electrons are transferred between reactants, an electric current can be generated.
3 Zinc is more reactive than iron. It can be oxidised and cause iron to be reduced, instead of oxidised (and therefore corroded).
$Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation of zinc)
$Fe^{2+} + 2e^- \rightarrow Fe$ (reduction of iron)
4 **a**



b



c



5

	Oxidation half-equation	Reduction half-equation	Overall equation
a	$Ni \rightarrow Ni^{2+} + 2e^-$	$Ag^+ + e^- \rightarrow Ag$	$Ni + 2Ag^+ \rightarrow 2Ag + Ni^{2+}$
b	$Mg \rightarrow Mg^{2+} + 2e^-$	$Pb^{2+} + 2e^- \rightarrow Pb$	$Mg + Pb^{2+} \rightarrow Pb + Mg^{2+}$
c	$Zn \rightarrow Zn^{2+} + 2e^-$	$Co^{2+} + 2e^- \rightarrow Co$	$Zn + Co^{2+} \rightarrow Co + Zn^{2+}$

6 Student answers will vary, but may include that corrosion of magnesium can provide electrons to iron to prevent iron from corrosion (sacrificial anodes).
7 **a** Primary cells do not align with the green chemistry principle of preventing waste because they are single-use and must be disposed of after use. They therefore generate considerable waste.
b Goal 7: Affordable and clean energy
c Student answers will vary but should indicate that secondary cells are rechargeable batteries. They help address the green chemistry principle of preventing wastes, since they can be reused.

Chapter 13 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Atoms of elements on the left of the periodic table (the metal atoms) are more likely to lose electrons. Atoms on the right-hand side of the periodic table (the non-metals) are likely to gain electrons.
12 An oxidising agent causes oxidation and is reduced in the process, whereas a reducing agent causes reduction and is oxidised. One is the opposite of the other.
13 Oxidation is the loss of electrons whereas reduction is the gain of electrons. You cannot have one process without the other occurring.
14 The reactivity series of metals is a list of metals arranged from most reactive to least reactive. The stronger reducing agents are at the top, the stronger oxidising agents are at the bottom.
15 **a** A displacement reaction could occur. The less reactive species is reduced as it accepts electrons to become a solid metal. The more reactive species loses electrons (is oxidised) to become an ion.
b This is a galvanic cell. The chemical energy is converted to electrical energy from the flow of electrons. Both metal species are consumed in the reaction.
16 **a** +7
b +6
c +4
d –3
17 **a** I sodium metal undergoes oxidation, bromine gas undergoes reduction
II aluminium metal undergoes oxidation, hydrogen from HCl undergoes reduction
III lithium metal undergoes oxidation, hydrogen from water undergoes reduction
b I sodium is the reducing agent, bromine is the oxidising agent
II aluminium is the reducing agent, hydrogen from HCl is the oxidising agent
III lithium is the reducing agent, hydrogen from water is the oxidising agent
c I oxidation: $Na \rightarrow Na^+ + e^-$
reduction: $Br_2 + 2e^- \rightarrow 2Br^-$
II oxidation: $Al \rightarrow Al^{3+} + 3e^-$
reduction: $2H^+ + 2e^- \rightarrow H_2$
III oxidation: $Li \rightarrow Li^+ + e^-$
reduction: $2H^+ + 2e^- \rightarrow H_2$
d I $2Na(s) + Br_2(g) \rightarrow 2NaBr(s)$
II $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
III $2Li(s) + 2H_2O\ (l) \rightarrow 2LiOH(aq) + H_2(g)$
18 **a** NO is the reducing agent, NO_2^- is the oxidising agent.
b SO_4^{2-} is the oxidising agent, SO_2 is the reducing agent.
c TeO_3^{2-} is the oxidising agent, Te is the reducing agent.
d IO^- is the reducing agent, IO_3^- is the oxidising agent.
e As is the reducing agent, $H_sAsO_4^-$ is the oxidising agent.
f ReO_4^- is the oxidising agent, Re is the reducing agent.
19 **a** Ag+ is reduced and Mn is oxidised
b Cl_2 is reduced and K is oxidised
c Br_2 is reduced and Ca is oxidised
d Hg^{+2} is reduced and Fe is oxidised
e O_2 is reduced and C is oxidised
20 **a** $2MnO_4^-(aq) + 5SO_2(g) + 2H_2O(l) + H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq)$
b $2NO_3^-(aq) + 2H^+ + H_2S(aq) \rightarrow S(s) + 2NO_2 + 2H_2O$
c $3PbS(s) + 2NO_3^-(aq) + 8H+(aq) \rightarrow 3S(s) + 3Pb^{2+}(aq) + 2NO(g) + H_2O(l)$
d $6I^-(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 3I_2(s) + 2Cr^{3+}(aq) + 7H_2O(l)$
e $Sn^{2+}(aq) + H_3AsO_4(aq) + 2H^+(aq) \rightarrow Sn^{4+}(aq) + H_3AsO_3(aq) + H_2O(l)$

21 **a** Yes, a reaction will occur.
Chlorine is a strong oxidising agent and magnesium is a strong reducing agent.

b Yes, a reaction will occur. Hydrogen is a strong oxidising agent and zinc is a strong reducing agent.

c No, a reaction will not occur. The sulfate ion is not a stronger reducing agent than bromine.

d Yes, a reaction will occur. Iron(III) ion is a strong oxidising agent and hydrogen sulfide is a strong reducing agent.

22 **a**

	Oxidation	Reduction
I	$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	$Br_2(g) + 2e^- \rightarrow 2Br^-(aq)$
II	$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$	$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$
III	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$	$2H^+(aq) + 2e^- \rightarrow H_2(g)$

b

I	$H_2(g) + Br_2(g) \rightarrow 2HBr(aq)$
II	$2Al(s) + 3Fe(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3Fe(s)$
III	$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

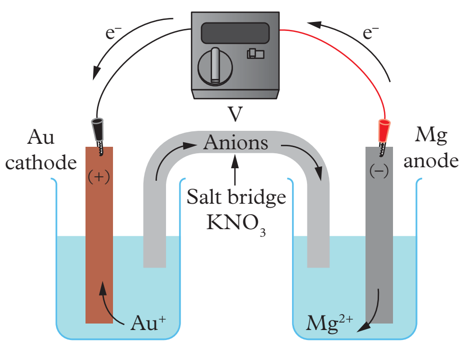
c

	Oxidation	Reduction
I	$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	$Br_2(g) + 2e^- \rightarrow 2Br^-(aq)$
II	$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$	$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$
III	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$	$2H^+(aq) + 2e^- \rightarrow H_2(g)$

d

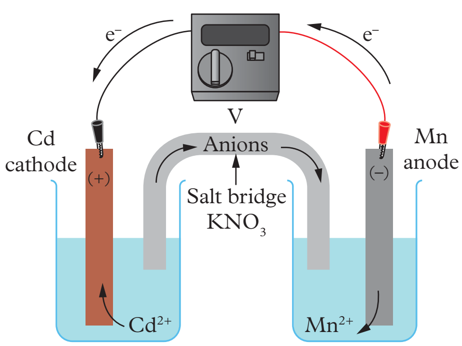
	Reducing agent	Oxidising agent
I	$H_2(g)$	$Br_2(g)$
II	$Al(s)$	$Fe^{2+}(aq)$
III	$Mg(s)$	$2H^+(aq)$

23 **a**



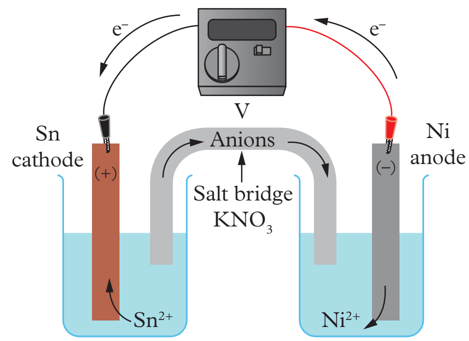
Oxidation half-equation	$Mg \rightarrow Mg^{2+} + 2e^-$
Reduction half-equation	$Au^+ + e^- \rightarrow Au$
Overall equation	$2Au^+ + Mg \rightarrow 2Au + Mg^{2+}$

b



Oxidation half-equation	$Mn \rightarrow Mn^{2+} + 2e^-$
Reduction half-equation	$Cd^{2+} + 2e^- \rightarrow Cd$
Overall equation	$Mn + Cd^{2+} \rightarrow Mn^{2+} + Cd$

c



Oxidation half-equation	$Ni \rightarrow Ni^{2+} + 2e^-$
Reduction half-equation	$Sn^{2+} + 2e^- \rightarrow Sn$
Overall equation	$Ni + Sn^{2+} \rightarrow Ni^{2+} + Sn$

24 Student answers will vary.

25 **a** Cu^{2+} and Au^+ are strong oxidising agents (low on the reactivity series of metals). They will cause Zn and Fe to undergo oxidation into ions, while they themselves are reduced to form solid metals.

b K and Li are both very strong reducing agents. H^+ ions are much stronger oxidising agents than Al^{3+} ions and are more likely to undergo reduction than Al^{3+}. Therefore, H_2 will be formed and not solid Al.

26 Student answers will vary but could suggest measuring the change in mass of each electrode. Based on the reactivity series of metals, the expected result is, from most to least reactive: Mg, Cd, Sn.

UNIT 2 Area of Study 1 Checkpoint

MULTIPLE CHOICE

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

SHORT ANSWER

1 **a** $HIn(aq) + H_2O(l) \rightarrow In^-(aq) + H_3O^+(aq)$

b KNO_3: neutral; K_3PO_4: base; NH_4NO_3: acidic.

c $K_2S(s) + H_2O(l) \rightleftharpoons 2K^+(aq) + S^{2-}(aq)$

d $S^{2-}(aq) + H_2O(l) \rightleftharpoons HS^-(aq) + OH^-(aq)$; alkaline.

2 a $2CH_2O + H_2SO_4 \rightarrow H_2S(g) + 2CO_2(g) + 2H_2O\ (l)$

b H_2SO_4

3 a Carbonate ions are more basic in solution and more readily accept protons to form hydroxide ions than ammonia. Therefore, sodium carbonate has a higher pH than ammonia.

b 1×10^{-5} M

c $CO_3^{2-}(aq)$

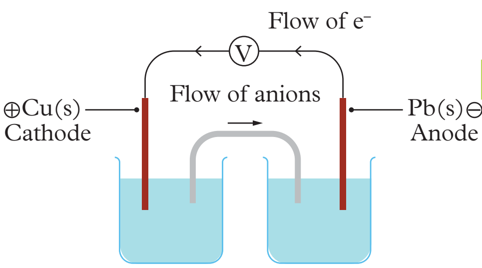
4 a Full equation: $2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$

Ionic equation: $2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$

b Full equation: $Na_2CO_3(aq) + H_2CrO_4(aq) \rightarrow Na_2CrO_4(aq) + CO_2(g) + H_2O(l)$

Ionic equation: $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

5 a



i Anode is negative, cathode is positive

ii Copper is the cathode, lead is the anode

iii Electrons flow from the anode to the cathode

iv Anions in the salt bridge flow from the cathode to the anode

b i $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$

ii $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

c $Pb(s) + Cu^{2+}(aq) \rightarrow Pb^{2+}(aq) + Cu(s)$

6 a The correct answer is 3.

b Iron should react with Pb because Fe is above Pb is the reactivity series.

c i $Cu(NO_3)_2(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$

ii Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

7 a i $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

ii Carbonic acid reacts with water to produce free $H^+(aq)$ ions, lowering the pH of the ocean: $H_2CO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HCO_3^-(aq)$

iii $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)$

iv Shell dissolves: $CaCO_3(s) + H_3O^+(aq) \rightleftharpoons Ca^{2+}(aq) + HCO_3^-(aq) + H_2O$

Free carbonate ions bond with hydrogen ions. Shell-building animals cannot extract carbonate ions from $HCO_3^-(aq)$.

b i Formation of sulfuric acid in acid rain

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

$H_2SO_3(aq) + H_2O(l) \rightarrow 2H_2SO_4(aq)$

Formation of nitric acid in acid rain

$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$

ii Student answers will vary, but could include harm to the health of humans, plants and animals, and erosion of buildings and statues.

Chapter 14: Measuring solubility and concentration

GROUNDWORK

14A $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaCl(aq)$

14B 0.032 (2 sig fig)

14C Polar molecules interact strongly with and can therefore be dissolved in other polar molecules because they can form dipole–dipole attractions — this is *like dissolves like*. Meanwhile, polar molecules cannot form dipole–dipole attractions with non-polar molecules (because these have no dipoles).

14D When the intermolecular bonds are very weak, the particles are not attracted to each other very strongly. Therefore, they remain in the gaseous state. If the intermolecular bonds are stronger, such as between polar molecules, the molecules tend to form a condensed state such as a liquid or solid.

14.1 Solution concentration

14.1 CHECK YOUR LEARNING

1 The units g L^{-1} and ppm are useful in a laboratory setting when an electronic balance can be used to measure out mass. The unit mol L^{-1} is more suited to analytical laboratory work or when the stoichiometry of a reaction is discussed. This could be more useful when working with a known concentration of solute.

2 Both the solute and solution are liquids, so it is appropriate that the percentage ratio be expressed in terms of volumes. The unit mol L^{-1} is not a convenient unit for drinks.

3 a 6 M (1 sig fig)

b 0.3 M (1 sig fig)

c 0.0655 M (3 sig fig)

4 a 620 ppm (2 sig fig)

b 930 ppm (2 sig fig)

c 0.00469 ppm (3 sig fig)

5 a 10 mL (2 sig fig)

b 2 g (1 sig fig)

c 0.1 L (1 sig fig)

d 28 kg (2 sig fig)

e 12 M (2 sig fig)

6 a 16000 ppm (2 sig fig)

b 0.025 g L^{-1} (2 sig fig)

c 0.84 M (2 sig fig)

d 5.8×10^{-5}% m/v (2 sig fig)

14.2 Solubility tables and graphs

14.2 SKILL DRILL

1 Student answers will vary but should include temperature (°C) on the *x*-axis, solubility (g L^{-1} of water) on the *y*-axis, and three data points.

2 As temperature increases, solubility of carbon dioxide decreases. As the solutions of dissolved gas and water are heated, the kinetic energy of all molecules increases. The increased motion of the particles breaks the intermolecular forces between gas and water. Thus, its solubility decreases with an increase in temperature.

3 The student may have forgotten to tare the balance between measurements. This would make their masses inaccurate.

4 Student answers will vary but may refer to theoretical solubility values of carbon dioxide for accuracy. Students may indicate that precision cannot be determined because no repeats were performed. For repeatability, students may evaluate the method used.

14.2 CHALLENGE

As gases are heated, their kinetic energy increases, breaking the intermolecular forces between gas and water. The solubility of ionic compounds increases as temperature increases as the heat breaks apart the intermolecular forces holding together the ions, allowing them to be surrounded by and solubilised by the water molecules.

14.2 CHECK YOUR LEARNING

1 As temperature increases, there is more energy available to break apart solute molecules held together by intermolecular attractions.

2 Remember *like dissolves like*. Non-polar molecules are likely to be insoluble in polar water.

3 Ionic substance > polar molecular substance > non-polar molecular substance

4 a Soluble, therefore $NaNO_3(aq)$
 b Insoluble, therefore $Pb(OH)_2(s)$
 c Insoluble, therefore $Hg_2Cl_2(s)$
 d Soluble, therefore KCl(aq)
 e Insoluble, therefore $BaSO_4(s)$

5 Similarity: solubility tables and solubility graphs both provide information about the solubility of substances in water. Differences: a solubility table provides information to determine whether a salt is soluble or insoluble (qualitative), whereas a solubility graph shows the relationship between temperature and the solubility of substances (quantitative).

6 a 62 g/100 g water
 b 116 g/100 g water
 c 63°C
 d 32 g (2 sig fig)
 e 3.4 g (2 sig fig)

7 Student answers will vary but should describe the steps to make a saturated solution of NH_4Cl, including the cooling of the solution. Students may also indicate the minimum mass of NH_4Cl that needs to be added to water.

8 Student answers will vary.

9 Student answers will vary.

14.3 Precipitation reactions

14.3 REAL-WORLD CHEMISTRY

1 $Hg^{2+}(aq) + 2I^-(aq) \rightarrow HgI_2(s)$

2 Mercury pollutants came from the side reaction of ethanal and mercury sulfate, which produced methyl mercury chloride (CH_3HgCl). CH_3HgCl was dumped into the Minamata Bay.

3 Designing safer chemicals, prevention of wastes

4 By designing safer chemicals and preventing wastes, we address Goal 6: Clean water and sanitation (improve waste disposal processes), Goal 9: Industry, innovation and infrastructure (support industry while also making sure their practices are sustainable and not generating chemicals that are harmful to humans or the environment), Goal 11: Sustainable cities and communities (making sure humans live in an environment that is safe) and Goal 14: Life below water (protect marine life by changing how waste is disposed of).

14.3 CHECK YOUR LEARNING

1 The cation–anion pairs of the two species reacting together will swap in a precipitation reaction to form a solid product that is not soluble in water.

2 Unwanted salt impurities can be precipitated from water by adding a chemical that will react with the impurity to form a precipitate. The solid precipitate can then be removed from the water by filtration.

3 a Yes. $PbCl_2$ is insoluble.
 b No. Neither of the products of the reaction are insoluble.
 c No. Neither of the products of the reaction is insoluble.

4 $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

5 a A precipitation reaction is qualitative as it shows the particles reacting but not the amounts in moles or grams.
 b You can measure the mass of the precipitate in relation to the concentration and volume of reactants.
 c Student answers will vary.

6 Sometimes not all ions react. They remain in aqueous state as reactants and products. This makes them 'spectator ions'.

Chapter 14 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Water is a polar solvent and will dissolve other polar molecules or solutes because it can form strong solute–solvent intermolecular bonds. Non-polar solutes will only form weak dispersion forces with water, so they will have low solubility. They won't mix with water and will therefore form layers.

12 Swapping of anion–cation pairs can produce an ionic substance that has low solubility in water. Sometimes, the solubility is so low that the product formed is a solid.

13 If the same units are used, it is easier to compare the solubility of different substances.

14 Student answers will vary, but should refer to the different intermolecular forces between molecules and how this affects the solubility.

15 As the solutions of dissolved gas and water are heated, the kinetic energy of all molecules increases. The increased motion of the particles breaks the intermolecular forces between gas and water. Thus, its solubility decreases with an increase in temperature.

16 It can dissolve more substances than any other solvent on Earth and can form many different solutions.

17 a 20 mL (1 sig fig)
 b 0.08 M (1 sig fig)
 c 13 g (2 sig fig)

18 a 16 g L^{-1} (2 sig fig)
 b 8.4×10^{-5} mol L^{-1} (2 sig fig)
 c 0.1 g (1 sig fig)

19 a 17 ppm (2 sig fig)
 b 0.22 mg (2 sig fig)
 c 2 L (1 sig fig)

20 At 20°C, sucrose has a solubility of 205 g/100 g water, whereas glucose has a solubility of 90 g/100 g water at the same temperature. Sucrose has 8 O–H groups, and glucose has 4 O–H groups. Because sucrose has more O–H groups that can hydrogen bond, it has a higher solubility than glucose.

21 a $Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + BaSO_4(s)$
 b $Pb(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$
 c $AlCl_3(aq) + 3KOH(aq) \rightarrow Al(OH)_3(s) + 3KCl(aq)$

22 a 190 g/100 g water (2 sig fig)
 b 20 g/100 g water (2 sig fig)
 c 32°C (2 sig fig)
 d 86 g (2 sig fig)
 e 10 g (2 sig fig)
 f Student answers will vary.

23 Student answers will vary.

24 Student answers will vary.

25 Student answers will vary, but should identify the unwanted ions present in wastewater, identify suitable ions to be added for precipitation and explain the choice using relevant solubility rules, state the method for achieving the precipitation and include a balanced precipitation reaction.

Chapter 15: Analysis for acids and bases

GROUNDWORK

15A A quantity that describes the amount of a substance. 1 mole is 6.02×10^{23} particles.

15B The concentration of a substance is the number of moles of solute in a specific volume of solvent. It can be calculated as: $c = n/V$

15C Acids, which form solutions with a low pH, will react with bases, which form solutions with a high pH, to form neutral solutions of salt and water with a pH of 7.

15.1 CHECK YOUR LEARNING

1 Volume–volume stoichiometry is used to calculate the concentration of an unknown solution (with a known volume) by reacting it with a known volume and concentration of a second solution.

2 Step 1. Calculate the number of moles of the 'known' or standard solution.
Step 2. Calculate the number of moles

of the unknown solution, using a ratio statement.

Step 3. Calculate the volume or concentration of the unknown solution depending on what the question is asking you to do.

3 A dilution involves adding extra solvent, in most cases water, to a solution.

4 KOH

5 a $H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2\ H_2O(l)$

b 1.6 mL (2 sig fig)

6 0.480 M (3 sig fig)

7 8.3 mL (2 sig fig)

8 a 4 b 16.67 c 10

9 a 0.10 M (2 sig fig)

b 0.090 M (2 sig fig)

c 2.0 M (2 sig fig)

10 600 mL (2 sig fig)

11 a HCl

b 0.21g (2 sig fig)

15.2 Acid-base titrations

15.2 CHALLENGE

1 $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

2 0.62 M (2 sig fig)

3 15 M (2 sig fig)

15.2 SKILL DRILL

1 $Al(OH)_3(aq) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

2 0.0063 mol (2 sig fig)

3 0.011 mol (2 sig fig)

4 0.82 g (2 sig fig)

5 15 % (2 sig fig)

6 The experimental percentage yield of 14.67% is lower than the 48% claimed on the packaging of the antacid tablets.

7 All errors must focus on the percentage (and therefore number of moles of $Al(OH)_3$) being lower than predicted.

8 Student answers may vary.

15.2 CHECK YOUR LEARNING

1 Reduce errors and increase the precision of results

2 a A volume of solution delivered by the pipette

b The volume of solution delivered by the burette

3 The solution in the pipette would be diluted, resulting in a lower number of moles in the aliquot, a lower titre volume from the burette and therefore an underestimation of the unknown concentration.

4 This is a random error because it could be eliminated with repeats.

5 a $Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

b 0.018 mol (2 sig fig)

c 0.0091 mol (2 sig fig)

d 0.60 M (2 sig fig)

6 a 24.20 mL (4 sig fig)

b $Al(OH)_3(aq) + H_3PO_4(aq) \rightarrow AlPO_4(aq) + 3H_2O(l)$

c 0.484 M (3 sig fig)

d 3.87 M (3 sign fig)

7 a $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$

b 0.978 M (3 sig fig)

c The titration was not repeated to obtain concordant titres and therefore its precision cannot be determined.

8 a $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

b 0.0433 M (3 sig fig)

c The students were accurate because their answer agrees with the known concentration. Their precision cannot be determined unless the individual titre volumes are shown. Although their average titre resulted in an accurate concentration, it does not mean that this was an average of concordant titres.

d Systematic errors: parallax error from reading the burette; Random error: incorrectly reading the titre volume.

9 Student answers will vary.

Chapter 15 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Concordant titres reduce errors and increase the precision of the results by ensuring that titre volumes are within a 0.1 mL range.

12 a Random: a bubble in the burette that comes out during the titration, washing it with water but not the solution, rinsing with the incorrect solution.

Systematic: parallax errors, incorrect reading of the scale on the burette, not calibrating the burette.

b Random: washing it with water but not the solution, rinsing with the incorrect solution, not leaving the final drop in the bottom of the pipette.

Systematic: parallax errors, not calibrating the pipette.

c Random: not washing the flask before use to remove contaminants, not filling to the bottom of the meniscus, not dissolving all the solid properly, overfilling the volumetric flask so that the meniscus is above the line.

Systematic: parallax error, not calibrating the volumetric flask.

13 Student answers will vary.

14 Standard solutions have a highly accurately known concentration. By using a specific volume of this solution to reach the equivalence point in a reaction, the number of moles of other reactants and products can be calculated.

15 Primary standards must have a known chemical formula, a high state of purity, have a relatively high molar mass, be easy to store without reacting with the atmosphere or deteriorating, and be cheap and readily available.

16 a 12.28, 12.34 and 12.36 mL because they are within a range of 0.1 mL

b 12.33 mL

17 a Rinse with water to remove impurities

b Rinse with water, empty and then rinse with HCl solution and empty before dispensing a titre

c Rinse with water, empty and then rinse with NaOH solution and empty before measuring an aliquot.

18 Methyl orange or bromophenol blue: both change colour at the pH of the equivalence point

19 NaOH absorbs moisture from the atmosphere, so its concentration can change easily, and it does not have a high state of purity.

20 4.3 L(2 sig fig)

21 16 mL (2 sig fig)

22 0.2668 M (4 sig fig)

23 a A strong acid (due to a very low initial pH) and a strong base (due to a very high final pH)

b 7

c Bromothymol blue 6.0–7.6 or phenol red 4.8–8.4

24 a A strong acid (due to a very low initial pH) and a weak base (due to a low final pH)

b 4–5

c Bromophenol blue 3.0–4.6, methyl red 4.4–6.2 or bromothymol blue 6.0–7.6

25 0.3 M (1 sig fig)

26 a $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

b 0.211 M (3 sig fig)

27 a $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

b 0.4 M (1 sig fig)

c The burette was not washed with HCl before it was filled with HCl. This would dilute the HCl in the burette, decreasing the number of moles in a set volume, and increasing the volume of HCl required to neutralise the NaOH in the conical flask.

d The experimental concentration of NaOH will be higher than the actual (theoretical) concentration of NaOH.

28 61.5 g (3 sig fig)

29 There will be enough for four titrations of NaOH, but not five because the burette must be washed with this solution.

30 a KOH

b 0.02 mol (1 sig fig)

31 a No. They haven't used the concordant titres, and therefore did not obtain the correct average of 17.05 mL.

b 0.1 M (1 sig fig)

c Student answers will vary but should not focus on the solution in the burette, because the same solution was in the burette in the second titration. Instead, answers can include: parallax error in burette or pipette, judgment/reading error on the burette, or washing the pipette with water.

d Because no theoretical value is given, accuracy cannot be determined.

32 Student answers will vary based on the concentrations that students use, but should include how to make a standard solution, the equipment required, the steps of the titration and relevant stoichiometric calculations.

33 Student answers will vary, but can include: pregnancy testing and quality control (to verify that the claims on packaging are correct) for various acids and bases in cleaners etc.

34 a 33.36 mL (4 sig fig)

b $3KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$

c 0.50 M (2 sig fig)

d The experimental value of 0.50 M is the same as the theoretical value of 0.50 M, so the analysis was accurate. The results are precise as there are three concordant titres.

e The first and fourth titres are much larger than the concordant ones. Student answers may vary to account for these.

35 a $Ba(OH)_2$

b 0.0021 mol (2 sig fig)

c 3.3 g (2 sig fig)

d 22 mL (2 sig fig)

e In this example, the concentrations of both reactants are known. Therefore, this is not a titration. However, it is an example of volumetric analysis because it uses volume–volume stoichiometry calculations.

Chapter 16: Measuring gases

GROUNDWORK

16A $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

16B 9.82 mol (3 sig fig)

16C As temperature increases, the kinetic energy of particles increases.

16D Compressibility refers to the ability to reduce the space between individual particles. This can occur because gas particles are energetic particles with space between them.

16.1 Gases contributing to greenhouse effect

16.1 REAL-WORLD CHEMISTRY

1 Atmospheric gases are trapped in the layers of ice that build up each winter/summer cycle. Ice core samples can be taken and the gas extracted to determine the atmospheric carbon dioxide and methane levels. This data can be used to predict global temperatures from that time period.

2 If the concentration of greenhouse gases increases, more infrared radiation will get trapped, resulting in higher atmospheric temperatures.

3 'Climate action' aims is to reduce climate-related hazards and natural disasters by understanding climate change through analysis of the ice core data. Ice core samples provide a historical climate record and provide comparisons with current greenhouse gas levels in the atmosphere.

16.1 CHECK YOUR LEARNING

1 The glass of a greenhouse traps infrared radiation to keep the inside of the greenhouse warmer than the external environment. Carbon dioxide, methane and water vapour are called greenhouse gases because they also trap infrared radiation in the atmosphere.

2 Water vapour (0.4–1.0 %), carbon dioxide (0.0407%) and methane (0.00018%)

3 Greenhouse gas molecules absorb infrared radiation and prevent it from being released from the Earth's atmosphere.

4 There are natural levels of carbon dioxide (produced by cellular respiration of plants and animals) and water vapour (from the evaporation cycle and methane produced by animals, such as cows). The enhanced greenhouse effect refers to increased levels of these gases due to human activity and the trapping of more heat energy.

5 More heat energy will become trapped, which may increase global temperatures, water evaporation and the amount of water vapour in the atmosphere. This could then result in global warming, glacial melting, sea levels increasing, a loss of land, and an increase in climate-related hazards.

6 Student answers may vary.

16.2 Gas pressure and standard laboratory conditions

16.2 CHECK YOUR LEARNING

1 Gas pressure is the force exerted by gas particles on the walls of the container.

2 Standard laboratory conditions (SLC) define the temperature and pressure that gas experiments are conducted at, since gas particles behave differently depending on temperature and pressure. At SLC, gases are measured at 25°C and 100 kPa (298 K and 1 atm)

3 Kinetic molecular theory explains how particles behave in terms of kinetic energy. In the gaseous state, particles experience greater random motion and therefore have a high level of kinetic energy.

4 a Gas particles have very little attraction to one another.

b In the gaseous state, particles are in constant random motion. The particles in liquids and solids have less movement and greater attraction forces.

c The space between gas particles is much larger than the space that the particles themselves occupy. Therefore, gases expand to occupy the whole space of a container, unlike solids and liquids.

5 a The kinetic molecular theory states that gas particles have more kinetic energy when the temperature is higher. If the temperature decreases, the kinetic energy will also decrease.

b The kinetic molecular theory states that gas particles do not lose energy when they collide. If pressure is the result of collisions that have increased, then kinetic energy has increased.

c Kinetic molecular theory states that gases expand to fill a container. If the volume is unchanged, then the kinetic energy is unchanged.

6 As the kinetic energy of the molecules in the scented gas increases from the burning candle, the scented gas particles move more freely and expand across the room, filling the room with the scent.

16.3 The ideal gas equation

16.3A CHALLENGE

Start with the ideal gas equation, $PV = nRT$ where P = 1 atm = 101.3 kPa, V = 22.4 L, T = 0°C = 273 K and n = 1 mol. Rearrange the equation and substitute in the values:

$$R = \frac{PV}{nT} = \frac{100 \times 22.4}{1 \times 273} = 8.31\,\text{J mol}^{-1}\,\text{K}^{-1}$$

16.3 SKILL DRILL

1 Volume
2 Pressure
3 Temperature
4 The experiment is not valid because temperature was not kept constant.

16.3B CHALLENGE

CH_3CH_3, i.e. ethane.

16.3 CHECK YOUR LEARNING

1 (1) Particles do not interact with one another and (2) particles occupy no volume.
2 **a** Volume would double.
b Pressure would double.
c Volume would halve.
3 **a** 41.9°C (3 sig fig)
b 350 K (2 sig fig)
c 36.4 kPa (3 sig fig)
d 0.49 atm (2 sig fig)
e 0.025 L (2 sig fig)
4 **a** 79 L (2 sig fig)
b 2.2 L (2 sig fig)
c 0.386 L (3 sig fig)
5 **a** 0.213 g (3 sig fig)
b 5.11×10^{22} (3 sig fig)
6 **a** When temperature is increased, the kinetic energy of the gas particles increases. They collide more readily with the walls of a container, so pressure increases. They are proportional and are placed on opposite sides of the equals sign in the ideal gas equation: $P\_ = \_\_T$.
b When temperature is increased, the kinetic energy of the gas particles increases. They push against the walls of a container more and will cause an increase in volume of a variable volume system. They are proportional and are placed on the opposite sides of the equals sign in the ideal gas equation: $\_\_V = \_\_T$.
c When volume is increased, the space between gas particles and the walls of a container increases. The pressure will decrease since there are fewer collisions. They are inversely proportional and are placed on the same side of the ideal gas equation: $PV = \_\_T$.

16.4 Chemical reactions involving gases

16.4 SKILL DRILL

1 Atom economy, design for energy efficiency, prevention of wastes, use of renewable feedstocks
2 Student answers will vary.
3 Student answers will vary.
4 Academic institutions or journals, government organisations, media organisations

16.4 CHECK YOUR LEARNING

1 A reaction with oxygen to produce energy, water and carbon dioxide
2 They can produce carbon dioxide and water, which are both greenhouse gases.
3 Burning of fuels from renewable sources, such as plant-based biomass, may not contribute significantly to the greenhouse effect. For example, plants take up the CO_2 from the environment for photosynthesis and 'recycle' or remove CO_2 from the atmosphere.
4 Similarity: involve a hydrocarbon fuel and oxygen, and the production of water; Difference: a complete combustion reaction occurs in an excess of oxygen to produce CO_2, while an incomplete combustion reaction occurs when there is insufficient oxygen and may produce C, CO_2 and CO.
5 **a** $n(CO_2) = n(HC1) \div$ coefficient of HC1 $\times$ coefficient of CO_2
$n(CO_2)$ = 2 mol (1 sig fig)
b $n(CO_2) = n(CH_4) \div$ coefficient of $CH_4 \times$ coefficient of CO_2
$n(CO_2)$ = 7.4 mol (2 sig fig)
c $n(CO_2) = n(C_6H_{12}O_6) \div$ coefficient of $C_6H_{12}O_6 \times$ coefficient of CO_2
$n(CO_2)$ = 72 mol (2 sig fig)
6 **a** 1 mol (1 sig fig)
b 160 g (2 sig fig)
c 450 g (3 sig fig)
d 0.184 L (3 sig fig)
7 **a** 42.9 g (3 sig fig)
b 1.85×10^3 kg (3 sig fig)
c 3.65×10^3 kg (3 sig fig)
8 **a** Combustion of fuel sourced from renewable plant material, such as ethanol, produces carbon dioxide that is, theoretically, consumed when the plants are regrown to produce more fuel. This makes them renewable and contributes to a circular economy.
b Because the carbon dioxide produced by combustion of plant-based fuels is consumed again through photosynthesis, there is theoretically a net zero production of carbon dioxide. Therefore, fuels sourced from plant materials contribute towards carbon neutrality.
c To produce plant-based fuels, forests will need to be cleared and replaced with fast-growing plants to keep up with the speed of fuel consumption. This is likely to result in a loss of habitat for different species of fauna and flora.
9 Student answers will vary.

16.5 Calculating molar volume or mass of a gas

16.5 CHECK YOUR LEARNING

1 Molar volume is the volume that one mole of any gas will occupy at a specific temperature and pressure.
2 **a** 79 L (2 sig fig)
b 2.3 L (2 sig fig)
c 0.386 L (3 sig fig)
3 **a** 0.213 g (3 sig fig)
b 5.10×10^{22} atoms (3 sig fig)
4 3.00 L (3 sig fig)
5 60.0 L (3 sig fig)
6 **a** 28 000 g (2 sig fig)
b $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$
c 99 000 (2 sig fig)
d $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
e 4.7×10^3 kg (2 sig fig)
f 60 L (2 sig fig)
7 7.4 g (2 sig fig)
8 Student answers will vary.
9 If P decreases, V_m will increase.

Chapter 16 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Gas pressure is the force exerted by gas particles on the walls of the container.

12 Standard laboratory conditions (SLC) define the temperature and pressure that gas experiments are conducted at because gas particles are influenced by changes in temperature and pressure.

13 There is a natural level of carbon dioxide produced by cellular respiration of plants and animals, water vapour from the evaporation cycle and methane produced by animals. This produces an average temperature that sustains life on the planet.

14 **a** The volume is fixed, so when the temperature increases, the pressure increases.

b The system expands (volume increases) when the temperature increases, and the pressure will remain constant.

15 The temperature will increase because the volume is fixed.

16 The difference between SLC and STP is the temperature. Because the temperature is higher at SLC (25°C) than at STP (0°C), the particles will have more energy and the gas expands in order for the pressure to remain constant.

17 $V(H_2) = 22.5$ L; $V(O_2) = 22.4$ L; $V(He) = 22.3$ L; $V(Ar) = 22.4$ L

18 59.999 L (5 sig fig)

19 46.8 g (3 sig fig)

20 852 g (3 sig fig)

21 **a** 0.242 g (3 sig fig)

b 0.545 mol (3 sig fig)

c 10.6 g (3 sig fig)

d 3.39 g (3 sig fig)

e 8.71 mL (3 sig fig)

22 **a** 16.0 g (3 sig fig)

b 5.37 g (3 sig fig)

23 Reaction I

24 129 g (3 sig fig)

25 0.194 g (3 sig fig)

26 28 g (2 sig fig)

27 171 g (3 sig fig)

28 0.13 L (2 sig fig)

29 **a** 7.2 L (2 sig fig)

b 110 L (3 sig fig)

c 23.0 L (3 sig fig)

d 1.863 L (4 sig fig)

e 1.8 L (2 sig fig)

f 3.2 L (2 sig fig)

30 **a** Radon

b Argon

c Argon

31 Student answers will vary.

32 The temperature change creates a difference in pressure, and the sides of the bottle crumple inwards because the pressure in the bottle decreases.

Chapter 17: Analysis for salts

GROUNDWORK

17A An ion is an atom or a molecule with a net electrical charge.

17B Ionic compounds contain ions and are held together by the attractive forces between the oppositely charged ions.

17C A precipitation reaction is a reaction in which cations and anions in aqueous solution combine to form an insoluble ionic solid called a precipitate.

17.1 Sources of salts in water and soil

17.1 CHECK YOUR LEARNING

1 Primary salinity is the natural occurrence of salts in the landscape, whereas secondary salinity is when additional salts are added to the soil or water through human activity.

2 Heavy metals can make their way into the waterways through their use in construction, life and agriculture. Examples include:

- Mercury from industry emissions or spills can enter water through their uses in batteries, electrical components, fungicides and preservatives.
- Cadmium from impurities in the zinc of galvanised pipes or in fittings in water heaters, water coolers and taps. It is also from waste water, industry and fertilisers.
- Lead from natural sources, household plumbing sources and industrial waste

3 Because the light bulb of a circuit only indicates whether ions are present or not present in the water instead of providing a quantitative value for ion concentration.

4 **a** Lead should not exceed 0.01 mg L^{-1}, cadmium should not exceed 0.002 mg L^{-1}, arsenic should not exceed 0.01 mg L^{-1} and mercury should not exceed 0.001 mg L^{-1}.

b Sample 2 fails because cadmium is too high. Sample 1 fails because mercury is too high. Sample 3 passes.

c Samples 1 and 2 are not safe to drink because they contain unsafe levels of cadmium and mercury. Sample 3 is safe to drink.

5 Safe: Mg^{2+}, Na^+, Cl^-, K^+, Ca^+.
Toxic: Hg^+, NO_3^-, Cu^{2+}, Pb^{2+}

6 Because shallow-rooted plants require less water, this leads to excess water entering the groundwater system. The increased amounts of water pass through deep soils that have high concentrations of salts and absorb these salts as they are added into the groundwater system.

7 Because high-density metals can have a harmful toxic effect on living organisms. They might interfere with some biochemical pathways in the body and harm the body.

8 Testing for salinity and quality of water ensures that residents consume sufficient but not excess amounts of salt. It also ensures that heavy metals, which can be toxic to the body, are in a safe range. It is relevant to the goal because it enables access to clean water and sanitation.

17.2 Quantitative analysis of salts

17.2 CHECK YOUR LEARNING

1 A hydrated salt is a salt that contains water molecules incorporated into its lattice structure.

2 Gravimetric analysis is a set of methods used to determine an unknown mass of a substance in a sample. Two examples of gravimetric analysis are: i) heating a hydrated sample to remove water and calculating the molar ratio of water in the sample and ii) precipitating a component of the sample into a solid form and determining the composition of a sample.

The molar ratio of hydration of a salt can be determined by heating a sample to remove all the water and weighing the sample to determine the amount of dry salt remaining. The mass of a precipitate can be determined by filtering the participate and then heating the sample to remove the water and weighing the precipitate to calculate the mass of interested ion.

3 A calibration curve is a curve representing the relationship between the absorbance (interested outcome) and the known concentration of standard solutions. By plotting the calibration curve, the unknown concentration can be calculated.

4 **a** 0.00200 mol (3 sig fig)

b $NiCl_2{\bullet}6H_2O$

5 Colorimetry and UV-visible spectroscopy are both used to determine the concentration of a sample by measuring the amount of light absorbed by the sample. Colorimetry is used to determine the concentration of coloured compounds in a solution. UV–visible spectroscopy is similar to colorimetry but is far more selective because a much wider spectrum of visible and UV light is used for detection, allowing non-coloured compounds to be detected.

6 1.52 g

7 a $25\ mg\ L^{-1}$ (2 sig fig)

b $125\ mg\ L^{-1}$

c No, $<0.05\ mg\ L^{-1}$ is the allowable level in drinking water but the concentration in this water tank is much higher than the safe limit.

d Process:

- Use a colorimeter with the relevant filter.
- Use a series of standard solutions with known concentrations to construct a standard curve.
- Use absorbance to calculate the concentration of the chromium ions.

8 The sample for UV–visible spectroscopy can be prepared by dissolving it in a solvent and then scanning the sample to determine the highest absorbance wavelength. The measurement of the absorbance is performed at this wavelength. For colorimetry, if the metal cation does not have a strong colour, an oxidant or other reagent can be used to make a coloured complex in solution. This coloured metal complex can then be analysed by colorimetry.

9 a 11.9 g water (3 sig fig)

b 25 g (3 sig fig)

10 a 1.46 g (3 sig fig)

b 00.020 mol (2 sig fig)

c Ca

11 a $AgNO_3(aq) + Cl^-(aq) \rightarrow AgCl(s) + NO_3^-(aq)$

b 1.13 g (3 sig fig)

c 1.13% (3 sig fig)

d The weight of the precipitate would be higher than the actual amount. Therefore, the calculated chloride would be higher than the actual amount.

Chapter 17 review

MULTIPLE CHOICE

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

SHORT ANSWER

11 Electrical conductivity is used to determine the amount of ions in water and soil. As the amount of ions increases, the conductivity increases. This testing ensures that the amount of ions in water is within safe limits for residents.

12 The higher the concentration of salt in a sample, the higher the electrical conductivity. This occurs because salt ionises in water and the water is able to conduct electricity.

13 0.0004, 0.0008, 0.0012 and $0.0016\ mg\ L^{-1}$

14 To ensure that the amount of salts and toxic heavy metal in the water is below the safe allowable levels for drinking. It is also used to monitor how clean the water is.

15 The colour of the sample is the complementary colour of the colour that is absorbed by the sample. Therefore, to maximise the absorbance, a complementary coloured filter is used to ensure that the sample absorbs as much of the coloured light as possible. In colorimetry, absorbance is used to determine the concentration, so the complementary colour is used to maximise the absorbance.

16
- Heat a sample to remove all the water.
- Weigh the sample to determine the amount of water lost during heating and the amount of dry nickel(II) chloride remaining.
- Use the amount of water lost to determine the hydrated formula of the salt.

17 It is best to make another calibration curve using the same machine because with a different machine the strength of the light source and the monochromator might be slightly different, which might affect the result.

18 Because it is difficult to precipitate potassium ions or sodium ions out of solution. There are no negative ions that can precipitate potassium or sodium. Therefore, it is not ideal to use gravimetric analysis.

19 a 1:2

b $Na_2CO_3{\bullet}2H_2O$

20 a 1:4

b $MgCO_3{\bullet}4H_2O$

21 a 9.04 g (3 sig fig)

b 17.4 g (3 sig fig)

22 a 14.67 g (4 sig fig)

b $142.1\ g\ mol^{-1}$ (4 sig fig)

c Na

23 7.07 g (3 sig fig)

24 a $2AgNO_3(aq) + BaCl_2(aq) \rightarrow 2AgCl(s) + Ba(NO_3)_2(aq)$

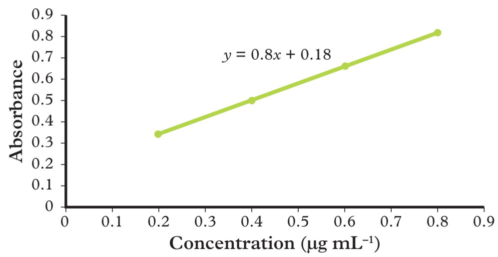
b AgCl

c 1.38 g (3 sig fig)

25 a $0.0097\ mg\ L^{-1}$ (2 sig fig)

b Sodium iodine could be used to react with Pb^{2+} ions, which would precipitate lead iodide. The precipitate can then be dried until the mass is consistent. The weight of the dried precipitate can then be used to calculate the amount of lead iodide and then the concentration of lead ion.

26 a



b $0.49\ \mu g\ mL^{-1}$ (2 sig fig)

27 a $CuCO_3(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2CO_3(aq)$

b New house 1: 8.81 g

New house 2: 6.05 g

Old house 1: 9.79 g

Old house 2: 7.46 g

c Technically, their hypothesis is correct because the four samples showed water in the old houses contained more copper than that in the new houses. However, the number of samples was not sufficient to draw a definite conclusion as one of the samples from the new houses had more copper than one of the samples from an old house. Therefore, more samples are needed.

d The mass of copper carbonate does not necessarily reflect the total amount of copper ions in the water because copper ions might exist as part of other salts in the water.

e The results would be higher than the actual amount because the mass of water would be included as the mass of the precipitate.

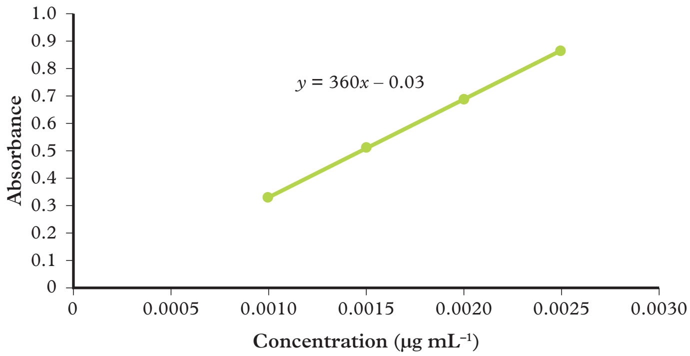
28 a Student answers will vary but should include the addition of silver nitrate in excess amount to precipitate the chloride ion.

b $AgNO_3(aq) + Cl^-(aq) \rightarrow AgCl(s) + NO_3^-(aq)$

c $13.1\ mg\ L^{-1}$ (3 sig fig)

d Colorimetry and UV–visible spectrometry can potentially be used. A metal complex can be formed with chloride ions to obtain a coloured solution. Known concentration of standard solutions can then be used to draw a calibration curve. The calibration curve can then be used to determine the unknown concentrations.

29 a



b Sample 1 = 0.0016 mg L^{-1} (2 sig fig)
Sample 2 = 0.0027 mg L^{-1} (2 sig fig)

c <0.002 mg L^{-1} is the allowable level in drinking water.
The water in sample 1 is safe to drink. The water in sample 2 is not safe to drink.

d The absorbance of sample 2 is outside the range of the calibration curve. One cannot be certain that the curve would still be linear. Therefore, standards of higher concentrations should be used to address this issue.

e Student answers will vary.

UNIT 2: Area of Study 2 Checkpoint

MULTIPLE CHOICE

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

SHORT ANSWER

1 a 8000 mg L^{-1} (1 sig fig)
b 0.8% (m/v) (1 sig fig)
c 2 mol L^{-1} (1 sig fig)

2 a $HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$
b 21.23 mL (4 sig fig)
c 2.17×10^{-3} mol (3 sig fig)
d 2.17×10^{-3} mol (3 sig fig)
e 0.108 M (3 sig fig)
f When the first permanent colour change is achieved, from magenta (in base) to colourless (in acid)
g The second student's results had only two concordant titres (of the four). Therefore, they were not precise because of the wider range of titre volumes. Their titre volumes were less precise because they were only recorded to one decimal place instead of two.

3 a i 0.48% (w/w) (2 sig fig)
ii Repeat the analysis at least two more times and average the results
b The concentration of phosphate ions in the dam water is higher than the limit and, therefore, does not comply.
c i $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
ii 1730 ppm (3 sig fig)

4

Quantity 1	>, <, or =	Quantity 2
Number of atoms of helium in 1.0 mol of He	=	Number of atoms of Ne in 1.0 mol of Ne
Number of molecules of CO_2 in 2 mol of CO_2	=	Number of oxygen atoms in 1 mol of CO_2
Volume of 0.20 M HNO_3 (a strong acid) required to completely react with 20.0 mL of 0.20 M KOH solution	=	Volume of 0.20 M HCN (a weak acid) required to completely react with 20.0 mL of 0.20 M KOH solution

5 a Unsaturated: a solution in which more solute can still be dissolved at a given temperature; Saturated: a solution in which the maximum amount of solute is dissolved at a given temperature; Supersaturated: a solution in which more solute is dissolved than theoretically possible at a given temperature
b Add a crystal of solute to each of the beakers. If the crystal dissolves, then the solution is unsaturated. If the crystal does not dissolve, the solution is saturated.

6 a C_2H_4F
b 0.0233 mol of vapour (3 sig fig); 94.0 g mol^{-1} (3 sig fig)
c $C_4H_8F_2$

7 a Compound Y
b 55 g/100 g H_2O
c 111 mL
d 46 g (2 sig fig)

UNIT 2 review

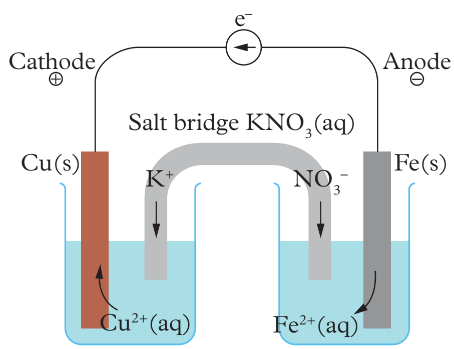
MULTIPLE CHOICE

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

SHORT ANSWER

1 a A strong acid readily donates one or more protons to a base and completely ionises in water.
b $NO_3^-(aq)$
c $HNO_3(aq)$
d They will have the same pH.
e i Zinc nitrate, $Zn(NO_3)_2(aq)$
ii Copper nitrate, $Cu(NO_3)_2(aq)$
iii Sodium carbonate, $Na_2CO_3(s)$

2 a $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
b $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
c $Fe(s) + Cu^{2+}(aq) \rightarrow$
$Fe^{2+}(aq) + Cu(s)$
d Iron is lower on the electrochemical series than copper, so it is more reactive and will undergo oxidation first; the iron will be a sacrificial anode and corrode before copper will.
e



3 a i 18 g $CaCl_2$/25 g H_2O (2 sig fig)
ii 8.5 g KNO_3/25 g H_2O (2 sig fig)
b i Unsaturated. For the solution to be saturated at 90°C, there would need to be 52 g $KClO_3$/100 g of water.
ii 32 g

4 a Phenol red or phenolphthalein
b 7.92 M (3 sig fig)
c If the burette is rinsed with water, this will dilute/lower the concentration of the NaOH solution. The titres will be higher in volume, meaning the calculated concentration of the final ethanoic acid will be lower.
If the pipette is rinsed with water, this will lower the concentration

of the ethanoic acid, meaning less NaOH will have to be added and the final calculated concentration will be lower.

The conical flasks can be rinsed with deionised water.

5 **a** Ethane

b 107 kPa (3 sig fig)

c $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

d 23.1 g (3 sig fig)

6 **a** 270 mg L^{-1} (3 sig fig)

b **i** 5

ii 1.4 g L^{-1} (2 sig fig)

c You could analyse the absorbance of two more solutions of known higher concentration to extend the calibration curve. This will allow you to read the absorbance from the curve to determine the unknown concentration.

GLOSSARY

%

% m/v
a unit of concentration measured in the mass of solute (g) per 100 mL of solution

% v/v
a unit of concentration measured in the volume of solute (mL) per 100 mL of solution

A

Aboriginal and/or Torres Strait Islander Peoples
people with family heritage from, and membership in, the ethnic groups that lived in Australia before British colonisation

absolute zero
the temperature, in kelvin, at which kinetic energy is the lowest; 0 K

absorbance
a measure of the amount of light that is absorbed by a sample during colorimetry of UV–visible spectroscopy

acid
a molecule that can donate a proton to a base

acid–base reaction
the reaction that occurs between an acid and a base

acid–base titration
a titration that uses an acid–base neutralisation reaction to find an unknown concentration

acid rain
rain with higher-than-normal acidity because of the atmospheric reactions of water with sulfur dioxide (SO_2) or nitric oxide (NO)

acid reflux
a digestive problem caused by the liquid content of the stomach flowing up into the oesophagus

acronym
an abbreviation formed from the initial letters of each word from a group of words

active ingredient
the compound in a medicine that is responsible for its medicinal effect

addition reaction
a reaction in which two or more molecules add together to form one larger molecule and no other products

adsorption
the attraction of a substance within a sample to the stationary phase; how the substance 'sticks' to the stationary phase

affinity
the strength of the interaction of a substance within a sample with the mobile or stationary phase

aim
a statement that describes the purpose of an investigation

aliquot
a volume of liquid delivered by the pipette

alkaline
having a pH greater than 7

alkane
a saturated hydrocarbon; contains only carbon–carbon single bonds

alkene
an unsaturated hydrocarbon; contains at least one carbon–carbon double bond

alkyl group
a group where a hydrogen atom has been removed from an alkane

allotrope
a different structural form of an element

amorphous
lacking a clearly defined structure or shape

amount
the quantity of a substance measured in moles

amphiprotic
can both donate and accept a proton, allowing it to act as either an acid or a base

anecdote
evidence from an individual's personal experience or observations that can be used to support a judgment or view

anhydrous
(of a crystalline compound) containing no water

anhydrous salt
an ionic compound from which the water has been removed

anion
a negatively charged ion formed after an atom gains electrons

anode
the negatively charged electrode

anomalous
unique; does not necessarily follow trends

antacid
a substance that works to neutralise an acid

anthocyanin
a water-soluble pigment that changes colour in response to pH

aqueous
the state given to a substance that is dissolved in water, (aq)

aqueous solution
the type of solution formed when water is used as the solvent

asymmetrical
opposite of symmetrical, made up of parts that are not mirror images of each other

atmosphere (atm)
a unit of pressure that is based on the atmospheric pressure at sea level; 1 atm is approximately 100 kPa, 100 000 Pa or 760 mmHg

atmosphere (atm) (unit of pressure)
a unit of pressure that is based on the atmospheric pressure at sea level; 1 atm is approximately 100 kPa, 100 000 Pa or 760 mmHg

atmosphere (of Earth)
the mixture of gases that surrounds Earth

atom
the smallest unit of matter, which consists of protons, neutrons and electrons

atomic mass unit (u)
a unit of mass where one atom of carbon-12 has a mass of exactly 12 u

atomic number
the number of protons in an atom

atomic radius
half the distance between two nuclei of the same element that are bonded together

Aufbau principle
a principle that states electrons always occupy available orbitals with the lowest energy

Avogadro's constant (N_A)
6.02×10^{23}

B

ball-and-stick model
a way of visualising the 3D structure of a molecule using balls to represent atoms and sticks to represent bonds

base
a molecule that can accept a proton from an acid

battery
a container in which reactants are stored to generate electrical energy from chemical energy

bent
arranged in a V-shape; the central atom is bonded to two atoms

bibliography
a full list of all the resources used in some research

bioaccumulate
when a substance concentrates in the bodies of living things

biocompatible
not harmful to the human body

biodegradable
the ability of a substance to be decomposed by bacteria or other living organisms, therefore avoiding pollution

biomass
material from animals and plants that can be fermented to derive fuel and other organic products

bioplastic
a plastic made from organic renewable materials

biopolymer
a polymer that occurs in nature or is made by a living organism, e.g. silk, DNA, cellulose, proteins; also known as natural polymer

boiling point
the temperature at which a substance changes state from liquid to gas

bond
a chemical bond formed by the electrostatic attraction of cations to anions

bond angle
the angle between adjacent bonds from the same atom

bonding electrons
valence electrons that are shared with another atom in a covalent bond

by-product
a secondary product formed in a chemical reaction

C

calibration curve
a graph of absorbance versus concentration of standard solutions so that an unknown concentration can be determined

capillary action
the movement of a liquid through a narrow space (e.g. the cellulose network in absorbent paper) without any help, usually against gravity

carbon-12 (^{12}C)
an isotope of carbon with 6 protons, 6 neutrons and 6 electrons

carbon–oxygen cycle
the cycle in which atmospheric oxygen is converted to carbon dioxide in animal respiration and regenerated by plants during photosynthesis

carboxyl
a functional group consisting of a carbon double bonded to an oxygen and single bonded to a hydrogen (–COOH)

cathode
the positively charged electrode

cation
a positively charged ion formed after an atom loses electrons

chemical species
the general name given to an atom, element, ion, molecule or compound

chemistry
the study of matter, including its structure, properties and behaviour

chromatogram
the pattern of bands, spots or peaks formed on the chromatography paper or TLC plate, demonstrating the separation of a mixture

chromatography
an analytical technique that separates the components of a sample mixture based on the properties of the molecules

circular economy
a model of production and consumption that involves sharing, leasing, re-using, repairing, refurbishing and recycling existing materials and products as long as possible

climate change
the gradual increase in the temperature of the Earth's surface, oceans and atmosphere, and consequent changes in climate; generally caused by the enhanced greenhouse effect

collate
collect and combine texts, information or data

colorimetry
a method for determining the concentration of a coloured compound in solution by measuring absorption of a particular wavelength of light

column
the tube-like structure that contains the stationary phase in column chromatography, through which the mobile phase and sample flow

column chromatography
a technique used to separate and purify individual components from mixtures of compounds

combustion
a chemical reaction with oxygen to form a metal oxide, covalent compound or carbon dioxide and water

command word
a word that provides direction on what you need to do, e.g. identify, explain, analyse, evaluate

complex half-equation
a half-equation for an oxidation and reduction equation that occurs under acidic or alkaline conditions

complex overall redox equation
an overall redox equation for a redox reaction that occurs under acidic or alkaline conditions

compostable
can break down and safely decompose in compost, and can be beneficially added to soil, leaving no trace

compound
two or more atoms bonded together, where the atoms belong to two or more different elements, e.g. O_2 is a molecule, but not a compound; HCl is both a molecule and a compound

concentrated
a solution that contains many particles per unit of volume

concentration
the amount of solute in a specific volume of solution

concordant
volumes of three or more titres that fall within 0.1 mL of each other

condensation reaction
a reaction in which two molecules with functional groups react to form a larger molecule and a small molecule by-product

condensed semi-structural formula
a simpler form of the semi-structural formula, in which brackets are used to show repeating units

conjugate acid
a molecule formed when a base accepts a proton

conjugate base
a molecule formed when an acid loses a proton

conjugate pair
two compounds that differ by the presence or absence of an H^+ or a proton in their formulas

conjugate redox pair
a pair of chemical species that represents the gain or loss of electrons, e.g. Na/Na^+ and O/O^{2-}

continuous data
data or information that can be any numerical value

controlled variable
a variable that is not changed or is constant during an investigation

coolant
a substance that removes heat

copolymer
a polymer made up of two or more different monomers

core charge
the number of protons minus the number of inner electrons

corrosion
the degradation of a metal to form a more stable metal oxide when exposed to gases and liquids

cosmetic
a product used to help maintain or improve appearance

counting prefix
letters put at the start of a chemical name to indicate the number of branches or functional groups; e.g. 2 = di, 3 = tri, 4 = tetra

covalent bond
a bond that forms between two nonmetal atoms when they share one or more pairs of electrons

covalent molecular substance
a molecule that is covalently bonded because of the sharing of valence electrons

critical element
a highly used element at risk of depletion in the near future

crystal
solid material that forms

crystal grain
an organised lattice of metal cations

crystal lattice
a three-dimensional arrangement of atoms or ions that repeats to make up a compound

crystalline
having a set, specifically ordered 3D structure

cure
a process in polymer manufacturing in which material is toughened

D

dalton (Da)
a unit of mass where one atom of carbon-12 has an mass of exactly 12 Da

data evaluation
the critical analysis of data that has been personally collected or provided to identify contradictory or incomplete data, or issues such as a personal bias

decomposition
the breakdown of a molecule into smaller molecules

degree of hydration
the amount (in moles) of water absorbed by a crystalline compound per formula unit

degree of saturation
the amount of solute dissolved in a volume of solvent

delocalised electron
an electron that is not linked to a particular atom or single covalent bond; it is free to move within the structure

density
the amount of matter contained in a specific volume; the compactness of a substance

dependent variable
the variable that is observed or measured when the independent variable is changed during an investigation

desorption
the release of a substance within a sample from the stationary phase into the mobile phase; how the substance 'unsticks' from the stationary phase

diatomic molecule
a molecule made of two atoms, which can be of the same element or different elements

dilute
a solution that contains few particles per unit of volume

dilution
a decrease in the concentration of a solution by adding solvent

dilution factor
the ratio of the final volume to the initial volume in a dilution

dipole
an uneven separation of charge; formed when a bond or a molecule has a partial positive charge ($\delta+$) at one end and a partial negative charge ($\delta-$) at the opposite end

dipole–dipole attraction
an intermolecular force that forms as a result of attraction between permanent dipoles in polar molecules

diprotic
can donate two protons per molecule

discrete data
data or information that can only be certain numerical values

dispersion force
an intermolecular force that results from attraction between temporary dipoles in polar and non-polar molecules

displacement reaction
a type of redox reaction in which a more reactive element replaces (or displaces) a less reactive element to form a new compound; the metal ions of one metal must be above the other metal in the reactivity series of metals

dissociate
break apart into smaller atoms, ions or molecules

dissolve
when solute–solute bonds break and the solute molecules mix evenly with the solvent molecules

double covalent bond
a bond that forms between two non-metal atoms when they share four electrons (two pairs)

double displacement reaction
a chemical reaction that occurs when two reactants exchange cations or anions to form two new products

ductile
can be drawn into wire

E

elastomer
a polymer that has elastic or stretching properties

electrical conductivity
the ability to allow charged particles to move through a substance

electrical insulator
a material that does not allow the movement of charged particles through it

electrochemical series
a table of the strongest oxidising and reducing agents written as reversible half-equations; the equations are always written with the strongest oxidising agent at the top left of the table and the strongest reducing agent at the bottom right

electrode
a strip or rod that conducts electricity; this is the surface on which the oxidation or reduction reactions occur

electrolyte
an electrically conductive solution that contains free moving charged ions

electron
a negatively charged subatomic particle

electron pair
a pair of electrons that occupy the same shell and can be part of a bond with another atom or non-bonding as a lone pair

electron transfer diagram
a visual representation that shows the movement of electrons between atoms and/or molecules

electron-dense area
a region of a molecule where there are one or more pairs of bonding or non-bonding electrons; e.g. a lone pair, or electrons in a single, double or triple covalent bond

electronegativity
the tendency for an atom to attract electrons when in a chemical bond

electronic configuration
the distribution of electrons in an atom

electrostatic attraction
the attraction force between oppositely charged particles

element
a pure substance that consists of only one type of atom

elute
come out of the bottom of a chromatography column

empirical formula
the simplest whole-number ratio of atoms in a compound

end point
the point during an acid–base reaction where the colour of an indicator changes

enhanced greenhouse effect
the accelerated warming of Earth due to human activities increasing the amount of greenhouse gases in the atmosphere

enzyme
a substance produced in living things that speeds up chemical reactions

equilibrium arrows
reactions arrows that point in both directions to indicate that a reaction is reversible

equivalence point
the point in a titration where the acid and base are mixed in exact stoichiometric ratios given by the balanced chemical equation

error
the difference between an accepted or theoretical value and the experimental, observed or measured value

ethics
moral principles that govern a person's behaviour or how an activity is conducted; the branch of knowledge that deals with moral principles

evidence
information or data collected on a topic that can help form a conclusion

excess reagent (reactant)
a reactant that is available in a greater amount than is needed for a reaction and is not completely consumed

excited state
the state of an atom when electrons are not in their highest energy state

expanding question
a smaller and more specific question that you can derive from the investigation question to help you answer it

external circuit
the pathway for movement of electrons from the anode to the cathode

F

feedstock
a raw material used to produce other goods

fermentation
a chemical process in which sugars are broken down in the absence of oxygen

first ionisation energy
the energy required to remove the first valence electron

fixed volume (rigid wall system)
a system where the volume of the container does not change

flammable
can be burned in a combustion reaction

fluorescent
emitting visible light when exposed to radiation

food additive
a compound that is added to food to enhance flavour, nutritional value, appearance, texture or safety

fossil fuel
a mixture of hydrocarbon compounds that are mined from the Earth's crust; includes coal, natural gas, crude oil

fossil-fuel-based plastic
a plastic made from petrochemicals

frequency
the number of times an observation occurred in an experiment or a study

functional group
a group of atoms that is responsible for the characteristic reactions of a compound

G

g L^{-1}
a unit of concentration measured in the number of mass of solute (g) per litre of solution

galvanic cell
a circuit in which a redox chemical reaction occurs to generate electrical energy; the reduction and oxidation half-reactions occur in separate compartments (or cells)

gas laws
the laws that explain the properties of gases and how they behave at different temperatures, volumes and pressures

gas pressure
the force exerted, per unit area, by a gas particle on another gas particle or the wall of a container

general formula
a rule allowing you to determine the molecular formula of an organic compound; each hydrocarbon family has a different general formula

grain boundary
the point at which different crystal grains meet in a metal

gravimetric analysis
the quantitative determination of the mass of a substance

greatest common divisor (GCD)
the largest positive integer that divides into all the numbers in the set (e.g. 6 is the GCD of 12 and 18)

green chemistry
an area of chemistry that focuses on designing safer and more sustainable new products

greenhouse effect
a phenomenon where greenhouse gases (e.g. methane, carbon dioxide, water vapour) trap heat in the atmosphere and cause the Earth to warm

greenhouse gases
the gases responsible for trapping the Sun's UV rays

ground state
the state of an atom when electrons are in their lowest energy state as predicted by the Aufbau principle

groundwater
the water beneath the surface in soil, in rocks or soil pore spaces

guiding word
a word used to ask a question, e.g. how, what, when, why, where, who

H

half-cell
a compartment that contains either the oxidation or the reduction half-reaction

half-equation
an equation that represents either the oxidation or the reduction half of a chemical equation; it shows the transfer of electrons

haloalkane
an alkane that contains a halogen functional group

halogen
a group 17 atom, including fluorine, chlorine, bromine and iodine

hardness
the ability to withstand deformation and scratching

herbicide
a pesticide that specifically controls plants

high-performance liquid chromatography (HPLC)
a form of column chromatography in which the mobile phase is pumped through a column at a controlled flow rate; the particles of the stationary phase are much smaller than in column chromatography and are densely packed to make separation more sensitive and efficient

homologous series
a family of compounds that contain the same functional groups and have similar properties; they differ by a $-CH_2$

hydrate
an ionic compound that has water molecules surrounding the charged ions

hydrated salt
an ionic compound that contains water molecules in a set mole ratio within the ionic lattice

hydrocarbon
an organic compound that contains one or more carbon atoms bonded to hydrogen atoms

hydrogen bond
an intermolecular force between a hydrogen atom that is connected to a fluorine, oxygen or nitrogen atom in one molecule, and a fluorine, oxygen or nitrogen atom on a nearby molecule, which has at least one lone pair of electrons

hydrolysis reaction
a chemical reaction in which water breaks down a larger molecule into smaller molecules

hydroxyl
a functional group consisting of an oxygen bonded to a hydrogen ($-OH$)

hypothesis
a testable statement that includes a prediction about the outcome of an investigation based on scientific reasoning

I

ideal gas
a gas in which particles do not interact with one another and occupy no volume

ideal gas equation
an equation that relates temperature, pressure, volume and moles of a gas; also called the universal gas equation

impurity
a small amount of chemical that has contaminated a pure substance

in-text reference
an acknowledgement of the source immediately after the research or information is referred to

incomplete combustion
a chemical reaction in which limiting oxygen reacts with a hydrocarbon to form either carbon monoxide or carbon and water

independent variable
the variable that is changed or manipulated during an investigation

indicator
a substance that undergoes a distinct observable change (often in colour) when pH conditions change

infrared radiation
waves of energy emitted from Earth

inner shell
any shell of electrons that is not a valence shell

insecticide
a pesticide that specifically controls insects

insoluble
will not dissolve

instrument error
an error due to the limited accuracy of a scientific instrument; usually, very small (and not significant) if used with standard operating and/ or laboratory procedures

intermolecular force
an attraction between molecules

internal circuit
the pathway for movement of ions between half-cells

intramolecular bond
a bond within a molecule

ion
a charged atom formed by gaining or losing electrons

ion–dipole attraction
the interaction between an ion with a whole charge and the partial charge caused by a polar bond or molecule

ionic bond
a chemical bond that results from the electrostatic attraction between a positive metal cation that has lost electrons and a negative non-metal anion that has gained electrons

ionic compound
a chemical compound that is held in place by ionic bonds

ionic product of water
the product of H_3O^+ and OH^- molar concentrations

ionic substance
a cation (metal) interacting with an anion (non-metal)

ionisation reaction
a reaction that results in molecules losing or gaining an electrostatic charge

ionise
form charged ions

isolated system
a thermodynamic system that cannot exchange matter or energy outside its boundaries

isotopes
atoms of the same element that have different numbers of neutrons

IUPAC systematic nomenclature
a set of rules set by the International Union of Pure and Applied Chemistry to consistently name organic compounds

J

jargon
special words or expressions used by a profession or group that are difficult for others to understand

K

kelvin
a unit of temperature that is based on levels of kinetic energy; one kelvin is equal to one degree Celsius

key term
a word that relates to the key topic you are investigating

kinetic energy
the energy associated with objects based on their movement or motion

kinetic molecular theory
the theory that states that all particles are in constant random motion

L

latent heat
the amount of energy, in kJ, required to change the state of one mole of a substance

latent heat of fusion
the amount of energy, in kJ, required to melt one mole of substance from solid to liquid state at its melting point

latent heat of vaporisation
the amount of energy, in kJ, required to evaporate one mole of substance from liquid to gaseous state at its boiling point

lattice
an interlaced structure consisting of regular repeated atoms or molecules

law of conservation of mass
a law that states that in an isolated system, mass cannot be created or destroyed

Lewis structure
a way of representing the structure of a covalent molecule that shows all bonding and non-bonding valence electrons as dots or crosses; also called an electron dot structure

limewater
a saturated aqueous solution of calcium hydroxide

limiting reagent (reactant)
a reactant that is completely consumed in a reaction and which determines the amount of product formed or excess reactant used

line of best fit
a line drawn through a scatterplot of data points that expresses a relationship between those points

linear
arranged in a straight line; the central atom is bonded to one or two atoms

linear economy
a way of managing resources that operates on a take–make–dispose model; new resources are used and disposed of after use

litmus paper
treated paper that changes colour in acidic or basic solutions

livestock
agricultural animals kept for produce such as meat, milk, eggs and fur

logarithmic scale
a non-linear scale used to display numerical data over a wide range of values in a neat way

lone pair
a pair of valence electrons that are not shared with another atom in a covalent bond

M

m/z
the mass-to-charge ratio; the x-axis on a mass spectrum; charge is +1 so m/z represents the relative isotopic mass of each isotope

main chain
the longest hydrocarbon chain

malleable
can be shaped or hammered without breaking

mass number
the sum of protons and neutrons in an atom

mass spectrometer
an apparatus for measuring the masses and relative abundances of isotopes (and molecules and molecular fragments) by ionising them and determining their trajectories in electric and magnetic fields

mass spectrometry
an analytical technique in which a sample is bombarded with electrons to form charged fragments; the fragments are analysed and put together like a puzzle to identify the substances present

mass spectrum
a column graph that shows the relative abundance of each isotope in a sample of an element

medicine
a compound or mixture of compounds used to treat disease

melting point
the temperature at which a substance changes state from solid to liquid

metabolism
the chemical processes in the body that convert energy from food into energy that our cells can use

metal oxide
an ionic compound where a metal cation is electrostatically attracted to an oxide anion (O^{2-})

metallic bond
a chemical bond that results from the electrostatic attraction between positive metal cations and negative delocalised electrons

metallic character
how readily an atom loses a valence electron

metastable
a highly energetic and unstable state; in a metastable supersaturated solution, the dissolved solute is likely to ‘crash out’ of solution and form crystals

methodology
a system of methods used in a particular area of study or activity

methyl group
a hydrocarbon branch consisting of one carbon and three hydrogen atoms

millimetres of mercury (mmHg)
a unit of pressure that is based on the pressure needed to raise the volume of mercury by 1 mm in a manometer

mobile phase
the solvent phase that flows, moving the components of a sample over the stationary phase

molar mass
the mass in grams of one mole of substance

molar volume
the volume occupied by one mole of a gas at a specified temperature and pressure

molarity (mol L^{-1} or M)
a unit of concentration measured by the number of moles per litre of solution; also known as molar concentration

mole
an amount of substance that contains 6.02×10^{23} particles

molecular formula
a way of representing the structure of a covalent molecule that shows the number and type of atoms in the molecule

molecular shape
the arrangement of atoms in an individual covalently bonded molecule

molecular substance
a molecule that is covalently bonded because of the sharing of valence electrons

molecule
a group of atoms covalently bonded together

monatomic
made from one type of atom

monomer
a small molecule that reacts with other molecules that are the same or different to form a polymer

monoprotic
can donate one proton per molecule

N

natural acidity
the slight acidity of rain caused by the reaction of CO_2 with water

natural indicator
an indicator from natural sources such as plants

net ionic equation
a chemical equation in which electrolytes in aqueous solution are expressed as dissociated ions

neurological disorder
a disorder that affects the brain or nervous system

neutral
a solution which is neither acidic nor basic

neutralisation
the reaction of an acid with a base that produces a metal salt and water

neutron
a neutral subatomic particle

new scrap
scrap metal that comes from the manufacturing of metal products

non-bonding electrons
valence electrons that are not shared with another atom in a covalent bond

non-ferrous scrap
scrap metal that does not contain iron

non-metallic character
how readily an atom gains a valence electron

non-polar
not having a positive and a negative end

non-polar covalent bond
a covalent bond in which electrons are shared equally; the two bonded atoms have equal electronegativity

non-renewable
resources that take an extremely long time to replenish and are available in a limited supply

nucleus
the dense region of an atom, which consists of protons and neutrons

nuclide notation
an abbreviated way to show basic information about an atom or ion, e.g. $^{7}_{3}Li^{+}$

O

ocean acidification
the continual decrease of ocean pH caused by increased atmospheric carbon dioxide levels

octet rule
a rule that states atoms gain or lose electrons to have eight electrons in their valence shell (excluding hydrogen and helium)

old scrap
used scrap metal that has been collected from consumers

opinion
a person or organisation's judgment or view on a topic; may not necessarily be based on evidence or fact

orbital
a subdivision of an electron shell where two electrons can orbit the nucleus; each electron shell contains 1, 3, 5 or 7 orbitals

organic compound
any compound that contains carbon

organometallic compound
a chemical compound that has at least one bond between a metal and a carbon that is part of an organic group

origin
the line applied to a chromatogram to mark the point where the sample or standard is placed

overall dipole
the presence of positive and negative charges at opposite ends of a molecule; unbalanced sharing of electrons

overall redox equation
an equation formed by combining the two half-equations; spectator ions are not shown, and electrons are balanced and cancelled out

oxidation
a process in which one chemical species loses electrons to another

oxidation number
the number of electrons that an atom gains or loses to form a chemical bond with another atom

oxidising agent
a chemical species that causes oxidation and is itself reduced

P

paint
a liquid mixture that, when dry, enhances the appearance of objects or protects them from the environment

paper chromatography
an analytical technique for separating and identifying mixtures; the stationary phase is a thin strip of absorbent paper

parallax error
an error made by not having your eye directly in line with the measurement

particles
atoms, molecules, ions or electrons

parts per million (ppm) or mg L^{-1}
a unit of concentration measured in the mass of solute (mg) per litre of solution

pascal (Pa)
a unit of pressure that is the force of one newton per square metre of a substance; 1 kilopascal (kPa) is equal to 1000 pascals (Pa)

percentage abundance
the abundance of each isotope, where the sum of all the percentage abundances is exactly 100%

percentage composition
percentage by mass of each element in a compound

percentage error
the calculated percentage difference between an observed or measured value and an accepted or theoretical value

pesticide
a compound or mixture of compounds used to protect plants or animals by controlling unwanted plants or pests

pH
a qualitive measure of H^{+} concentration

pH meter
an electronic device that measures hydrogen ion activity (acidity or alkalinity) in solutions

phase (in chromatography)
a form of matter that has uniform chemical and physical properties; can be a pure substance or a mixture

photochemical smog
a haze formed when volatile organic compounds react with nitrogen dioxide, ozone or other volatile organic compounds in the air

photosynthesis
a chemical process in which energy from sunlight is converted to sugars and oxygen

physical property
a characteristic that can be observed or measured without changing the identity of the substance

pigment
a chemical compound that gives colour

plant-based biomass
specifically, biomass that comes from plants

pOH
a qualitive measure of OH^- concentration

polar
having a positive end and a negative end

polar covalent bond
a covalent bond in which electrons are not shared equally; the two bonded atoms have different electronegativities

polarity
the separation of electric charge in a bond or molecule that depends on the electronegativity of its atoms

polarity index
a measure of a substance's polarity determined by the resistance when a voltage is applied

polyatomic
made from more than one type of atom

polyatomic molecule
a molecule made up of three or more atoms

polymer
a large molecule made up of many repeating units called monomers

polymerisation
a chemical reaction in which monomers are connected to form a polymer

polyprotic
can donate more than one proton per molecule

potable
safe to drink

precipitate
the solid product formed from a precipitation reaction

precipitation reaction
a reaction between two soluble ionic substances that forms an insoluble product

preservative
a substance that prolongs the lifetime of a product by protecting it from harmful bacteria

primary cell
a type of galvanic cell that cannot be recharged; it is disposed of after the redox reaction is complete

primary standard
a very pure substance that is used to make a primary standard solution

procrastination
the action of delaying or postponing something

proton
a positively charged subatomic particle

pyramidal
arranged in a triangular shape; the central atom is bonded to three atoms with one lone pair

Q

qualitative data
data that is not numerical; can be text, images or audio

quantitative data
data that can be counted or measured and is expressed as numbers

R

radioisotope
an isotope with too many or too few neutrons, which breaks down by undergoing radioactive decay

ratio statement
an equation used to determine the quantity of an unknown substance (usually in moles) relative to another substance in the same chemical reaction

raw material
a compound used as starting materials for the synthesis of larger and more complex compounds

reactivity
how readily a chemical substance undergoes a chemical reaction

reactivity series
an analytical arrangement of metals from lowest to highest reactivity

reactivity series of metals
a list that orders the metals in the periodic table from most reactive (strongest reducing agent) to least reactive (weakest reducing agent)

reading error
an error made by approximating a measurement if it falls between the markings of a measuring device

redox
a chemical reaction involving the transfer of electrons from one reactant to another

reducing agent
a chemical species that causes reduction and is itself oxidised

reduction
a process in which one chemical species gains electrons from another

relative abundance
the abundance of each isotope, where the abundance of the most abundant isotope is set to a value of 100%; the sum of all the relative abundances of all the isotopes will be greater than 100%

relative atomic mass (RAM)
the weighted mean of all the relative isotopic masses of an element

relative isotopic mass (RIM)
the ratio of the average mass of one atom of an isotope to one-twelfth of the mass of an atom of carbon-12; RIM has no units and has a value equal to the mass of the atom in Da or u

renewable
resources that are available in a continuous/unlimited supply or can replenish naturally

research outline
a tool to help you organise your ideas; can be a skeleton for presenting information later

retardation factor (R_f)
the ratio of the distance travelled by a component of a sample, from the origin, in relation to the distance travelled by the mobile phase

retention time (R_t)
the time that a component is retained by a chromatography column

risk assessment
a process of evaluating the potential risks that may be involved in an activity, e.g. performing an experiment

S

salinity
the concentration of salt in water or soil

salt bridge
a connection between solutions in a galvanic cell that allows the flow of charge by acting as a pathway for moving ions

saturated (compound)
containing only carbon–carbon single bonds

saturated (solution)
a solution in which the maximum amount of solute is dissolved

saturation point
the point at which the maximum amount of solute has dissolved in the known quantity of solute at a given temperature

scientific method
a series of steps used to acquire knowledge in science, which involves observation, developing and testing hypotheses; collecting, analysing and communicating results

secondary cell
a type of galvanic cell that can be recharged; it can be reused because the redox reaction is reversible

self-ionisation
when water molecules interact with each other to form positive and negative ions

semi-structural formula
a way of representing the structure of an organic compound, in which the positions of all the atoms and functional groups, but not the bonds, are shown

separation science
the science of separating a mixture on the basis of the different properties of its components

shielding effect
the shielding of valence electrons by inner electrons that alters the nuclear charge felt by valence electrons

significant figures
the digits of a number that are used to express it to the required degree of accuracy, starting from the first non-zero digit

single covalent bond
a bond that forms between two non-metal atoms when they share two electrons (one pair)

skeletal structure
a simpler form of the structural formula, in which carbon atoms are represented as vertices and hydrogen atoms bonded to these carbon atoms are not shown

slightly soluble
only partially dissolves in a solvent

solar radiation
sunlight or waves of energy emitted from the Sun

solubility
the ability of a substance to dissolve

solubility graph or curve
a graph that shows the relationship between the solubility of substances and temperature

soluble
will dissolve

solute
the substance being dissolved in the solvent

solution
an even mixture that forms when a solute dissolves in a solvent

solvent
the substance that another is being dissolved in, e.g. water

solvent front
the point the mobile phase reaches on a chromatogram before the analysis is terminated

specific heat capacity
the energy in joules required to increase the temperature of 1 g of a substance by 1°C

spectator ion
an ion that does not participate in a reaction and has the same state and oxidation number as both a reactant and a product

standard laboratory conditions (SLC)
standard conditions for temperature and pressure under laboratory conditions; gases are measured at 1 atm and 25°C or 298 K

standard solution
a solution with an accurately known concentration

standard temperature and pressure (STP)
standard conditions for temperature and pressure; gases are measured at 1 atm and 0°C or 273 K

starting point
the carbon that is the nearest to an alkyl branch, functional group or double bond; this is the first carbon

stationary phase
the solid phase to which the components of a sample are adsorbed; can sometimes be a liquid coated onto a solid support

stoichiometry
the numerical relationship between the amounts of reactants and products in a reaction

stomach acid
a digestive fluid consisting of salts and HCl that aids in the breakdown and digestion of food

strong acid
an acid that completely dissociates in water

strong base
a base that completely dissociates in water

structural formula
a way of representing the structure of a covalent molecule that shows the covalent bonds between the atoms as lines

structural isomers
compounds with the same molecular formula, but a different structure

subatomic particle
a particle that makes up an atom, e.g. a proton, a neutron and an electron

sublime
to change state directly from a solid to a vapour

subshell
a subdivision of an electron shell that can hold a maximum of 2, 6, 10 or 14 electrons depending on its type (s, p, d or f)

supersaturated
a solution in which more than the usual maximum amount of solute is dissolved because a saturated solution has been heated to dissolve more solute and on cooling the solute stays in solution

sustainability
using natural resources more efficiently, to ensure supply of resources is continued into the future

symmetrical
made up two halves that are mirror images of each other

synthesis
a chemical reaction in which two atoms, elements or molecules react to form a new molecule; also known as addition

synthetic polymer
an artificially produced or synthesised polymer; generally derived from fossil fuels, but increasingly sourced from biomass

system
the gas particles and the container they occupy

T

temperature
a measure of the average heat energy of the particles within a system

temporary dipole (or dipole moment)
when a bond gains a temporary negative charge at one end and a temporary positive charge at the other end

terminal carbon
the end carbon in a hydrocarbon chain

tetrahedral
arranged in a triangular pyramid shape; the central atom is bonded to four atoms

thermoplastic
a polymer that has weak intermolecular forces between the polymer chains; also known as linear

thermosetting
a polymer that has strong covalent bonds between the polymer chains; also known as cross-linked

thin-layer chromatography (TLC)
an analytical technique for separating and identifying mixtures; the stationary phase is typically a thin layer of silica gel, aluminium oxide or cellulose supported on a piece of glass or plastic

titrant
the solution being delivered by burette in a titration

titration
a quantitative analytical technique that is used to find the unknown concentration of a solution

titre
the volume delivered by the burette in a titration

total artificial heart
a device that replaces the ventricles of a heart; used for patients whose ventricles are no longer functional

triple covalent bond
a bond that forms between two non-metals when they share six electrons (three pairs)

triprotic
can donate three protons per molecule

turbidity
the cloudiness (or opacity) of a liquid due to suspended solids

two-dimensional paper (or thin-layer chromatography)
an analytical technique used to separate components; a first paper or thin-layer chromatography analysis is completed, and the sheet is rotated 90° and a second analysis is run using a different mobile phase

U

universal gas constant (*R*)
the proportionality constant (8.314 J mol^{-1} K^{-1}) that is used to define gas behaviour under ideal conditions

universal indicator
a pH indicator consisting of several compounds that can present a range of colours in response to acidic, basic or neutral solutions

unsaturated (compound)
containing at least one carbon–carbon double bond

unsaturated (solution)
a solution in which more solute can be dissolved

V

valence electron
an electron in the outermost shell of an atom

valence shell
the outer shell of an atom where electrons are found

valence shell electron pair repulsion (VSEPR) theory
a theory that states that the valence electron pairs (either bonding or non-bonding) arrange themselves to be as far apart as possible

valence structure
a way of representing the structure of a covalent molecule that shows the covalent bonds between all atoms as lines, and lone pair electrons as dots or crosses

variable volume system
a system where the volume of a container can change to accommodate the gas

ventricle
a chamber of a heart

volatile organic compound
an organic compound that can easily evaporate and become vapour

volume
a measure of the space occupied by a substance

volume–volume stoichiometry
calculations using the mole ratio for a chemical reaction where one solution with known volume and concentration is used to find the unknown volume or concentration of another solution

volumetric analysis
a quantitative analytical technique for determining concentration of a solution by titrating it against another solution of known concentration and volume

W

water table
the underground boundary between the soil and the area where groundwater saturates spaces between sediments and cracks in rock

weak acid
an acid that does not completely dissociate in water

weak base
a base that does not completely dissociate in water

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APPENDIX

Periodic table

Period	1 Group	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.0 Hydrogen																	2 He 4.0 Helium
2	3 Li 6.9 Lithium	4 Be 9.0 Beryllium											5 B 10.8 Boron	6 C 12.0 Carbon	7 N 14.0 Nitrogen	8 O 16.0 Oxygen	9 F 19.0 Fluorine	10 Ne 20.2 Neon
3	11 Na 23.0 Sodium	12 Mg 24.3 Magnesium											13 Al 27.0 Aluminium	14 Si 28.1 Silicon	15 P 31.0 Phosphorus	16 S 32.1 Sulfur	17 Cl 35.5 Chlorine	18 Ar 39.9 Argon
4	19 K 39.1 Potassium	20 Ca 40.1 Calcium	21 Sc 45.0 Scandium	22 Ti 47.9 Titanium	23 V 50.9 Vanadium	24 Cr 52.0 Chromium	25 Mn 54.9 Manganese	26 Fe 55.8 Iron	27 Co 58.9 Cobalt	28 Ni 58.7 Nickel	29 Cu 63.5 Copper	30 Zn 65.4 Zinc	31 Ga 69.7 Gallium	32 Ge 72.6 Germanium	33 As 74.9 Arsenic	34 Se 79.0 Selenium	35 Br 79.9 Bromine	36 Kr 83.8 Krypton
5	37 Rb 85.5 Rubidium	38 Sr 87.6 Strontium	39 Y 88.9 Yttrium	40 Zr 91.2 Zirconium	41 Nb 92.9 Niobium	42 Mo 96.0 Molybdenum	43 Tc (97) Technetium	44 Ru 101.1 Ruthenium	45 Rh 102.9 Rhodium	46 Pd 106.4 Palladium	47 Ag 107.9 Silver	48 Cd 112.4 Cadmium	49 In 114.8 Indium	50 Sn 118.7 Tin	51 Sb 121.8 Antimony	52 Te 127.6 Tellurium	53 I 126.9 Iodine	54 Xe 131.3 Xenon
6	55 Cs 132.9 Caesium	56 Ba 137.3 Barium	57 to 71	72 Hf 178.5 Hafnium	73 Ta 180.9 Tantalum	74 W 183.8 Tungsten	75 Re 186.2 Rhenium	76 Os 190.2 Osmium	77 Ir 192.2 Iridium	78 Pt 195.1 Platinum	79 Au 197.0 Gold	80 Hg 200.6 Mercury	81 Tl 204.4 Thallium	82 Pb 207.2 Lead	83 Bi 209.0 Bismuth	84 Po (210) Polonium	85 At (210) Astatine	86 Rn (222) Radon
7	87 Fr (223) Francium	88 Ra (226) Radium	89 to 103	104 Rf (267) Rutherfordium	105 Db (270) Dubnium	106 Sg (269) Seaborgium	107 Bh (270) Bohrium	108 Hs (270) Hassium	109 Mt (278) Meitnerium	110 Ds (281) Darmstadtium	111 Rg (281) Roentgenium	112 Cn (285) Copernicium	113 Nh (286) Nihonium	114 Fl (289) Flerovium	115 Mc (290) Moscovium	116 Lv (289) Livermorium	117 Ts (294) Tennessine	118 Og (294) Oganesson

Metals

Key: 6 — Atomic number; C — Chemical symbol; 12.01 — Atomic mass; Carbon — Name of element

Rare earth elements Lanthanide series	57 La 138.9 Lanthanum	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.1 Ytterbium	71 Lu 175.0 Lutetium
Actinide series	89 Ac (227) Actinium	90 Th 232.0 Thorium	91 Pa 231.0 Protactinium	92 U 238.0 Uranium	93 Np (237) Neptunium	94 Pu (244) Plutonium	95 Am (243) Americium	96 Cm (247) Curium	97 Bk (247) Berkelium	98 Cf (251) Californium	99 Es (252) Einsteinium	100 Fm (257) Fermium	101 Md (258) Mendelevium	102 No (259) Nobelium	103 Lr (260) Lawrencium

METALS
- alkali metal
- alkaline earth metal
- lanthanide
- actinide
- transition metal
- post-transition metal

NON-METALS
- diatomic non-metal
- polyatomic non-metal
- noble gas

OTHER
- metalloid
- unknown chemical properties

This front cover shows a shower of sparks produced by the friction of grinding metal. In this process, fine metal particles become loose. These particles are so hot from the friction that they glow, producing the sparks.

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